Questions are for both separate science and combined science students unless indicated in the question

Q1.

Potash alum is a chemical compound.

Potash alum contains potassium ions, aluminium ions and sulfate ions.

(a) Which **two** methods can be used to identify the presence of potassium ions in potash alum solution?

Tick (✓) **two** boxes. **(separate only)**

Flame emission spectroscopy	✓
Flame test	✓
Measuring boiling point of solution	
Paper chromatography	
Using litmus paper	

(2)

(b) Sodium hydroxide solution is used to test for some metal ions.

Sodium hydroxide solution is added to a solution of potash alum until a precipitate forms.

Complete the sentence. **(separate only)**

Choose the answer from the box.

blue	**brown**	**green**	**white**

The colour of the precipitate formed is white.

(1)

(c) Complete the sentence.

Choose the answer from the box. **(separate only)**

barium chloride solution	**limewater**
red litmus paper	**silver nitrate solution**

Sulfate ions can be identified using dilute hydrochloric acid

and barium chloride solution.

(1)

(d) A solution of potash alum has a concentration of 258 g/dm^3

Calculate the mass of potash alum needed to make 800 cm^3 of a solution of potash alum with a concentration of 258 g/dm^3

Give your answer to 3 significant figures.

Solution of potash alum: 258g in 1 dm^3

1 dm^3 = 1000 cm^3 " in 1000 cm^3

258 g in 1000 cm^3 $\times \frac{800}{1000}$ or 0.8

$\times$ 0.8 ↓ x 800 cm^3

206.4 g

206 g (3 s.f.) in 800 cm^3

$\frac{206\text{ g}}{0.8\text{ dm}^3} = 258\text{ g/dm}^3$

Mass (3 significant figures) = 206 g

(4)

(Total 8 marks)

Q2.

This question is about displacement reactions.

(a) The displacement reaction between aluminium and iron oxide has a high activation energy.

What is meant by 'activation energy'?

The minimum energy needed for the particles to react.

(1)

(b) A mixture contains 1.00 kg of aluminium and 3.00 kg of iron oxide.

The equation for the reaction is:

$$2\ Al + Fe_2O_3 \rightarrow 2\ Fe + Al_2O_3$$

Show that aluminium is the limiting reactant.

Relative atomic masses (A_r): O = 16 Al = 27 Fe = 56

Convert masses of Al and Fe_2O_3 to moles

2 Al : Fe_2O_3

37.04 : 18.75

$\frac{37.04}{18.75} : \frac{18.75}{18.75}$

1.98 : 1

< 2 !

N° moles Al: 1.00 kg = 1.00 × 1000 = 1000 g

A_r (Al) = 27 g/mol

n (Al) = 1000 g / 27 g/mol = 37.04 moles

N° moles Fe_2O_3: 3.00 kg = 3.00 × 1000 = 3000 g

M_r (Fe_2O_3) = (56 × 2) + (16 × 3) = 160 g/mol

n (Fe_2O_3) = 3000 g / 160 g/mol = 18.75 moles

1 mole Fe_2O_3 needs 2 mole Al

18.75 " " 2 × 18.75 " = 37.5 mole Al

37.04 mol Al is less than the 37.5 needed.

(4)

Magnesium displaces zinc from zinc sulfate solution.

(c) Complete the ionic equation for the reaction.

You should include state symbols.

$Mg(s) + Zn^{2+}(aq) \rightarrow$ $Mg^{2+}(aq)$ + $Zn(s)$

(2)

(d) Explain why the reaction between magnesium atoms and zinc ions is both oxidation and reduction.

$Mg \longrightarrow Mg^{2+} + 2e^-$ loses e^-, oxidation

$Zn^{2+} + 2e^- \longrightarrow Zn$ Gains e^-, reduction

(2)

(Total 9 marks)

Q3.

This question is about the halogens.

Table 1 shows the melting points and boiling points of some halogens.

Table 1

Element	Melting point in °C	Boiling point in °C
Fluorine	–220	–188

Chlorine	–101	–35
Bromine	–7	59

(a) What is the state of bromine at 0 °C **and** at 100 °C?

Tick (✓) **one** box.

State at 0 °C	State at 100 °C	
Gas	Gas ✓	☐
Gas	Liquid	☐
Liquid ✓	Gas ✓	✓
Liquid ✓	Liquid	☐
Solid	Gas ✓	☐
Solid	Liquid	☐

(1)

(b) Explain the trend in boiling points of the halogens shown in **Table 1**.

Boiling point increases down the group because the relative formula mass (size of the molecule increases, so the intermolecular forces increase in strength. thus more energy is needed to overcome the intermolecular forces.

(4)

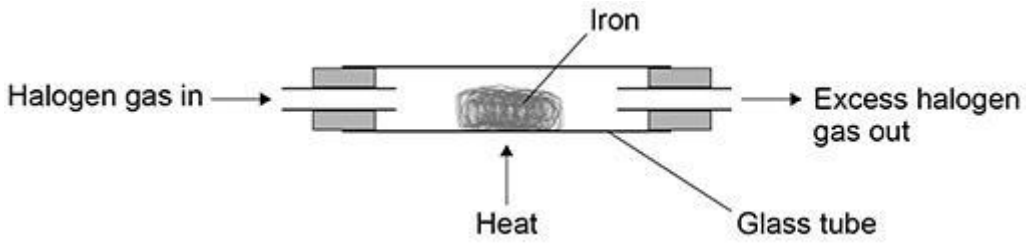
(c) Why is it **not** correct to say that the boiling point of a single bromine molecule is 59 °C?

Boiling point is a bulk property

(1)

Iron reacts with each of the halogens in their gaseous form.

The diagram below shows the apparatus used.



(d) Give **one** reason why this experiment should be done in a fume cupboard.

The gas halogen is toxic

(1)

(e) Explain why the reactivity of the halogens decreases going down the group.

Going down the group, the outer electrons became further from the nucleus.
So the nucleus has less attraction on the outer shell.
So electrons are gained less easily.
(not attracted as strongly)

(3)

(f) A teacher investigated the reaction of iron with chlorine using the apparatus in the above diagram.

The word equation for the reaction is:

iron + chlorine → iron chloride

The teacher weighed:

- the glass tube
- the glass tube and iron before the reaction
- the glass tube and iron chloride after the reaction.

Table 2 shows the teacher's results.

Table 2

	Mass in g
Glass tube	51.56
Glass tube and iron	56.04
Glass tube and iron chloride	64.56

Calculate the simplest whole number ratio of:

moles of iron atoms : moles of chlorine atoms

Determine the balanced equation for the reaction.

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

Mass of Fe = 56.04 − 51.56 = 4.48 g

N° moles Fe = $\frac{4.48\text{ g}}{56\text{ g/mol}}$ = 0.08 mol

Mass of Cl^- = 64.56 − 56.04 = 8.52 g

N° moles Cl^- = $\frac{8.52\text{ g}}{35.5\text{ g/mol}}$ = 0.24 mol

n(Fe) : n(Cl)

0.08 : 0.24

$\frac{0.08}{0.08}$: $\frac{0.24}{0.08}$

1 : 3

$FeCl_3$

Moles of iron atoms : moles of chlorine atoms = 1 : 3

Equation for the reaction $2Fe + 3Cl_2 \rightarrow 2FeCl_3$

(6)

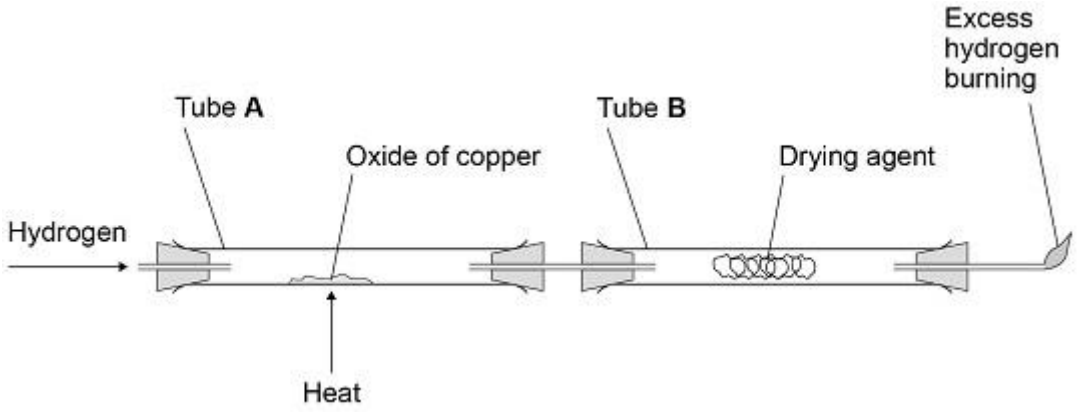
(Total 16 marks)

Q4.

Copper forms two oxides, Cu_2O and CuO

A teacher investigated an oxide of copper.

The following figure shows the apparatus.



This is the method used.

1. Weigh empty tube **A**.

2. Add some of the oxide of copper to tube **A**.

3. Weigh tube **A** and the oxide of copper.

4. Weigh tube **B** and drying agent.

5. Pass hydrogen through the apparatus and light the flame at the end.

6. Heat tube **A** for 2 minutes.

7. Reweigh tube **A** and contents.

8. Repeat steps 5 to 7 until the mass no longer changes.

9. Reweigh tube **B** and contents.

10. Repeat steps 1 to 9 with different masses of the oxide of copper.

(a) Suggest **one** reason why step 8 is needed.

To make sure all the Cu has reacted

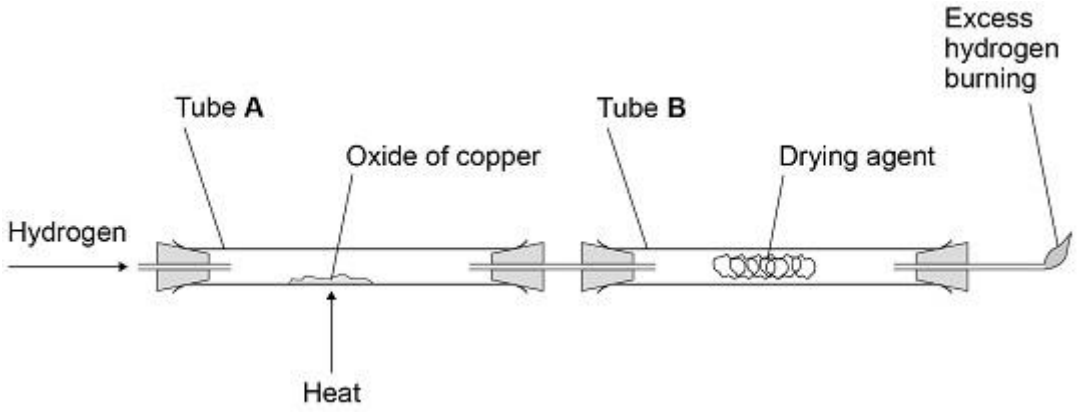
(1)

(b) Explain why the excess hydrogen must be burned off.

To prevent H_2 escaping into the air because H_2 is explosive.

(2)

The figure above is repeated here.



The table below shows the teacher's results.

	Mass in g
Tube **A** empty	105.72
Tube **A** and oxide of copper before heating	115.47
Tube **A** and contents after 2 minutes	114.62
Tube **A** and contents after 4 minutes	114.38
Tube **A** and contents after 6 minutes	114.38
Tube **B** and contents at start	120.93
Tube **B** and contents at end	123.38

When an oxide of copper is heated in a stream of hydrogen, the word equation for the reaction is:

copper oxide + hydrogen → copper + water

(c) Determine the mass of copper and the mass of water produced in this experiment.

Use the table.

Mass of Cu = 114.38 g − 105.72 g = 8.66 g

Mass of H_2O = 123.38 g − 120.93 g = 2.45 g

Mass of copper = 8.66 g g

Mass of water = 2.45 g g

(2)

(d) The teacher repeated the experiment with a different sample of the oxide of copper.

The teacher found that the oxide of copper produced 2.54 g of copper and 0.72 g of water.

Two possible equations for the reaction are:

Equation 1: $Cu_2O + H_2 \rightarrow 2\ Cu + H_2O$

Equation 2: $CuO + H_2 \rightarrow Cu + H_2O$

Determine which is the correct equation for the reaction in the teacher's experiment.

$M_r(H_2O) = (1\times2)+16 = 18$

Relative atomic masses (A_r): H = 1 O = 16 Cu = 63.5

Mass of Cu = 2.54 g

N° moles " = $\frac{2.54}{63.5}$ = 0.04 moles

Mass of H_2O = 0.72 g

N° moles " = $\frac{0.72}{18}$ = 0.04 moles

Cu : H_2O

0.04 : 0.04

1 : 1

∴ Equation 2

$CuO + H_2 \rightarrow Cu + H_2O$

(3)

(Total 8 marks)

Q5.

A student investigated the temperature change in the reaction between dilute sulfuric acid and potassium hydroxide solution.

This is the method used.

1. Measure 25.0 cm^3 potassium hydroxide solution into a polystyrene cup.

2. Record the temperature of the solution.

3. Add 2.0 cm^3 dilute sulfuric acid.

4. Stir the solution.

5. Record the temperature of the solution.

6. Repeat steps 3 to 5 until a total of 20.0 cm^3 dilute sulfuric acid has been added.

(a) Suggest why the student used a polystyrene cup rather than a glass beaker for the reaction.

Polystyrene is a better (thermal) insulator so reduces energy exchange (with the surroundings)

_____

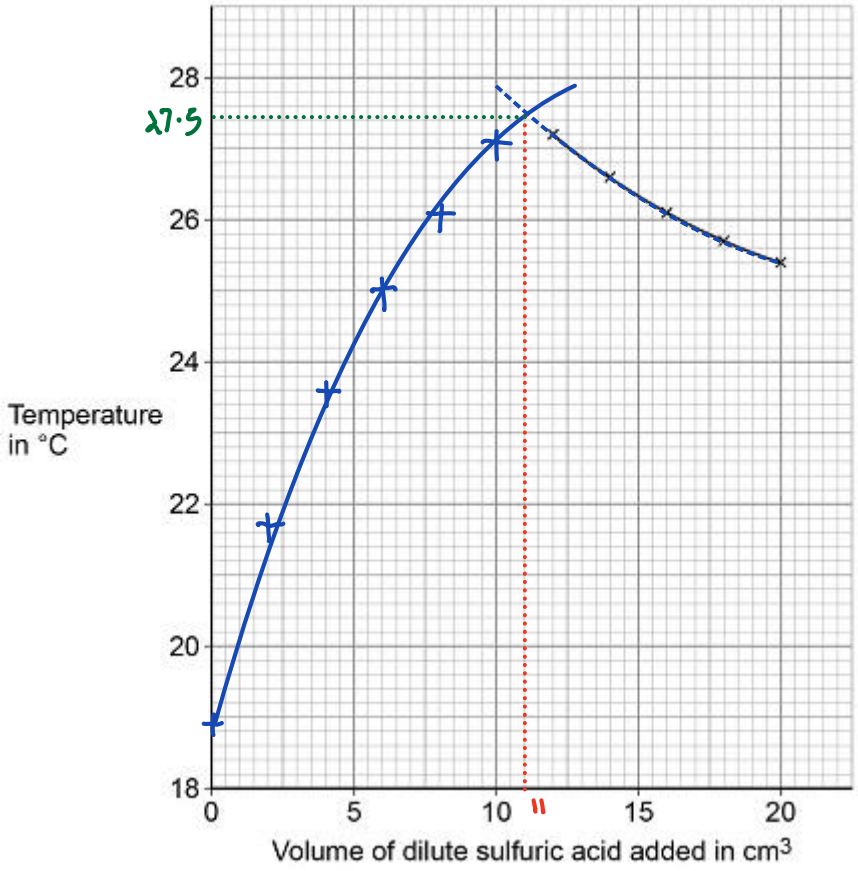
_____

(2)

The following table shows some of the student's results.

Volume of dilute sulfuric acid added in cm^3	Temperature in °C
0.0	18.9
2.0	21.7
4.0	23.6
6.0	25.0
8.0	26.1
10.0	27.1

The figure below shows some of the data from the investigation.



(b) Complete the figure:

- plot the data from the table
- draw a line of best fit through these points

- extend the lines of best fit until they cross.

(4)

(c) Determine the volume of dilute sulfuric acid needed to react completely with 25.0 cm^3 of the potassium hydroxide solution.

Use the figure above.

Volume of dilute sulfuric acid to react completely = 11 cm^3

(1)

(d) Determine the overall temperature change when the reaction is complete.

Use the figure above.

27.5 − 18.9 = 8.6°C

Overall temperature change = 8.6 °C

(1)

(e) The student repeated the investigation.

The student used solutions that had different concentrations from the first investigation.

The student found that 15.5 cm^3 of 0.500 mol/dm^3 dilute sulfuric acid completely reacted with 25.0 cm^3 of potassium hydroxide solution.

The equation for the reaction is:

$$2\ KOH + H_2SO_4 \rightarrow K_2SO_4 + 2\ H_2O$$

Calculate the concentration of the potassium hydroxide solution in mol/dm^3 and in g/dm^3

Relative atomic masses (A_r): H = 1 O = 16 K = 39 **(separate only)**

N° moles H_2SO_4 = $\frac{15.5}{1000} \times 0.500 = 7.75 \times 10^{-3}$ mol

2 KOH reacts with 1 H_2SO_4

N° moles KOH = $7.75 \times 10^{-3} \times 2$

= 0.0155 mol

Volume " = 25.0 cm^3 = $\frac{25}{1000}$ = 0.025 dm^3

Concentration KOH = $\frac{0.0155 \text{ mol}}{0.025 \text{ } dm^3}$

= 0.62 mol/dm^3

M_r (KOH) = 39 + 16 + 1 = 56 g/mol

Mass of KOH = 0.62 mol × 56 g/mol

= 34.72 g

Conc^n = 34.72 g/dm^3

Concentration in mol/dm^3 = 0.62 mol/dm^3

Concentration in g/dm^3 = 34.72 g/dm^3

(6)

(Total 14 marks)

Q6.

This question is about elements in Group 1.

A teacher burns sodium in oxygen.

(a) Complete the word equation for the reaction.

sodium + oxygen → sodium oxide

(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. (alkaline) pH > 7

What is the pH value of the solution?

Tick **one** box.

1	4	7	13
☐	☐	☐	☑
acid		neutral	alkaline

(1)

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

Give the name of the substance.

Sodium hydroxide

(1)

(e) All alkalis contain the same ion.

What is the formula of this ion?

Tick **one** box.

- 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?

NaOH: 40g per 1 dm^3

40g " 1000 cm^3 ÷4

÷4 x 250

10g

Mass = 10 g

(2)

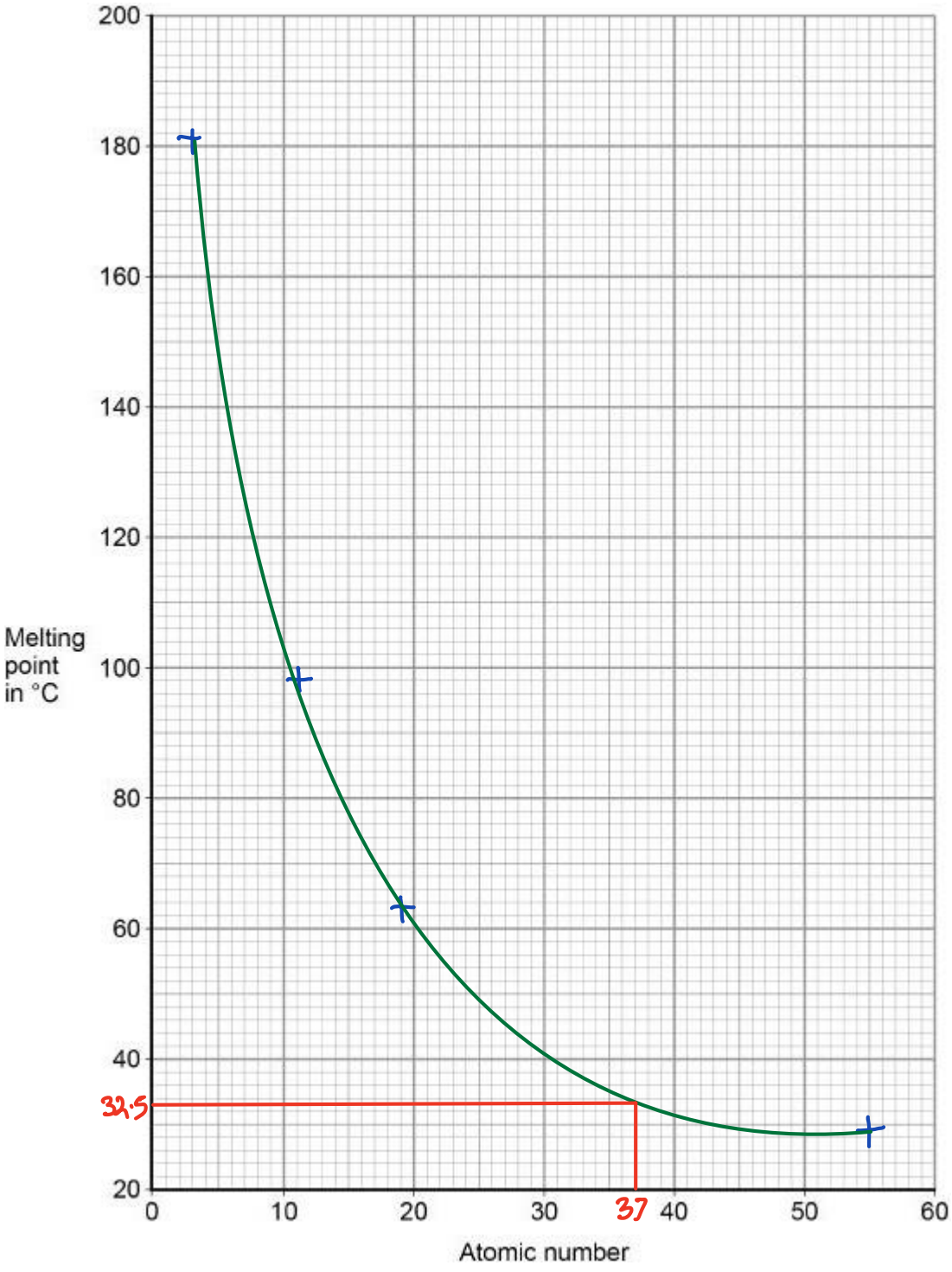
(g) The melting points of the elements in Group 1 show a trend.

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.



(2)

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

Use the graph above.

Melting point = 32.5 °C

(1)

(Total 10 marks)

Q7.

Titanium is a transition metal.

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**.

Chlorine is toxic

(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.

Na and H_2O would produce a very exothermic reaction and produces a corrosive solution.

(1)

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

Argon is unreactive (inert). Oxygen from the air would react with the sodium.

(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.

Metal chlorides are usually ionic, so metal chlorides

are solid at room temperature.

(3)

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

OIL RIG

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

Why is this an oxidation reaction?

Na atoms lose electrons

(1)

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

$Na \longrightarrow \underline{Na^+} + \underline{e^-}$

(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.

Mass of $TiCl_4$ = 40 kg = 40000 g

$M_r(TiCl_4) = 48 + (35.5 \times 4) = 190$ g/mol

N° moles of $TiCl_4 = \frac{40000}{190} = 210.53$ moles

Mass of Na = 20 kg = 20000 g

N° moles Na $= \frac{20000}{23} = 869.57$ moles.

1 $TiCl_4$ reacts with 4 Na

210.53 mol " " $210.53 \times 4 = 842.12$ mol Na

There are 869.6 mol Na, so $TiCl_4$ is limiting reagent.

(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. **(separate only)**

% Yield = Actual amount / Theoretical amount × 100%

92.3% = $\frac{x}{13.5}$ × 100

Mass of titanium = 12.5 kg

$\frac{92.3}{100} \times 13.5 = x$

$x = 12.46\,g$

(2)

(Total 15 marks)

Q8.

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	☑
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 mol CO : 2 H_2

$\times 2 \times 10^3$ ↓ ↓ $\times 2 \times 10^3$

x : 4.0×10^3

2.0×10^3

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. reactant ⇌ (exo / endo) product + heat

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

If temperature increases, the equilibrium shifts to the left to oppose the change. Endothermic reaction favoured, yield reduced.

(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.

$CO_{(g)} + 2\,H_{2(g)} \rightleftharpoons CH_3OH_{(g)}$

N° mol gas: 1 + 2 = 3; 1

Explain why.

Yield: Equilibrium moves to the product side as there are fewer moles of gas, reducing volume and opposing higher pressure.

Rate: Increased pressure causes more collisions per unit time as there are more molecules per unit volume

(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.

Catalyst provides a different reaction pathway which has a lower activation energy.

(2)

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

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

Less energy is needed.

(1)

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

Catalysts do not affect the yield, only the rate at which the equilibrium yield is reached.

(1)

(Total 12 marks)

Q9.

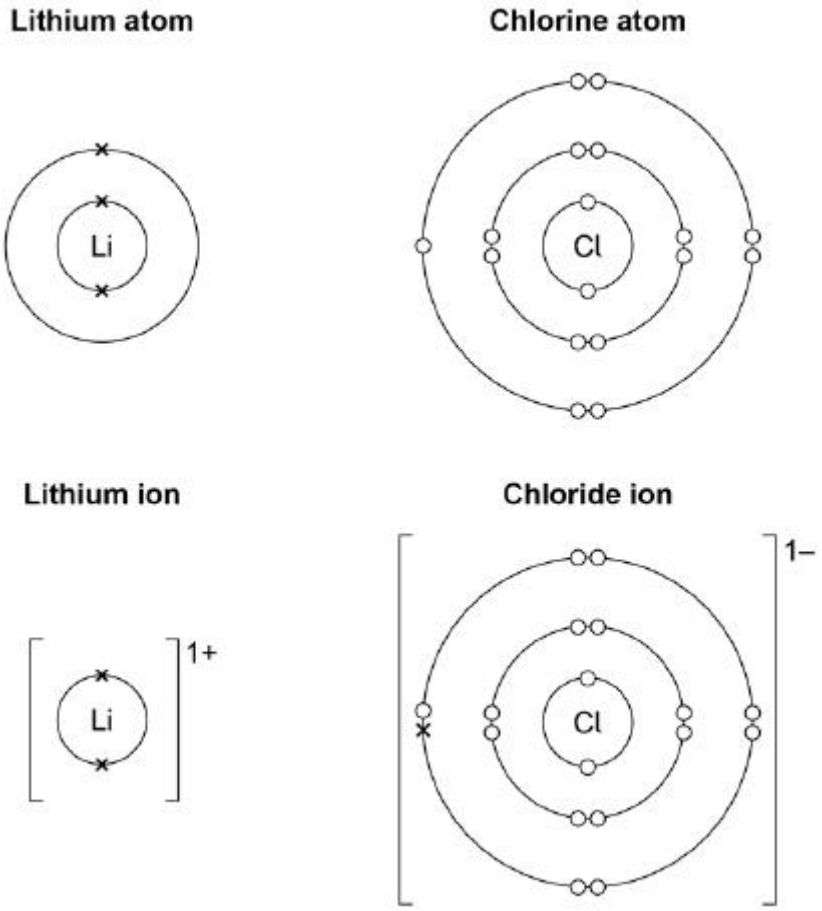
This question is about metal compounds.

(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.

Lithium atom loses an electron to form a positive ion.
This electron is transferred to the chlorine atom which gains the electron to form a negative ion.
Opposite charges attract each other to form an ionic bond.

(4)

Zinc sulfate can be made by two methods.

The equations for the two methods are:

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

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

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

Use the equation:

percentage atom economy =

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

161 / 81 / 98

Give your answer to 3 significant figures.

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

Method 1: % Atom Economy = $\frac{161}{81+98} \times 100$

$= 89.94\%$

Percentage atom economy = 89.9% (3 s.f.) %

(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. **(separate only)**

As it is more sustainable, producing less waste.

(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?

$ZnSO_4$ 80g per dm^3

80g per 1000 cm^3

$\frac{80}{1000}$ g per cm^3

50 cm^3 contains $0.08 \times 50 g/cm^3 = 4.0$ g

Mass = 4.0 g

(2)

(Total 10 marks)

Q10.

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.

Heat with a water bath.

(1)

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

To make sure that all the iodine reacts. As excess iodine would remain in solution and could not be filtered off. Whereas excess zinc can be filtered off.

(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

Mass of I_2 = 0.500g $M_r(I_2) = 2 \times 127 = 254$ g/mol

N° moles " = $\frac{0.500g}{254 g/mol} = 1.97 \times 10^{-3}$ mol

1 mole Zn reacts with 1 mole I_2

∴ N° moles Zn required = 1.97×10^{-3} mol

Mass " = $1.97 \times 10^{-3} \times 65$

= 0.128 g

Minimum mass of zinc = 0.128 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? **(separate only)**

$$\% \text{ yield} = \frac{\text{obtained mass}}{\text{theoretical mass}} \times 100$$

$$92 = \frac{12.5}{x} \times 100$$

$$92 = \frac{1250}{x} \qquad x = \frac{1250}{92} = 13.6 \text{ g}$$

Maximum theoretical mass = 13.6 g

(3)

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

Some product may be lost on separation.

(1)

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

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

$M_r(ZnI_2) = 65 + (2 \times 127) = 319$ g/mol

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

N° moles of ZnI_2 = $\frac{250}{1000} \times 0.100$ mol/dm^3

= 0.025 moles

Mass of ZnI_2 = 0.025 mol × 319 g/mol

= 7.98 g

Mass = 7.98 g

(3)

(Total 14 marks)

Q11.

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.

Salt, sodium chloride NaCl

(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.

Since desalination requires large amounts of energy, it is expensive.

(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?

To remove solids

(1)

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

Why is chlorine gas used to treat water?

To kill and sterilise the water microorganisms

(1)

(e) Describe a test for chlorine gas.

Give the result of the test if chlorine is present.

Test

Use damp litmus paper

Result Cl_2 bleaches the litmus paper

(2)

Some students investigated different water samples.

The table shows some of their results.

Water	**pH**	**Mass of dissolved solid in g / dm³**
Tap water	6.5	0.5
Seawater	8.1	35.0
Pure water	7.0	0.0

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

(2)

(g) What mass of dissolved solid is present in 100 cm³ of the sample of tap water?

Tick (✓) **one** box.

Tap water: 0.5 g / dm³

1 dm³ = 1000 cm³

$\text{Mass} = \frac{100\ cm^3}{1000} \times 0.5\ g/dm^3$

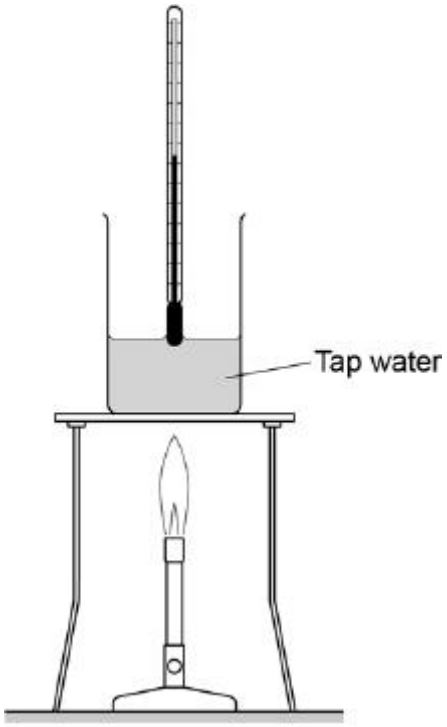
$= 0.05\ g$

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.



The students made a mistake setting up the apparatus.

What mistake did the students make?

The bulb of the thermometer is not immersed.

(1)

(Total 10 marks)

Q12.

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.

1 Add excess copper carbonate to dilute hydrochloric acid.

2 Filter to remove excess copper carbonate

3 Heat filtrate to evaporate some water (heat to the point of crystallisation)

4 Leave to cool, so crystals

$M_r(CaCl_2) = 63.5 + (35.5 \times 2) = 134.5$

$M_r(CuCO_3) = 63.5 + 12 + (16 \times 3) = 123.5$ **(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.

N° of moles of $CaCl_2$: $\frac{11.0}{134.5} = 0.0818$ moles

1 mole of $CuCl_2$ is produced from 1 mole $CaCO_3$

∴ N° moles $CuCO_3$ needed = 0.0818 mol

Mass " " = 0.0818×123.5 g/mol

Mass of copper carbonate = 10.1 g

(4)

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

Calculate the mass of copper chloride the student actually produced. **(separate only)**

$\%\ Yield = \frac{actual}{theory} \times 100$. $79.1 = \frac{x}{11} \times 100$

$x = 0.791 \times 11$
$= 8.70$

Actual mass of copper chloride produced = 8.70 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. **(separate only)**

% Atom economy = $\frac{CuCl_2}{CuO + 2HCl} \times 100$

$= \frac{134.5}{79.5 + (2 \times 36.5)} \times 100$

$= 88.2\ \%$

Percentage atom economy = 88.2 %

(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. **(separate only)**

Atom economy for Reaction 1 is lower since the additional product CO_2 is produced.

(1)

(Total 14 marks)

Q13.

Sodium carbonate reacts with dilute hydrochloric acid:

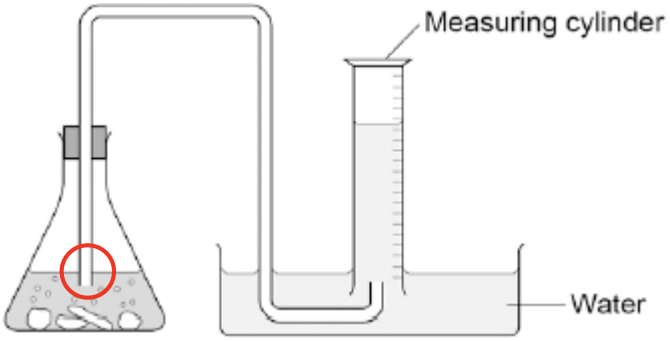
$$Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2O + CO_2$$

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.

The delivery tube sticks into the acid.
The acid would go into the water.

(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.

Gas lost before being placed
Lock.

(1)

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

All of the acid has reacted, no further CO_2 produced.

(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?

Take more readings between 0.23 and 0.34 g Na_2CO_3

(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. **(separate only)**

1 mole gas occupies 24 dm^3 (24×1000 cm^3)

24 000 cm^3 = 1 mole gas

95 cm^3 = $\frac{1}{24000} \times 95 = 3.96 \times 10^{-3}$ mol

3.96×10^{-3} 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.

Use a pipette to measure the acid.

Because it is a more accurate volume than a measuring cylinder.

(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.

Correct, they should be collected because CO_2 is left in the flask at the end, and it has the same volume as the air collected.

(2)

(Total 11 marks)