Surname	StemUp Model Solution	Centre Number	Candidate Number
First name(s)			2

GCE A LEVEL

A410U10-1

O20-A410U10-1

TUESDAY, 6 OCTOBER 2020 – AFTERNOON

CHEMISTRY – A level component 1
Physical and Inorganic Chemistry

2 hours 30 minutes

	For Examiner's use only		
	Question	**Maximum Mark**	**Mark Awarded**
Section A	**1. to 6.**	**15**	
Section B	**7.**	**11**	
	8.	**13**	
	9.	**20**	
	10.	**11**	
	11.	**24**	
	12.	**13**	
	13.	**13**	
	Total	**120**	

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ADDITIONAL MATERIALS

In addition to this examination paper, you will need a:
- calculator;
- **Data Booklet** supplied by WJEC.

INSTRUCTIONS TO CANDIDATES

Use black ink or black ball-point pen.

Write your name, centre number and candidate number in the spaces at the top of this page.

Section A Answer **all** questions in the spaces provided.
Section B Answer **all** questions in the spaces provided.

Candidates are advised to allocate their time appropriately between **Section A (15 marks)** and **Section B (105 marks)**.

INFORMATION FOR CANDIDATES

The number of marks is given in brackets at the end of each question or part-question.

The maximum mark for this paper is 120.

Your answers must be relevant and must make full use of the information given to be awarded full marks for a question.

The assessment of the quality of extended response (QER) will take place in **Q.10***(c)*(ii) and **Q.11***(c)*.

If you run out of space, use the additional page(s) at the back of the booklet, taking care to number the question(s) correctly.

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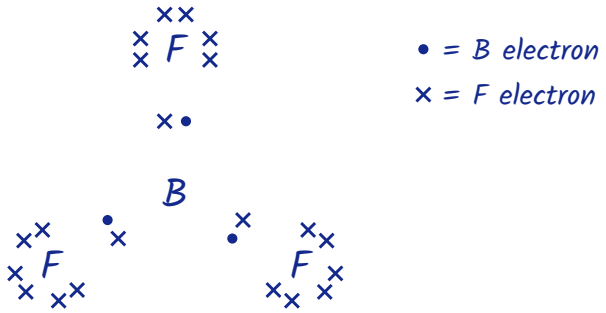
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SECTION A

Answer **all** *questions in the spaces provided.*

1. *(a)* Draw a dot and cross diagram for the molecule BF_3. [1]



(b) Use valence shell electron pair repulsion theory to explain why BF_3 has a trigonal planar structure. [1]

Boron has three bonding pairs and no lone pairs around it. The three bond pairs repel each other equally and arrange themselves as far apart as possible (minimum repulsion), at 120° in one plane.

2. Give the formulae of the **four** sulfur-containing products of the reaction of sodium iodide, NaI, with concentrated sulfuric acid. [2]

$NaHSO_4$, SO_2, S and H_2S

3. A radioactive sample of sodium-24 decays with a half-life of 15 hours to produce magnesium-24.

(a) Identify the type of radioactive decay occurring. [1]

Beta (β^-) decay

(b) A sample of 0.276 g of ^{24}Na decays for 45 hours. Calculate the mass of magnesium-24 produced in this time. [2]

45 ÷ 15 = 3 half-lives

Na-24 left = $0.276 \times (\frac{1}{2})^3 = 0.0345$ g

Mg-24 formed = $0.276 - 0.0345 = 0.2415$ g

Mass = 0.242 g

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4. Addition of excess potassium iodide to 25.0 cm^3 of aqueous copper(II) sulfate produces a brown solution of iodine and a white solid.

(a) Give the **formula** of the white solid. [1]

CuI (copper(I) iodide)

(b) The iodine is titrated using sodium thiosulfate solution of concentration 0.200 mol dm^{-3}.

$$2S_2O_3^{2-} + I_2 \longrightarrow 2I^- + S_4O_6^{2-}$$

The experiment is repeated and the results are shown below.

	Titration 1	Titration 2	Titration 3	Titration 4	Mean titre
Volume of sodium thiosulfate solution / cm^3	24.40	23.90	23.80	23.85	23.85

(i) Suggest why the value from titration 1 is not used in calculating the mean titre. [1]

It is a rough / anomalous titre: it is too far from the other values (not concordant, not within 0.10 cm^3 of them).

(ii) Calculate the number of moles of iodine produced from each 25.0 cm^3 of aqueous copper(II) sulfate solution. [1]

$n(S_2O_3^{2-}) = 0.02385 \times 0.200 = 4.77 \times 10^{-3}$ mol

$n(I_2) = 4.77 \times 10^{-3} \div 2 = 2.385 \times 10^{-3}$ mol

Moles of iodine = 2.39 × 10^{-3} mol

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5. The boiling temperatures of three compounds are shown below.

Compound	ammonia, NH_3	phosphine, PH_3	arsine, AsH_3
Boiling temperature / °C	–33	–88	–63

(a) Explain why the boiling temperature of ammonia is much higher than that of phosphine and arsine. [2]

Ammonia has hydrogen bonding between molecules, but phosphine and arsine only have van der Waals forces. Hydrogen bonds are stronger than van der Waals forces, so more energy is needed.

(b) Give a reason why the boiling temperature of arsine is higher than that of phosphine. [1]

Arsine molecules have more electrons, so there are stronger van der Waals forces between arsine molecules.

6. Iron metal reacts with chlorine to form iron(III) chloride and it reacts with iodine to form iron(II) iodide. Use the standard electrode potentials below to explain this difference. [2]

	Standard electrode potential, E^θ / V
$Cl_2(aq) + 2e^- \rightleftharpoons 2Cl^-(aq)$	+1.36
$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$	+0.77
$I_2(aq) + 2e^- \rightleftharpoons 2I^-(aq)$	+0.54

$E^\ominus$ for Cl_2/Cl^- (+1.36 V) is more positive than for Fe^{3+}/Fe^{2+} (+0.77 V), so chlorine can oxidise Fe^{2+} to Fe^{3+}, forming $FeCl_3$.

$E^\ominus$ for I_2/I^- (+0.54 V) is less positive than for Fe^{3+}/Fe^{2+}, so iodine cannot oxidise Fe^{2+} further to Fe^{3+}, and FeI_2 forms.

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SECTION B

Answer **all** *questions in the spaces provided.*

7. Carboxylic acids are weak acids. The table below gives the K_a values for some carboxylic acids.

Acid	K_a / mol dm^{-3}
ethanoic acid, CH_3COOH	1.74×10^{-5}
chloroethanoic acid, $ClCH_2COOH$	1.43×10^{-3}
dichloroethanoic acid, $Cl_2CHCOOH$	5.13×10^{-2}
trichloroethanoic acid, Cl_3CCOOH	2.24×10^{-1}

(a) Describe how the replacement of hydrogen atoms by chlorine atoms affects the strength of the acid. [1]

As the number of chlorine atoms increases the acid becomes stronger (K_a increases).

(b) Find the pH of a solution of chloroethanoic acid of concentration 0.200 mol dm^{-3}. [2]

$[H^+]^2 = K_a \times [HA] = 1.43 \times 10^{-3} \times 0.200 = 2.86 \times 10^{-4}$

$[H^+] = 0.0169$ mol dm^{-3}

$pH = -\log[H^+] = -\log(0.0169) = 1.77$

pH = 1.8

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(c) The soluble salt sodium ethanoate can be produced from ethanoic acid. A student suggests the following method:

- Measure 25 cm^3 of a solution of ethanoic acid of concentration 0.200 mol dm^{-3} in a measuring cylinder.
- Transfer to a conical flask and add a few drops of methyl orange indicator. This is selected because it changes colour in a pH range of 3.1-4.4.
- Add sodium hydroxide solution of concentration 2.00 mol dm^{-3} from a burette until the indicator changes colour.
- Record the volume of sodium hydroxide required.
- Measure another 25 cm^3 of the ethanoic acid solution.
- Add the same volume of sodium hydroxide as in the first experiment, this time without an indicator.
- Transfer the solution to an evaporating basin and evaporate most of the water.
- Leave the hot solution to cool and allow crystals to form. Filter these away from the solution.

(i) The teacher tells the student that their proposed method would form 0.68 g of sodium ethanoate trihydrate ($CH_3COONa.3H_2O$) and that the solubility of this compound is 2.42 mol dm^{-3} at 20 °C.

Find the minimum volume of water needed to dissolve the mass of sodium ethanoate trihydrate formed at 20 °C and hence suggest why the final step in the method may not be successful. [3]

$M_r(CH_3COONa.3H_2O) = 136.0$

moles = $0.68 \div 136.0 = 5.00 \times 10^{-3}$ mol

volume = $5.00 \times 10^{-3} \div 2.42 = 2.07 \times 10^{-3}$ dm^3

Miniumum volume = 2.07 cm^3

Only about 2 cm^3 of water is needed, so the volume of solution left is too small to filter easily (little or no crystals would be obtained).

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(ii) Suggest **three** changes that should be made to this method, giving reasons for each. [5]

Use a pipette (graduated / volumetric) instead of a measuring cylinder, as it is more precise.

Use phenolphthalein, not methyl orange: ethanoic acid is a weak acid so the equivalence point is above pH 7 and the pH changes only gradually in the range 3.1 to 4.4.

Use NaOH of similar concentration to the acid (e.g. 0.200 mol dm^{-3}): with 2.00 mol dm^{-3} the titre is very small (2.5 cm^3), giving a large percentage error.

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8. Magnesium and sodium lie adjacent to each other in the Periodic Table.

(a) The melting temperature of sodium is 98 °C and magnesium has a melting temperature of 639 °C.

Explain why magnesium has a much higher melting temperature than sodium. [2]

To melt a metal the electrostatic forces between the positive metal ions and the sea of delocalised electrons must be overcome.

Magnesium ions have a greater positive charge (Mg^{2+} vs Na^{+}) and each atom gives two delocalised electrons, so the forces are stronger and more energy is needed.

(b) Sodium and magnesium both conduct electricity. State which of these two metals you would expect to be the better conductor of electricity. Give a reason for your answer. [1]

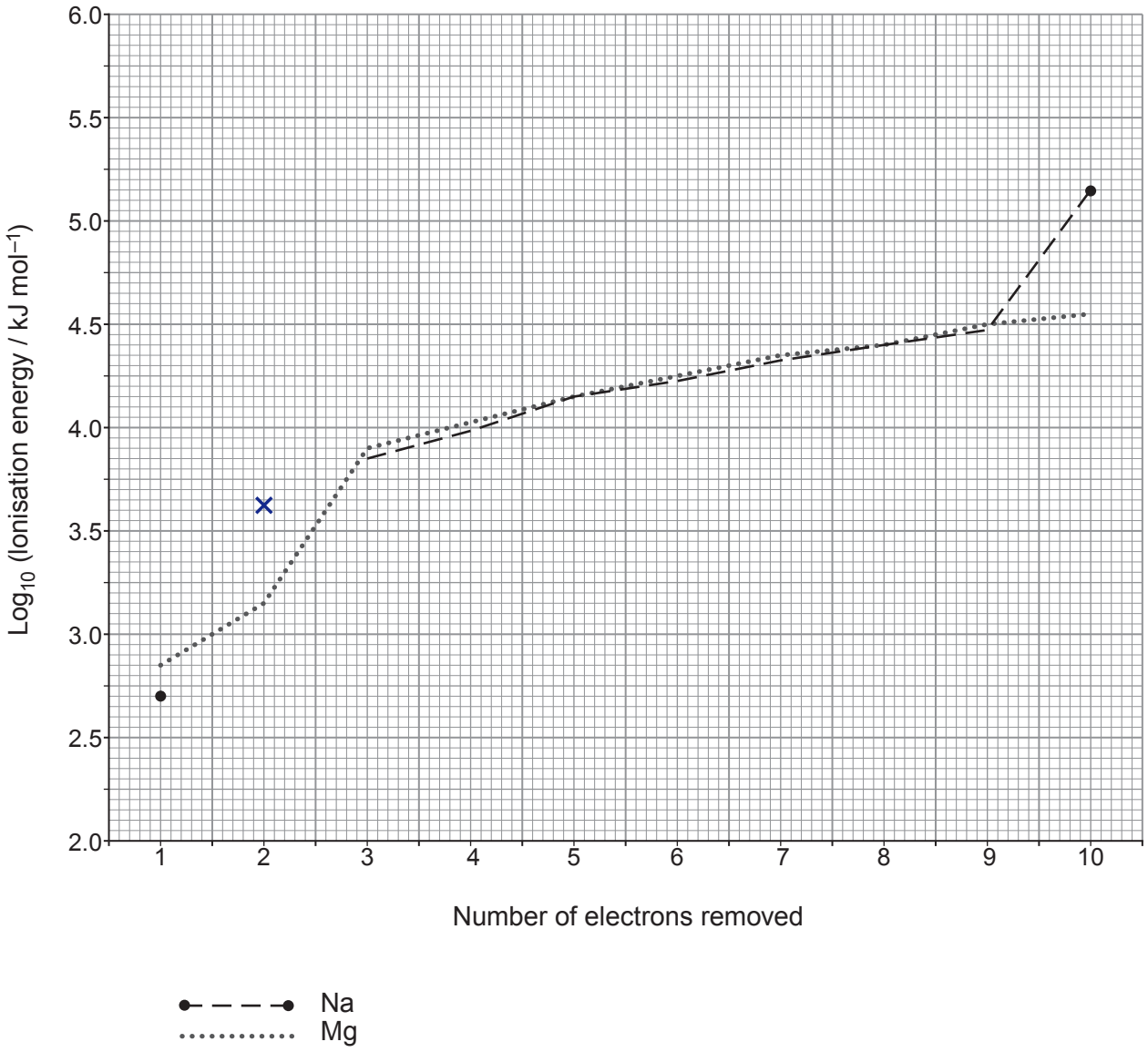
Magnesium, as it has more delocalised electrons per atom (two outer shell electrons compared with one) to move and carry charge.

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(c) The graph below shows the first ten log molar ionisation energies of sodium and magnesium.



(i) Give the equation for the **third** ionisation energy of magnesium. [1]

$Mg^{2+}(g) \rightarrow Mg^{3+}(g) + e^-$

(ii) Explain why the tenth ionisation energy of sodium is much larger than the tenth ionisation energy of magnesium. [2]

Sodium's tenth electron is removed from the innermost (first) shell, nearer the nucleus; magnesium's is removed from the second shell. No shielding in sodium; in magnesium the inner shell gives some shielding.

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(iii) The plot for the log of the second ionisation energy of sodium is missing from the graph. Estimate its value and hence calculate an approximate second ionisation energy. [2]

$\log_{10}(IE) \approx 3.65$ (estimated from the trend, plotted at 2 on the graph)

$IE = 10^{3.65} = 4467$ kJ mol^{-1}

Second ionisation energy = ... 4470 ..... kJ mol^{-1}

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(d) (i) Heating magnesium carbonate causes the compound to decompose forming magnesium oxide and carbon dioxide. The equation for this process is:

$$MgCO_3(s) \longrightarrow MgO(s) + CO_2(g)$$

	Standard enthalpy change of formation, $\Delta_f H^\theta$ / kJ mol^{-1}	Standard entropy, S^θ / J mol^{-1} K^{-1}
$MgCO_3(s)$	–1113	66
MgO(s)	–602	27
$CO_2(g)$	–394	214

Calculate the minimum temperature required, in °C, for the thermal decomposition to occur. [4]

$\Delta H = (-602 + -394) - (-1113) = +117$ kJ mol^{-1}

$\Delta S = (27 + 214) - 66 = +175$ J K^{-1} mol^{-1} $= 0.175$ kJ K^{-1} mol^{-1}

$\Delta G = \Delta H - T\Delta S \leq 0$, so $T = \Delta H \div \Delta S = 117 \div 0.175 = 669$ K

$669 - 273 = 396$ °C

Minimum temperature = 396 °C

(ii) The element calcium lies below magnesium in Group 2. Suggest a value for the minimum temperature required for the thermal decomposition of $CaCO_3$. Give a reason for your answer. [1]

About 900 °C (above 400 °C), because the carbonates become more thermally stable down Group 2 (Ca^{2+} is larger, so it polarises the carbonate ion less).

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9. Organometallic compounds contain metal atoms bonded directly to the carbon atom of an organic group. One example is methyllithium (CH_3Li).

(a) The electronegativity values of the elements present in methyllithium are given below.

Element	H	Li	C
Electronegativity	2.20	0.98	2.55

(i) State what is meant by the term *electronegativity*. [1]

The ability of an atom to attract a (bonding) pair of electrons in a covalent bond.

(ii) State which type of bond will form between lithium and carbon. Give a reason for your answer. [1]

(Polar) covalent: the electronegativity difference (2.55 – 0.98 = 1.57) is too small for ionic, between 0.5 and 2.0.

(b) Methyllithium is synthesised by treating bromomethane with a suspension of metallic lithium in a non-aqueous solvent. This forms a solution of methyllithium in the non-aqueous solvent.

$$CH_3Br + 2Li \longrightarrow CH_3Li + LiBr$$

(i) Calculate the atom economy for this method of producing methyllithium. [2]

Atom economy = $M_r(CH_3Li) \div M_r$(reactants) × 100

= 21.97 ÷ (94.93 + 2 × 6.94) × 100 = 21.97 ÷ 108.81 × 100

Atom economy = *20.2* %

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(ii) An alternative method of synthesising methyllithium is by using chloromethane as a source of the methyl group. Use the information given below to suggest which method is better for the production of methyllithium.

You should include **two** advantages and **one** disadvantage for your chosen method. [4]

	CH_3Cl	CH_3Br	CH_3Li	LiCl	LiBr
M_r	50.5	94.9	21.9	42.4	86.8
Melting temperature / °C	–97	–94			
Boiling temperature / °C	–24	+4			
Solubility in solvent used	very soluble	very soluble	soluble	insoluble	partly soluble
Relative cost per gram	1.5	1.0			

Atom economy for production of methyllithium from chloromethane = 34.0 %

Chloromethane is the better method.

Advantage 1: it has a higher atom economy (34.0 % compared with 20.2 %), so less waste.

Advantage 2: LiCl is insoluble in the solvent, so it is easily removed by filtration, whereas LiBr is partly soluble.

Disadvantage: CH_3Cl boils at −24 °C, so it is a gas at room temperature and harder to handle than CH_3Br, a liquid just below 4 °C.

(c) Methyllithium is a very strong base.

State what is meant by the term *strong base*. [1]

A substance that removes (accepts) H^+ ions in a non-reversible reaction (it is completely ionised / reacts completely).

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(d) When methyllithium is added to water the reaction below occurs.

$$CH_3Li(sol) + H_2O(l) \longrightarrow LiOH(aq) + CH_4(g)$$

When 10.00 cm^3 of a non-aqueous solution of methyllithium is added to 250.0 cm^3 of water a total volume of 391.8 cm^3 of methane is produced at a temperature of 298 K and a pressure of 1 atm. The pH of the aqueous layer formed is 12.8.

You may assume that LiOH is totally insoluble in the non-aqueous solvent.

(i) Use the volume of gas produced to calculate the concentration of the initial methyllithium solution. Give your answer to an **appropriate** number of significant figures.

You **must** show your working. [3]

$n(CH_4) = 391.8 \div 24\,500 = 0.01599$ mol

$n(CH_3Li) = n(CH_4) = 0.01599$ mol (1 : 1)

concentration $= 0.01599 \div 0.01000\ dm^3 = 1.599$ mol dm^{-3}

Concentration = ... 1.60 ..... mol dm^{-3}

(ii) Use the pH to calculate the concentration of the initial methyllithium solution.

You **must** show your working. [3]

$[H^+] = 10^{-12.8} = 1.585 \times 10^{-13}$ mol dm^{-3}

$[OH^-] = K_w \div [H^+] = 1.0 \times 10^{-14} \div 1.585 \times 10^{-13} = 0.0631$

$n(LiOH) = 0.0631 \times 0.250 = 0.01577$ mol $= n(CH_3Li)$

concentration $= 0.01577 \div 0.01000 = 1.58$ mol dm^{-3}

Concentration = ... 1.58 ..... mol dm^{-3}

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(iii) State which of these two methods gives the more accurate value. Give a reason for your answer. [2]

The gas volume method is more accurate, because the volume is measured more precisely, to more significant figures (391.8 cm^3), than the pH (12.8, only 1 d.p.).

(iv) The reaction between methyllithium and water is exothermic. The enthalpy change of the reaction is –198 kJ mol^{-1}.

Calculate the expected temperature rise when 0.010 mol of pure methyllithium is added to 250.0 cm^3 of water. [3]

heat released $q = 198\,000 \times 0.010 = 1980$ J

$\Delta T = q \div (m \times c) = 1980 \div (250.0 \times 4.18)$

$\Delta T = 1.89$ °C

Temperature rise = 1.9 °C

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10. Copper and cobalt are both considered to be transition elements whilst zinc is not.

(a) Give a reason why zinc is not considered to be a transition element. [1]

Zinc only forms Zn^{2+} ($3d^{10}$), so it does not form an ion with a partially filled d-subshell.

(b) Transition elements are able to form complex ions and these ions are usually coloured.

Give the colour of the $[Co(H_2O)_6]^{2+}$ ion. [1]

pink

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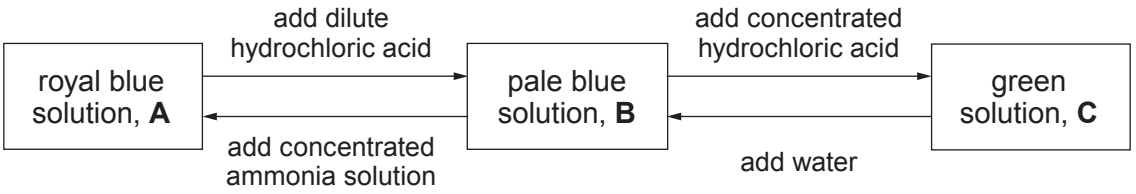
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(c) When an aqueous solution containing Cu^{2+} ions is treated with a concentrated aqueous solution of ammonia a pale blue precipitate forms. This dissolves when more ammonia is added to form a royal blue solution, labelled **A** in the scheme below.

(i) Identify the pale blue precipitate. [1]

copper(II) hydroxide, $Cu(OH)_2$

(ii) Some of the colour changes observed when other solutions are added to solution **A** are given in the following scheme.



Identify the species causing these colours and use Le Chatelier's principle to explain the changes that occur. Your answer should include appropriate equations. [6 QER]

A is $[Cu(H_2O)_2(NH_3)_4]^{2+}$ (royal blue), B is $[Cu(H_2O)_6]^{2+}$ (pale blue) and C is $[CuCl_4]^{2-}$ (green).

A ⇌ B: $[Cu(H_2O)_2(NH_3)_4]^{2+} + 4H_2O \rightleftharpoons [Cu(H_2O)_6]^{2+} + 4NH_3$

Dilute HCl is mostly water, which increases $[H_2O]$, and its H^+ ions protonate the ammonia ($NH_3 + H^+ \rightarrow NH_4^+$), lowering $[NH_3]$. By Le Chatelier's principle the equilibrium shifts to the right, forming pale blue B.

Adding concentrated ammonia increases $[NH_3]$, so the equilibrium shifts back to the left, re-forming royal blue A.

B ⇌ C: $[Cu(H_2O)_6]^{2+} + 4Cl^- \rightleftharpoons [CuCl_4]^{2-} + 6H_2O$

Concentrated HCl provides chloride ions, at a much higher concentration than the water, so the equilibrium shifts to the right, forming green C (ligand substitution, 6 to 4 coordinate).

Adding water decreases $[Cl^-]$ and increases $[H_2O]$, so the equilibrium shifts to the left, re-forming pale blue B.

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(d) Give another typical property of transition elements and use it to show that copper is a typical transition element but zinc is not. [2]

Variable oxidation states: copper forms compounds in the +1 and +2 states (e.g. Cu_2O and CuO), whereas zinc only forms the +2 state.

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11. Fuel cells are an electrochemical method of obtaining energy from fuels.

(a) Give **one** advantage and **one** disadvantage of the use of fuel cells. [2]

Advantage *higher efficiency; can use renewable fuels (carbon-neutral), and the only product is water.*

Disadvantage *H_2 is explosive, so it is difficult to store and transport (and needs expensive catalysts).*

(b) Hydrogen gas is a common fuel used in fuel cells. Hydrogen can be produced in the electrolysis of sodium chloride solution. The reaction occurring is shown below.

$$2NaCl(aq) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g) + Cl_2(g)$$

(i) Electrolysis was undertaken on 250 cm^3 of a solution of sodium chloride of concentration 5.00 mol dm^{-3} until 2856 cm^3 of hydrogen gas was produced at a temperature of 320 K and a pressure of 1.14×10^5 Pa.

Calculate the concentration of the sodium hydroxide present at the end of the process. [4]

$n = pV \div RT$

$n(H_2) = (1.14 \times 10^5 \times 2856 \times 10^{-6}) \div (8.31 \times 320) = 0.1224$ mol

$n(NaOH) = 2 \times n(H_2) = 2 \times 0.1224 = 0.2449$ mol

$[NaOH] = 0.2449 \div 0.250 = 0.979$ mol dm^{-3}

Concentration = *0.979* mol dm^{-3}

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(ii) The chlorine gas produced in the reaction was stored under a pressure of 5.05×10^5 Pa at a temperature of 283 K.

Calculate the volume of the chlorine gas under these conditions. [2]

$n(Cl_2) = n(H_2) = 0.1224$ mol

$V = nRT \div p = (0.1224 \times 8.31 \times 283) \div 5.05 \times 10^5$

$V = 5.70 \times 10^{-4}$ m^3 = 570 cm^3

Volume = 570 cm^3

(iii) At the end of the electrolysis process the solution contained a base and a salt. A student suggests that this could therefore be used as a buffer. Is the student correct? Explain your answer. [3]

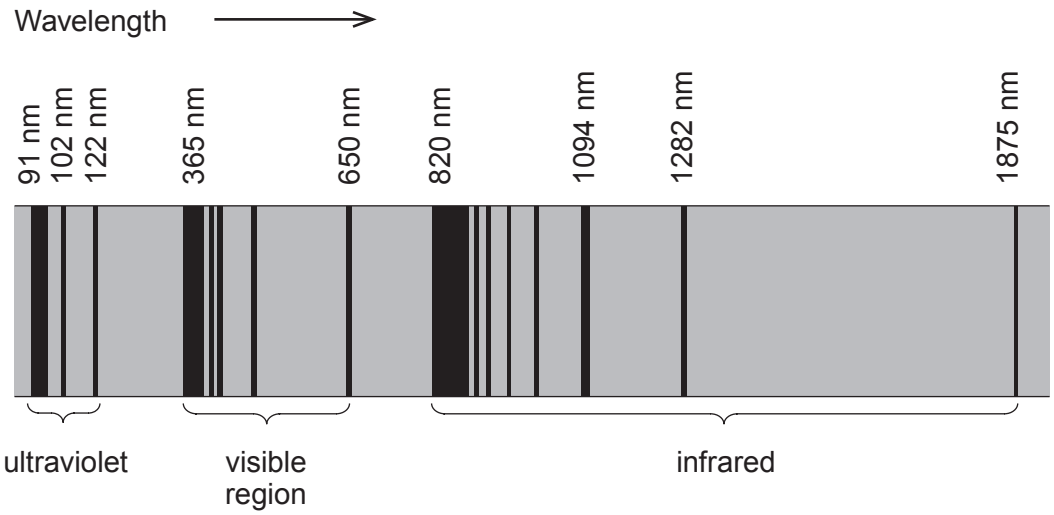
No, the student is incorrect. NaOH is a strong base; a buffer must use a weak acid or a weak base with its salt.

A buffer needs a reversible reaction involving gain / loss of H^+ ions to resist pH change.

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(c) When an electrical discharge is passed through hydrogen gas, the atoms emit light and other frequencies of electromagnetic radiation. These form the emission spectrum shown below.



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Explain why this pattern of sharp lines forms and how it can be used to find the ionisation energy of hydrogen. You should also calculate the ionisation energy of hydrogen in kJ mol^{-1}. [6 QER]

$E = hc \div \lambda = (6.63 \times 10^{-34} \times 3.00 \times 10^{8}) \div (91 \times 10^{-9})$

$E = 2.19 \times 10^{-18}$ J per atom

$\times N_A$: $2.186 \times 10^{-18} \times 6.02 \times 10^{23} = 1.316 \times 10^{6}$ J mol^{-1}

$= 1316$ kJ mol^{-1}

Ionisation energy = 1316 kJ mol^{-1}

The energy levels of the electron in a hydrogen atom are quantised: it can only have certain fixed energies.

The discharge excites electrons to higher levels. When an electron drops from a higher to a lower level, it emits a photon with energy equal to the difference between the levels, $\Delta E = hf = hc/\lambda$.

Only fixed energy differences exist, so only certain frequencies are emitted, giving sharp lines.

Each series corresponds to electrons dropping to one particular level: Lyman (ultraviolet) to $n = 1$, Balmer (visible) to $n = 2$, Paschen (infrared) to $n = 3$.

Energy levels get closer together at higher n, so the lines in each series converge.

The convergence limit of the Lyman series (electron falling from $n = \infty$ to $n = 1$) equals the ionisation energy. This is the 91 nm line: $E = hc/\lambda$ gives the energy per atom, and × Avogadro constant gives 1316 kJ mol^{-1}.

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(d) Another fuel that can be used in fuel cells is methanol. The equation for the combustion of gaseous methanol at a temperature of 120°C is given below.

$$2CH_3OH(g) + 3O_2(g) \longrightarrow 2CO_2(g) + 4H_2O(g) \qquad \Delta H = -1286\,kJ\,mol^{-1}$$

(i) Calculate the bond energy of the C—O bond in methanol. [3]

Bond	Bond energy / kJ mol^{-1}
C — H	414
O = O	498
O — H	464
C = O	803

Bonds broken: 6(C–H) + 2(C–O) + 2(O–H) + 3(O=O)

= 6(414) + 2(C–O) + 2(464) + 3(498) = 4906 + 2(C–O)

Bonds made: 4(C=O) + 8(O–H) = 4(803) + 8(464) = 6924

ΔH = broken – made: −1286 = 4906 + 2(C–O) – 6924

2(C–O) = 732, so C–O = 366 kJ mol^{-1}

Bond energy = ... 366 ..... kJ mol^{-1}

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(ii) Give a reason why the value calculated in part (i) may not precisely match the actual energy of this C—O bond. [1]

Bond energies given are average values from many compounds, not the same in every molecule.

(iii) The standard enthalpy change of combustion of methanol has a value of $-715\,kJ\,mol^{-1}$. Give reasons why this value is very different from the value associated with the equation. [3]

The standard value is for combustion of 1 mol of methanol; the equation is for 2 mol. Methanol and water are not in their standard states (liquids), they are gases here. Not standard conditions: 120 °C, not 298 K (25 °C).

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12. The oxidation of carbon monoxide to carbon dioxide in the presence of suitable catalysts is an important method of removing this toxic gas from gas mixtures produced during incomplete combustion.

$$2CO + O_2 \longrightarrow 2CO_2$$

Some catalysts for this reaction are produced by soaking aluminium oxide pellets in a solution of a transition metal chloride followed by drying. Two suitable transition metal chlorides are palladium chloride and ruthenium chloride.

(a) These catalysts are examples of heterogeneous catalysts.

(i) State what is meant by *heterogeneous* in this context. [1]

Catalyst is in a different (physical) state from the reactants.

(ii) Give another example of a heterogeneous catalyst, clearly identifying the reaction that it catalyses. [1]

Iron (Fe) in the Haber process: $N_2 + 3H_2 \rightleftharpoons 2NH_3$

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(b) Some information regarding these catalysed reactions is given below.

	Activation energy / kJ mol^{-1}	Frequency factor, A / mol dm^{-3} s^{-1}
palladium catalyst	61.7	6.1×10^9
ruthenium catalyst	79.4	14.1×10^9

(i) At a temperature of 600 K, the value of the rate constant for the reaction catalysed by palladium is 2.58×10^4.

I. Give the unit for this rate constant. [1]

mol dm^{-3} s^{-1}

II. Find the value of the rate constant for the ruthenium catalyst under the same conditions and hence identify which of these two catalysts is the more effective. [4]

$k = A\, e^{-Ea/RT}$ (E_a = 79.4 kJ mol^{-1} = 79 400 J mol^{-1})

$Ea/RT = 79\,400 \div (8.31 \times 600) = 15.92$

$k = 14.1 \times 10^9 \times e^{-15.92} = 14.1 \times 10^9 \times 1.213 \times 10^{-7}$

$k = 1.71 \times 10^3$ mol dm^{-3} s^{-1}

Palladium is the more effective catalyst, as its rate constant at 600 K (2.58×10^4) is higher (lower E_a).

(ii) Under certain conditions the oxidation of carbon monoxide can occur without a catalyst. The rate equation for this process is

$$\text{rate} = k[CO][O_2]$$

Suggest a two-step mechanism for the uncatalysed oxidation of carbon monoxide. Label the rate determining step clearly. [3]

Step 1: $CO + O_2 \rightarrow CO_2 + O$ (slow, rate determining step)

Step 2: $CO + O \rightarrow CO_2$ (fast)

Step 1 is the rate determining step as it involves one CO and one O_2, matching the rate equation.

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(c) Carbon monoxide is classed as a reducing agent.

(i) State what is meant by a *reducing agent*. [1]

A substance that loses (donates) electrons, so it is oxidised, reducing another species.

(ii) Explain why carbon monoxide is a reducing agent whilst the corresponding oxide of lead, PbO, is not. [2]

Carbon is more stable in the +4 oxidation state, so CO (C is +2) is easily oxidised to CO_2; lead is more stable in the +2 state, so PbO is not oxidised to +4.

This is due to the inert pair effect, which increases down Group 4.

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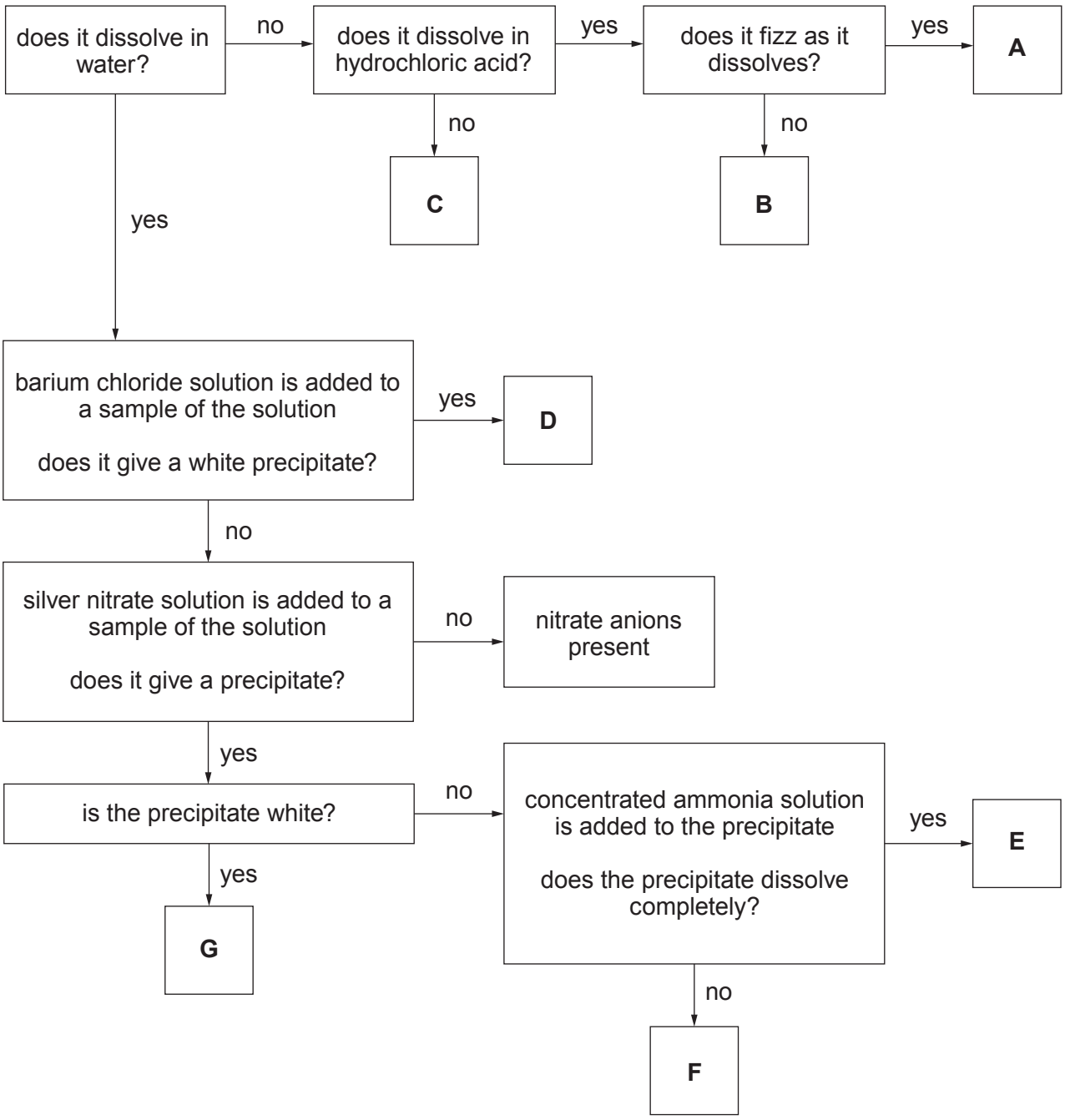
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FOR QUESTION 13

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13. A student is given several magnesium compounds containing common anions and designs the scheme below to identify the anion present in each compound.



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(a) The student labels seven possible results as **A**-**G**. He is able to suggest likely identities for anions that would give six of these results, and his teacher said that none of the common anions that he had studied would give the seventh result.

Suggest possible anions for six of the seven results below. Write '**UNKNOWN**' for the result that would not correspond to a common anion. [4]

A carbonate, CO_3^{2-}

B oxide, O^{2-} / hydroxide, OH^-

C UNKNOWN

D sulfate, SO_4^{2-}

E bromide, Br^-

F iodide, I^-

G chloride, Cl^-

(b) The planned method would not be suitable for identifying **all** these anions in **sodium** compounds. Explain why the method would not be suitable and suggest how it could be modified to make it suitable. [3]

All sodium salts are soluble, so Na_2CO_3 would give a white precipitate with barium chloride (like sulfate) and Na_2CO_3 / NaOH would give a precipitate with silver nitrate.

Modify: add dilute nitric acid to all samples first (fizzing shows carbonate), then add $BaCl_2$ or $AgNO_3$.

(c) Concentrated aqueous ammonia solution at 298 K contains 31.0 % ammonia by mass. The density of this solution is 0.900 g cm^{-3}.

Calculate the concentration of this solution in mol dm^{-3}. [3]

1 dm^3 (1000 cm^3) of solution has a mass of 1000 × 0.900 = 900 g

mass of NH_3 = 0.310 × 900 = 279 g

concentration = 279 ÷ 17.03 = 16.4 mol in 1 dm^3

Concentration = 16.4 mol dm^{-3}

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(d) The student is provided with a solution containing chloride as the only anion. He is told that the solution also contains sodium ions, magnesium ions, or a mixture of both sodium and magnesium ions.

He proposes using a flame test to identify the cation(s) present in the mixture.

State the observations expected for each possibility and explain whether this method would be appropriate. [3]

Sodium ions only: yellow / orange flame. Magnesium ions only: no flame colour.

Mixture of both: yellow / orange flame (the sodium colour).

Not appropriate: sodium only and the mixture give the same yellow flame, so they cannot be distinguished.

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