Please check the examination details below before entering your candidate information

Candidate surname	Other names

Centre Number | Candidate Number

Pearson Edexcel Level 1/Level 2 GCSE (9–1)

Monday 22 May 2023

Morning (Time: 1 hour 45 minutes)

Paper reference **1CH0/1H**

Chemistry

PAPER 1

Higher Tier

You must have:
Calculator, ruler

Total Marks

Instructions

- Use **black** ink or ball-point pen.
- **Fill in the boxes** at the top of this page with your name, centre number and candidate number.
- Answer **all** questions.
- Answer the questions in the spaces provided – *there may be more space than you need.*
- Calculators may be used.
- Any diagrams may NOT be accurately drawn, unless otherwise indicated.
- You must **show all your working out** with **your answer clearly identified** at the **end of your solution**.

Information

- The total mark for this paper is 100.
- The marks for **each** question are shown in brackets – *use this as a guide as to how much time to spend on each question.*
- In questions marked with an **asterisk** (*****), marks will be awarded for your ability to structure your answer logically, showing how the points that you make are related or follow on from each other where appropriate.
- There is a periodic table on the back cover of the paper.

Advice

- Read each question carefully before you start to answer it.
- Try to answer every question.
- Check your answers if you have time at the end.

Turn over ▶

Pearson

Answer ALL questions. Write your answers in the spaces provided.

Some questions must be answered with a cross in a box ☒. If you change your mind about an answer, put a line through the box ☒ and then mark your new answer with a cross ☒.

1 (a) Figure 1 shows information about two isotopes of hydrogen, **A** and **B**.

Complete the table to show the number of subatomic particles in each isotope.

(2)

	isotope A	**isotope B**
atomic number	1	1
mass number	1	2
number of protons	1	1
number of electrons	1	1
number of neutrons	0	1

Figure 1

(b) Hydrogen gas and oxygen gas are used in a hydrogen–oxygen fuel cell.

Separate containers of hydrogen and oxygen are used to supply the gases.

A student tests the voltage supplied by the fuel cell every 15 minutes.

The results are shown in Figure 2.

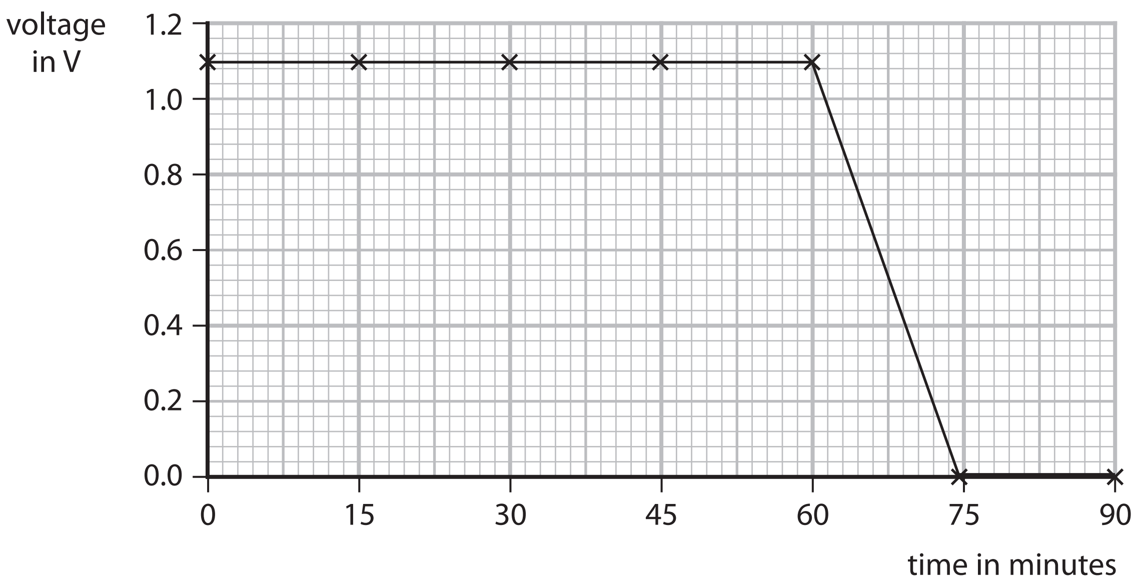

Figure 2

Describe what Figure 2 shows about how the voltage of this fuel cell varies with time.

(2)

An description including
• voltage {stays same/ constant/ steady/ stable/ at 1.1(V)} for 60 minutes/ 1hr (1)

• and then falls to 0(V) (during the time 60-75 mins) (1)

(c) A chemical cell is made by placing two electrodes into an aqueous electrolyte.

Figure 3 shows a chemical cell.

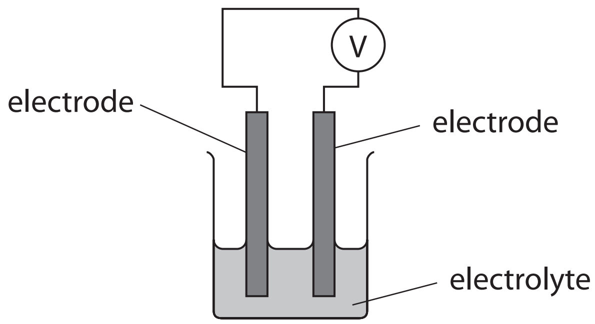

Figure 3

State why sodium and sulfur electrodes are **not** suitable for this experiment.

(2)

sodium

• sodium would react with the {electrolyte / water/ solution} (1)

sulfur

• sulfur does not conduct (electricity) (1)

(Total for Question 1 = 6 marks)

2 In an experiment, powdered calcium hydroxide was added to dilute hydrochloric acid and the pH was measured.

The method used was

step 1 measure 200 cm^3 dilute hydrochloric acid into a beaker

step 2 add 0.1 g of powdered calcium hydroxide to the beaker

step 3 find the pH of the mixture

step 4 repeat steps 2 and 3 until the pH stops changing.

(a) State what should be done after **step 2** to make sure that any reaction is complete.

(1)

stir/ swirl/ shake (the beaker)

....

(b) Complete the word equation for the reaction.

(2)

calcium hydroxide + hydrochloric acid → calcium chloride (1) + water (1)

(c) Which row of the table shows the state symbols for powdered calcium hydroxide and dilute hydrochloric acid in the balanced chemical equation?

(1)

		calcium hydroxide	hydrochloric acid
☐	**A**	aq	l
☐	**B**	l	aq
☒	**C**	s	aq
☐	**D**	s	l

(d) The results of the experiment are shown in Figure 4.

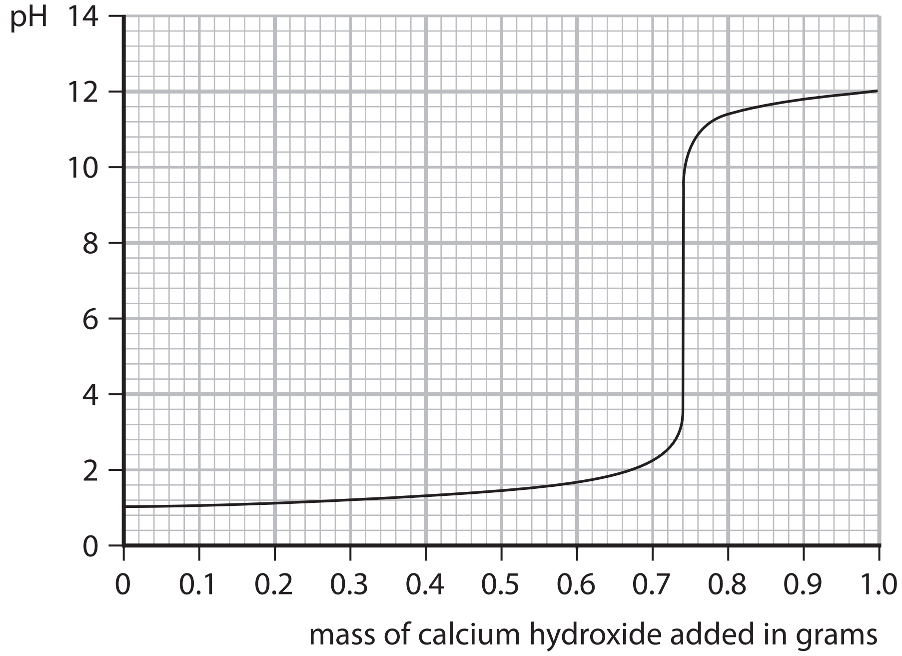

Figure 4

(i) Using Figure 4, give the pH of the acid at the start of the experiment.

(1)

pH = 1

(ii) Using Figure 4, give the mass of calcium hydroxide required to make a neutral mixture.

(1)

mass of calcium hydroxide = 0.74 (g) g

(iii) Explain why the pH starts at a low value and ends at a higher value.

(3)

START
• solution is acidic / acids have low pH / high {concentration/ amount} of H+ ions/ excess H+ ions (1)

REACTION
• neutralisation/ H+ + OH------> H2O/ {the hydroxide/ alkali} reacts with the {acid/ H+} (1)

END
• {amount/ concentration} of H+ ions has reduced/ {amount/ concentration} of OH- ions has increased / excess OH- ions/ (excess of) hydroxide ions have pH > 7 (1)

(Total for Question 2 = 9 marks)

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3 Figure 5 shows part of the reactivity series of metals.

calcium
magnesium
aluminium
zinc
iron
copper
silver
gold

↑ increasing reactivity

Figure 5

(a) Which metal reacts when added to cold water?

(1)

- ☒ **A** calcium
- ☐ **B** copper
- ☐ **C** gold
- ☐ **D** silver

(b) A student investigates the reactivity of four different metals.

The student adds an equal-sized piece of each metal to separate test tubes containing dilute hydrochloric acid.

The student's observations for zinc and copper are recorded in Figure 6.

metal	observations
magnesium	
zinc	bubbles produced at a steady rate test tube feels slightly warm
iron	
copper	no reaction

Figure 6

(i) Use the information in Figure 5 and in Figure 6 to predict the observations for the reactions of magnesium and of iron with dilute hydrochloric acid.

(2)

magnesium

many bubbles / bubbles produced quickly / bubbles vigorously
OR
test tube feels hot / warm / warmer than with zinc (1)

iron

few bubbles / bubbles produced slowly / some bubbles
OR
test tube feels very slightly warm (1)

(ii) When metals react with acids, hydrogen gas is produced.

Describe the test to show that the gas is hydrogen.

(2)

a description to include the following points

• apply lighted splint (to the gas) (1)

• (squeaky) pop (1)

(iii) When magnesium reacts with hydrochloric acid, magnesium chloride and hydrogen are formed.

Complete the balanced equation for the reaction.

(2)

Mg + 2 HCl → $MgCl_2$ + H2

(c) An excess of magnesium is added to some dilute hydrochloric acid of pH 2. The mass of hydrogen gas produced is measured.

The experiment is repeated with excess magnesium but with the same volume of dilute hydrochloric acid of pH 1.

(i) State how many times greater the concentration of hydrogen ions is in the acid of pH 1 than in the acid of pH 2.

(1)

ten (times) / 10 (x) / (x) 10

(ii) With the acid of pH 2, the mass of hydrogen gas produced when the reaction is complete is 0.005 g.

Predict the mass of hydrogen gas produced in the reaction with acid of pH 1.

(1)

0.05 (g)
OR
0.005 x factor from (c)(i)

mass = ... g

(Total for Question 3 = 9 marks)

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4 There are several stages to the production of sulfuric acid in industry.

(a) Sulfur dioxide is required for the production of sulfuric acid.

Sulfur dioxide can be obtained by heating copper sulfide, Cu_2S, in excess air.

$$Cu_2S + O_2 \rightarrow 2Cu + SO_2$$

Calculate the atom economy for the production of sulfur dioxide, SO_2, in this reaction.

(relative atomic mass: Cu = 63.5
relative formula masses: O_2 = 32.0, Cu_2S = 159.0, SO_2 = 64.0)

Give your answer to two significant figures.

(4)

total mass = 191 (1)

• $\frac{64}{191}$ = 0.33507 (1)

• x 100 = 33.507 (1)

atom economy = 34 %

(b) In one stage vanadium oxide, V_2O_5, is used.

Based on the position of vanadium in the periodic table, which row shows the most likely melting point of vanadium and colour of vanadium oxide?

(1)

		melting point of vanadium in °C	colour of vanadium oxide
☐	**A**	50	white
☐	**B**	1910	white
☐	**C**	50	orange
☒	**D**	1910	orange

(c) The equation shows a reaction forming sulfuric acid.

$$SO_3 + H_2O \rightarrow H_2SO_4$$

(i) Calculate the maximum mass of sulfuric acid that could be produced from 400 tonnes of sulfur trioxide, SO_3.

(relative formula masses: SO_3 = 80, H_2SO_4 = 98)

(2)

$\frac{400}{80}$ = 5 (1)

5 x 98 = 490 (1)

maximum mass of sulfuric acid = 490 tonnes

(ii) Using a different amount of sulfur trioxide, it was calculated that 700 tonnes of sulfuric acid could be made.

The actual mass produced was 672 tonnes.

Calculate the percentage yield of sulfuric acid.

(2)

$\frac{672}{700}$ = 0.96 (1)

x 100 = 96 (1)

percentage yield = 96

(iii) State **two** reasons why the percentage yield is less than 100%.

(2)

1 incomplete reaction

2

• loss {of substance/reactant/product} (during practical)/ practical losses

(Total for Question 4 = 11 marks)

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5 (a) Ammonia is manufactured in the Haber process by the reversible reaction between nitrogen and hydrogen.

(i) Write the balanced equation for the reversible reaction between nitrogen and hydrogen to make ammonia, NH_3.

(3)

N2 + 3H2 ------>2NH3 (3)
<------

(ii) Which row shows the typical conditions of temperature and pressure used in the Haber process?

(1)

		temperature in °C	pressure in atmospheres
☐	**A**	250	100
☐	**B**	250	200
☐	**C**	450	500
☒	**D**	450	200

(iii) In the Haber process, iron is added to the vessel where the nitrogen and hydrogen react.

State the purpose of the iron.

(1)

catalyst/ increase rate of reaction(s)/ lower activation energy/ increase rate of attainment of equilibrium

(iv) The reaction between nitrogen and hydrogen to make ammonia can reach dynamic equilibrium.

The reaction gives out heat.

Explain how the position of equilibrium changes if the temperature is decreased.

(2)

moves in exothermic direction (1)

• moves {right/ forwards / towards ammonia/ to products} (1)

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P 7 2 6 3 1 A 0 1 2 3 2

(b) Compound **A** is a dark brown gas.

Compound **B** is a colourless gas.

Two molecules of **A** combine to form one molecule of **B** in a reversible reaction.

You are given

- a sealed glass tube containing an equilibrium mixture of **A** and **B**
- a beaker
- a kettle
- some ice

At room temperature, the equilibrium mixture is a pale brown colour.

Devise an experiment to show how the position of equilibrium of this reaction is affected by temperature.

The sealed tube must **not** be opened.

(3)

METHOD OF HEATING AND COOLING
• put tube into hot water (1)

• then into cold water/ add cold water/ add ice (1)

OBSERVATIONS
• colour goes darker

AND
colour goes lighter/ colourless

(Total for Question 5 = 10 marks)

6 A student investigates the mass of copper produced when copper chloride solution in a beaker is electrolysed using inert electrodes.

(a) Where is copper formed during the electrolysis?

(1)

☐ **A** at the anode

☐ **B** at the bottom of the beaker

☒ **C** at the cathode

☐ **D** on the surface of the electrolyte

(b) The student investigated the change in the mass of copper formed when the current was altered.

The results are shown in Figure 7.

current in A	mass of copper formed in g
0.0	0.000
0.2	0.040
0.4	0.080
0.6	0.118
0.8	0.158
1.0	0.196

Figure 7

(i) State and explain the trend shown in these results.

(3)

as current increases mass increases / the mass is proportional to the current (1)

• because the higher the current the more electrons (per second) (1)

• so more copper ions {are reduced/ gain electrons/ are discharged} (1)

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(ii) Describe how, after the power supply has been switched off, the mass of copper formed can be measured.

(2)

(rinse and) dry {electrode / cathode} (1)

• measure mass of {electrode/ cathode} (on a balance) (and subtract original mass) (1)

(c) In another experiment, 74 mg of copper is formed.

Calculate the number of copper atoms in 74 mg of copper.

(relative atomic mass Cu = 63.5; Avogadro constant = 6.02×10^{23})

(3)

mass copper in g = $\frac{74}{1000}$

= 0.074 / 7.4 x 10-2 g (1)

• amount of copper = $\frac{0.074}{63.5}$

= 0.001165…./ 1.165…. x 10-3 mol (1)

• number of atoms = 0.001165… x 6.02 x 1023

number of atoms = 7.015 x 10^{20}

(Total for Question 6 = 9 marks)

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7 Titration is used to carry out some neutralisation reactions.

(a) Ammonium nitrate can be made by neutralisation.

(i) State the name of the two reactants that are neutralised to make the salt ammonium nitrate.

(2)

ammonia (solution) (1) and • nitric acid (1

(ii) Ammonium nitrate is a fertiliser.
Another fertiliser is ammonium phosphate.

Which elements are combined in ammonium phosphate?

(1)

☐ **A** nitrogen, oxygen and phosphorus only

☐ **B** hydrogen, oxygen and phosphorus only

☐ **C** hydrogen, nitrogen and phosphorus only

☒ **D** hydrogen, nitrogen, oxygen and phosphorus only

(b) Titrations involve aqueous solutions and the use of burettes.

(i) Figure 8 shows readings on part of a burette at the start and at the end of a titration.

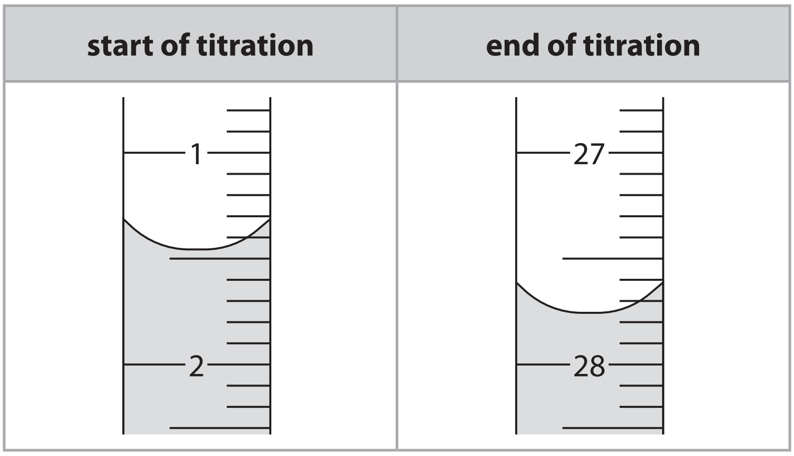

Figure 8

Calculate the volume of solution added from this burette.

Give your answer to a suitable number of decimal places.

(2)

two readings: 27.75 and 1.45 (1)

• 27.75 – 1.45 = 26.30 (1)

volume = 26.30 cm^3

P 7 2 6 3 1 A 0 1 7 3 2

Turn over

(ii) A student carries out a titration four times.

The volumes from the student's results table are shown in Figure 9.

	rough	titration 1	titration 2	titration 3
volume in cm^3	25.90	24.90	24.60	25.00
used to calculate mean volume		✓		✓

Figure 9

Tick the volumes that should be used to calculate the mean volume.

(1)

(iii) Figure 10 shows the burette and flask prepared for use by the student. The burette is supported vertically by a clamp that is not shown in the diagram.

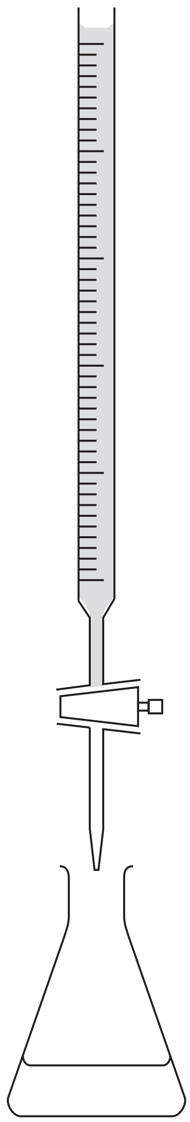

Figure 10

The student wrote a description of how they used the burette.

I took the burette from the cupboard. I closed the tap and filled the burette with the correct solution. I added the solution from the burette drop by drop to the flask until the indicator changed colour.

Give **three** improvements to the way that the student used the **burette**.

(3)

1 rinse the burette with titrant (1)

2 • fill the jet / tip (1)

3 • do not fill burette over 0 cm3 / overfill/ run liquid out until volume at or below 0 cm3 (1)

(c) In a titration a student placed alkali in the flask.

By mistake a few drops of litmus **and** a few drops of phenolphthalein were added to the flask.

The student added acid to the flask until the mixture was acidic.

Predict the colour change that would be seen.

(1)

from pink-blue/ pink-purple/ blue-purple/ purple to red

(d) In a titration a student rinsed out the flask with distilled water and did not dry it.

They used the flask for titration, adding the solution from the burette until the indicator changed colour.

State the effect, if any, on the titre volume of using the wet flask rather than a dry flask.

(1)

None / volume the same

(Total for Question 7 = 11 marks)

8 Crystals of copper sulfate are prepared by reacting copper oxide, a base, with dilute sulfuric acid.

(a) Name the other product of this reaction.

(1)

water

(b) During the experiment, a spatula measure of copper oxide, a black powder, is added to warm, dilute sulfuric acid in a beaker.

When the mixture is stirred, the black powder disappears and the mixture turns pale blue.

The student then adds more copper oxide until the maximum amount of copper sulfate is formed without wasting copper oxide.

Explain how the student knows when to stop adding copper oxide.

(3)

OBSERVATION

• when some powder remains in the beaker (after stirring) (1)

COPPER OXIDE

• there is an excess of copper oxide (1)

ACID

• all the acid {is neutralised/ has reacted}/ no acid remains (1)

(c) The reaction produces an aqueous solution of copper sulfate.

What is the best way to obtain crystals of copper sulfate from an aqueous solution?

(1)

☐ **A** pour the solution through filter paper in a funnel

☐ **B** heat the solution with a Bunsen burner until dry

☒ **C** heat the solution using a water bath

☐ **D** leave the solution in a cold, damp place

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(d) When some water is removed from the aqueous solution of copper sulfate, crystals of copper sulfate are made.

Describe how the arrangement and movement of the particles change as crystals are formed from a solution.

(3)

SOLUTION

• (the ions) are (freely) moving (1)

• (the ions) are randomly arranged (1)

SOLID

• (the ions) are fixed/ not moving/ vibrating (1)

• (the ions) are in a regular arrangement/ lattice/ giant structure (1)

(e) In this reaction, copper oxide, CuO, forms copper sulfate, $CuSO_4$.

Explain, in terms of electrons, whether the copper in copper oxide has been oxidised, has been reduced, or has not been oxidised or reduced.

(2)

the copper (ions are) neither oxidised nor reduced (1)

• the copper (ions) do not lose or gain electrons/ Cu2+ present at start and end (1)

(f) In another experiment, a copper sulfate solution with a concentration of 39.875 g dm^{-3} is used.

Calculate the mass of copper sulfate dissolved in 0.300 dm^3 of this solution.

(1)

11.9625 with or without working scores 1

11.9625

mass = 11.9 g

(Total for Question 8 = 11 marks)

9 (a) Figure 11 shows the structure of a molecule of compound **S**.

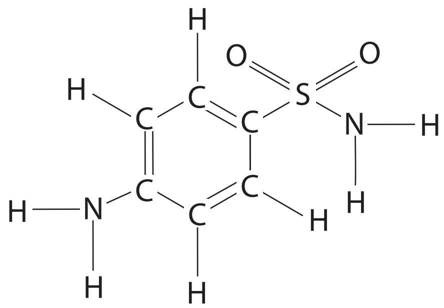

Figure 11

(i) Use Figure 11 to deduce the empirical formula of compound **S**.

(1)

C6H8N2SO2

(ii) The melting points of three samples of **S** are shown in Figure 12.

sample	melting point in °C
A	160–164
B	166
C	163–165

Figure 12

State whether each of these samples, **A**, **B** and **C**, is pure or impure and justify your answers using the information in Figure 12.

(3)

B is pure and A is impure and C is impure (1)

• B has a sharp/ single melting point (1)

• A and C have melting points {over a range / lower than (the sharp melting point of) B} (1)

(b) A scientist uses chromatography in an investigation of compound **S**.

In the conditions used, compound **S** has an R_f value of 0.22.

Calculate the distance the spot of compound **S** moves if the solvent front has moved by 2.4 cm.

(2)

0.528/ 0.53 with or without working scores 2

• distance = Rf x solvent front distance/ 0.22 x 2.4 (1)

= 0.528/ 0.53 (cm) (1)

distance = 0.53 cm

*(c) A solution of sodium chloride in water needs to be separated to obtain a sample of pure, dry sodium chloride and a sample of pure water.

Figure 13 shows the boiling points of sodium chloride and water.

substance	boiling point in °C
sodium chloride	1465
water	100

Figure 13

Explain this difference in boiling points in terms of the structure and bonding of sodium chloride and water and how this difference is used to choose a method to separate sodium chloride solution into pure, dry sodium chloride and pure water.

(6)

SODIUM CHLORIDE
- ionic compound
- giant lattice
- positive (sodium) ions and negative (chloride) ions
- strong electrostatic attraction between ions
- lots of energy to overcome attraction/ bonds

WATER
- simple covalent/ molecular
- strong covalent bonds between atoms in a molecule
- weak forces between molecules
- little energy needed to overcome the intermolecular forces

SEPARATION
- use distillation – with condenser or simple apparatus: delivery tube into test tube in ice water
- water has much lower boiling point
- water will distil but sodium chloride will not
- water collected after being condensed
- sodium chloride remains in flask

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(Total for Question 9 = 12 marks)

10 (a) Buildings sometimes have water sprinklers to put out fires.

The pipes in some water sprinklers are filled with nitrogen gas to prevent corrosion when the system is not in use.

(i) State what is meant by the term **corrosion**.

(2)

a description to include
• reaction of a metal (1)

• with oxygen/ oxidation (1)

(ii) Nitrogen can be made from sodium azide, NaN_3.

$$2NaN_3 \rightarrow 2Na + 3N_2$$

Calculate the maximum volume, in cm^3, of nitrogen produced from 110 g of sodium azide.

(relative formula mass: NaN_3 = 65;

1 mol of gas occupies 24 dm^3 in the conditions used)

(4)

amount of sodium azide = 110/65
= 1.692… mol (1)
• amount of nitrogen = 3/2 x 1.692
= 2.538… mol (1)

• volume of nitrogen = 24 x 2.538
= 60.923… dm3 (1)

• volume in cm3 = 60.923 x 1000
= 60923/ 61 000 cm3 (1)

volume = 6100 cm^3

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*(b) Compare and contrast the properties and uses of pure aluminium and pure copper with the alloys of aluminium and the alloys of copper.

Include in your answer an **explanation** of the similarities and the differences in the properties and the uses of a pure metal and its alloy.

(6)

METAL PROPERTIES

- good conductors of heat
- good conductors of electricity
- malleable
- ductile
- shiny (when pure)
- copper is unreactive
- aluminium samples are resistant to corrosion due to oxide layer
- aluminium has low density

ALLOY PROPERTIES

- good conductors of heat
- good conductors of electricity
- malleability is lower than pure metal
- ductility lower than pure metal
- specific properties of magnalium e.g. low density (ignore light)
- specific properties of brass

METAL USES

- aluminium for cooking foil, food trays
- copper for water pipes, electrical wires, roofing

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ALLOY USES
• aluminium alloys for aircraft parts, vehicles, ladders
(6)
A01 1

• copper alloys for coins, brass instruments, jewellery, plug prongs

COMPARE/ SIMILARITIES between metal and alloy
• both exist as lattices of ions with delocalised electrons
• conduct heat because delocalised electrons in both to carry thermal energy
• conduct electricity because delocalised electrons in both can move
• high melting and boiling points due to strong metallic bonds

CONTRAST/ DIFFERENCES between metal and alloy
• layers of ions/ atoms/ particles in alloys disrupted
• harder for layers to slide
• therefore alloys typically stronger, less malleable and less ductile
• alloys can be more corrosion resistant
• pure aluminium used (e.g. cooking foil) where conducting heat required but strength not
needed
• aluminium alloy used (e.g aircraft, ladders) where low density required but greater
strength of alloy needed
• pure copper used (e.g. electrical wires) where conducting electricity needed but high
strength not required
• copper alloys used (e.g.plug pins) where conducting electricity needed but needs to be
harder than pure copper
• metals have fixed melting point but alloys a range as a mixture

(Total for Question 10 = 12 marks)

TOTAL FOR PAPER = 100 MARKS

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The periodic table of the elements

Key

relative atomic mass **atomic symbol** name atomic (proton) number

1	2											3	4	5	6	7	0
							1 **H** hydrogen 1										4 **He** helium 2
7 **Li** lithium 3	9 **Be** beryllium 4											11 **B** boron 5	12 **C** carbon 6	14 **N** nitrogen 7	16 **O** oxygen 8	19 **F** fluorine 9	20 **Ne** neon 10
23 **Na** sodium 11	24 **Mg** magnesium 12											27 **Al** aluminium 13	28 **Si** silicon 14	31 **P** phosphorus 15	32 **S** sulfur 16	35.5 **Cl** chlorine 17	40 **Ar** argon 18
39 **K** potassium 19	40 **Ca** calcium 20	45 **Sc** scandium 21	48 **Ti** titanium 22	51 **V** vanadium 23	52 **Cr** chromium 24	55 **Mn** manganese 25	56 **Fe** iron 26	59 **Co** cobalt 27	59 **Ni** nickel 28	63.5 **Cu** copper 29	65 **Zn** zinc 30	70 **Ga** gallium 31	73 **Ge** germanium 32	75 **As** arsenic 33	79 **Se** selenium 34	80 **Br** bromine 35	84 **Kr** krypton 36
85 **Rb** rubidium 37	88 **Sr** strontium 38	89 **Y** yttrium 39	91 **Zr** zirconium 40	93 **Nb** niobium 41	96 **Mo** molybdenum 42	[98] **Tc** technetium 43	101 **Ru** ruthenium 44	103 **Rh** rhodium 45	106 **Pd** palladium 46	108 **Ag** silver 47	112 **Cd** cadmium 48	115 **In** indium 49	119 **Sn** tin 50	122 **Sb** antimony 51	128 **Te** tellurium 52	127 **I** iodine 53	131 **Xe** xenon 54
133 **Cs** caesium 55	137 **Ba** barium 56	139 **La*** lanthanum 57	178 **Hf** hafnium 72	181 **Ta** tantalum 73	184 **W** tungsten 74	186 **Re** rhenium 75	190 **Os** osmium 76	192 **Ir** iridium 77	195 **Pt** platinum 78	197 **Au** gold 79	201 **Hg** mercury 80	204 **Tl** thallium 81	207 **Pb** lead 82	209 **Bi** bismuth 83	[209] **Po** polonium 84	[210] **At** astatine 85	[222] **Rn** radon 86

* The elements with atomic numbers from 58 to 71 are omitted from this part of the periodic table.

The relative atomic masses of copper and chlorine have not been rounded to the nearest whole number.