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

1 Determination of the dependence of the rate of reaction between glucose and acidified potassium manganate(VII) on temperature

Read through the whole method before starting any practical work. Where appropriate, prepare a table for your results in the space provided.

Show your working and appropriate significant figures in the final answer to **each** step of your calculations.

Glucose, $C_6H_{12}O_6$, is a sugar that can act as a reducing agent. You will investigate how an increase in temperature affects the rate of the redox reaction between glucose and acidified potassium manganate(VII).

FA 1 is 0.020 mol dm^{-3} acidified potassium manganate(VII), $KMnO_4$.
FA 2 is 1.0 mol dm^{-3} sulfuric acid, H_2SO_4.
FA 3 is an aqueous solution containing 32.8 g dm^{-3} glucose, $C_6H_{12}O_6$.
FA 4 is a solid mixture containing an unknown percentage of glucose.

You will measure the time it takes for the purple colour to disappear. Your table of results on the following page should include the rate of reaction for each experiment.

(a) Method

Experiment 1

1. Fill the burette with **FA 1**.
2. Add 5.00 cm^3 of **FA 1** into the 250 cm^3 beaker.
3. Use the 50 cm^3 measuring cylinder to transfer 50.0 cm^3 of **FA 2** into the beaker containing **FA 1**.
4. Use the same measuring cylinder to transfer 50.0 cm^3 of distilled water into the same beaker.
5. Place the beaker on the tripod and heat its contents to between 75 °C and 80 °C.
6. While the solution in the beaker is being heated, pour 25.0 cm^3 of **FA 3** into the 25 cm^3 measuring cylinder.
7. When the temperature of the contents of the beaker has reached between 75 °C and 80 °C, turn off the Bunsen burner and **carefully** hold the top of the hot beaker with either a cloth or paper towel and place it onto the white tile.
8. Record the temperature of the solution in the beaker.
9. Add the 25.0 cm^3 of **FA 3** and **immediately** start timing.
10. Stir the contents of the beaker once and stop timing as soon as the solution turns colourless. Record the time to the nearest second.
11. Record the temperature of the solution as soon as it is colourless.
12. Calculate and record the average temperature of the reaction mixture to one decimal place.
13. Empty, rinse and dry the beaker so it is ready for use in **Experiment 2**.

Experiment 2

1. Add 5.00 cm^3 of **FA 1** into the 250 cm^3 beaker.
2. Use the 50 cm^3 measuring cylinder to transfer 50.0 cm^3 of **FA 2** into the beaker containing **FA 1**.
3. Use the same measuring cylinder to transfer 50.0 cm^3 of distilled water into the same beaker.
4. Place the beaker on the tripod and heat its contents to between 35 °C and 40 °C.
5. While the solution in the beaker is being heated, pour 25.0 cm^3 of **FA 3** into the 25 cm^3 measuring cylinder.
6. When the temperature of the contents of the beaker has reached between 35 °C and 40 °C, turn off the Bunsen burner and **carefully** hold the top of the hot beaker with either a cloth or paper towel and place it onto the white tile.
7. Record the temperature of the solution in the beaker.
8. Add the 25.0 cm^3 of **FA 3** and **immediately** start timing.
9. Stir the contents of the beaker once and stop timing as soon as the solution turns colourless. Record the time to the nearest second.
10. Record the temperature of the solution as soon as it is colourless.
11. Calculate and record the average temperature of the reaction mixture to one decimal place.
12. Empty, rinse and dry the beaker so it is ready for use in **Experiment 3**.

Experiments 3, 4 and 5

1. Repeat the method for **Experiment 2** at three different temperatures.
2. Keep the temperature of the contents of the beaker between room temperature and 80 °C.
3. Record all your results in your table.

Results:

The rate of reaction can be calculated as shown.

$$\text{rate} = \frac{1000}{\text{reaction time}}$$

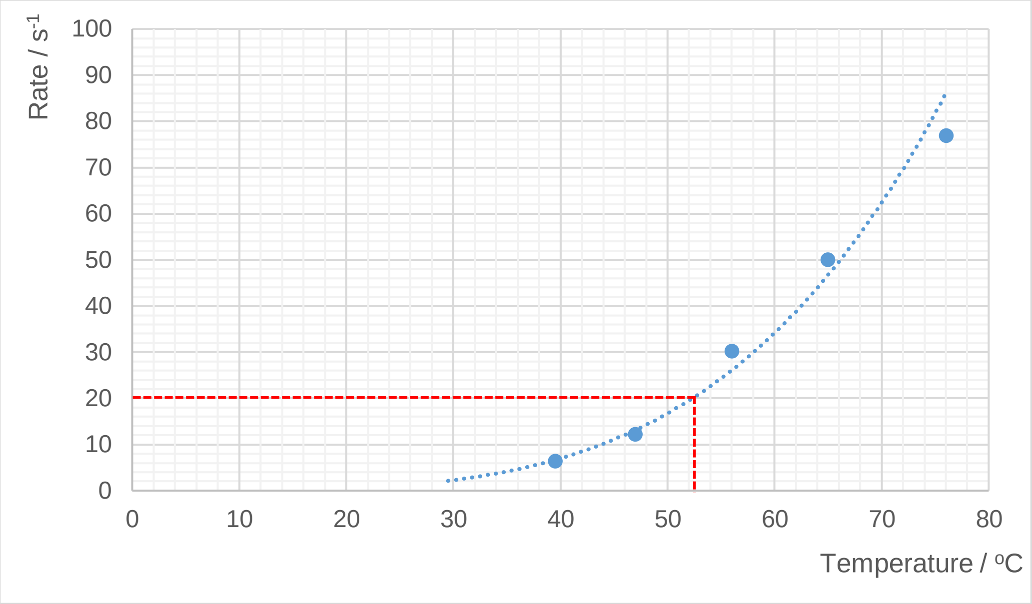
Calculate the rate of reaction for each experiment and include this in your table.

Experiment	T_{initial} / °C	T_{final} / °C	T_{average} / °C	time / s	rate / s^{-1}
1	80.0	72.0	76.0	12	76.9
2	40.0	39.0	39.5	155	6.45
3	49.0	45.0	47.0	82	12.2
4	59.0	53.0	56.0	33	30.3
5	69.0	61.0	65.0	20	50.0

[5]

(b) Plot a graph of rate (*y*-axis) against average temperature (*x*-axis) on the grid below. Select a scale on the *x*-axis to include an average temperature of 30.0 °C. Label any points you consider anomalous.

Draw the most appropriate best-fit curve taking into account all of your plotted points and extrapolate it to 30.0 °C.



[3]

(c) **Use your graph** to calculate the **time** to the nearest second that the reaction would have taken if the average temperature had been 52.5°C.
Show **on the grid** how you obtained your answer.

From the graph, at 52.5 °C,

$$\textbf{rate} = \frac{\textbf{1000}}{\textbf{reaction time}} = \textbf{20 s}^{-1}$$

$$\textbf{Hence, reaction time} = \frac{\textbf{1000}}{\textbf{20}} = \textbf{50 s}$$

time = **50 s** [2]

(d) Explain, by referring to your graph or your table of results, how the rate of reaction is affected by an increase in temperature.

The rate of reaction increases with increase in temperature. The gradient increases with temperature / graph is exponential / acceleration of rate with temperature increase.

[1]

(e) **(i)** Calculate the maximum percentage error in the reaction time recorded for **Experiment 1**. Assume the error of the timer is ±1 s.

Maximum percentage error = 1 / 12 x 100
= 8.33%

maximum percentage error in **Experiment 1** = **8.33** % [1]

(ii) You have carried out experiments at five different temperatures.

Identify an experiment, if any, you should have repeated. Give a reason for your answer.

Identifies no anomaly as all points are near line of best fit.

[1]

(iii) Suggest one way to improve the accuracy of the results for this investigation.

Use thermostatically controlled water bath to heat both reagents and keep them at constant temperature.

[1]

(f) **(i)** Calculate the concentration of glucose in **FA 3** in mol dm^{-3}.

$$\textbf{Concentration of glucose} = \frac{32.8}{180.0} = 0.182 \text{ mol dm}^{-3}$$

concentration of glucose in **FA 3** = **0.182** mol dm^{-3} [1]

(ii) Calculate the volume of 0.020 mol dm^{-3} acidified $KMnO_4$ that would react with **all** the glucose present in 25.0 cm^3 of **FA 3**.

$$5C_6H_{12}O_6 + 24MnO_4^- + 72H^+ \rightarrow 30CO_2 + 24Mn^{2+} + 66H_2O$$

Amount of glucose in 25.0 cm^3 = 0.182 x 25.0/1000
= 4.55 x 10^{-3} mol

Amount of $KMnO_4$ reduced = 4.55 x 10^{-3} x $\frac{24}{5}$
= 2.18 x 10^{-2} mol

Volume of $KMnO_4$ reduced = 2.18 x 10^{-2}/0.020
= 1.09 dm^3

Volume of acidified $KMnO_4$ required = **1.09 dm^3** [2]

(iii) Comment on the feasibility of conducting a redox titration between the given concentration of glucose and acidified $KMnO_4$.

It is not feasible as the volume required for acidified $KMnO_4$ is too large.

[1]

(g) **Determination of the percentage of glucose in FA 4 by titration**

1. Weigh accurately 1.5 g of **FA 4** in a weighing bottle. Record your weighing appropriately in the space below. If you use **TARE** facility of the balance, please indicate clearly in your recording.
2. Dissolve the solid in a beaker and quantitatively transfer into a 250 cm^3 volumetric flask. Make up to the mark with distilled water. Label this **FA 4 solution**.
3. Pipette 25.0 cm^3 of **FA 4 solution** into a conical flask.
4. Using a measuring cylinder add 50.0 cm^3 of **FA 2** into the conical flask.
5. Place the conical flask on the tripod and heat its contents to between 75 °C and 80 °C.
6. When the temperature of the contents of the conical flask has reached between 75 °C and 80 °C, turn off the Bunsen burner and **carefully** hold the neck of conical flask with either a cloth or paper towel.
7. Titrate with **FA 1** until a permanent colour change is observed.

8. Record the burette readings in the suitable format. Repeat titration to achieve consistent results.

Results:

*Weighing of **FA 4***

Mass of weighing bottle and FA 4 / g	**1.825**
Mass of empty weighing bottle / g	**0.325**
Mass of FA 4 used / g	**1.500**

OR

Mass of weighing bottle only / g	**TARE**
Mass of FA 4 / g	**1.500**

Tabulation with ALL readings recorded, clear headers and units for mass table
3 d.p. correct subtraction and mass of solid within stated range (1.450 – 1.550 g)

Titration

Experiment	**1**	**2**
Final burette reading / cm^3	**22.40**	**42.40**
Initial burette reading / cm^3	**0.00**	**20.00**
Volume of FA 1 used / cm^3	**22.40**	**22.40**

Tabulation with ALL readings recorded, clear headers and units for titration table
2 d.p. for titration reading
At least two consistent readings ±0.10 cm^3

[5]

(h) **(i)** From your titrations, obtain a suitable volume of **FA 1** to be used in your calculations. Show clearly how you obtained this volume.

Average volume of FA 1 used = (22.40 + 22.40) ÷ 2
= 22.40 cm^3

volume of **FA 1** = **22.40 cm^3** [1]

(ii) Calculate the amount of **FA 1** required for titration.

$$5C_6H_{12}O_6 + 24MnO_4^- + 72H^+ \rightarrow 30CO_2 + 24Mn^{2+} + 66H_2O$$

Amount of FA 1 used = 22.40/1000 x 0.020
= 4.48 x 10^{-4} mol

amount of **FA 1** = **4.48 x 10^{-4} mol** [1]

(iii) Hence, calculate the amount of glucose in **FA 4**.

Amount of glucose in 25.0 cm^3 of FA 4 solution = 4.48 x 10^{-4} x $\frac{5}{24}$
= 9.33 x 10^{-5} mol

Amount of glucose in 250 cm^3 of FA 4 solution = 9.33 x 10^{-5} x $\frac{250}{25.0}$
= 9.33 x 10^{-4} mol

amount of glucose in **FA 4** = **9.33 x 10^{-4} mol** [1]

(iv) Use your answer in **(h)(iii)** to determine the percentage of glucose in the mixture.
[A_r: C, 12.0; O, 16.0; H, 1.0]

Mass of glucose in 250 cm^3 of FA 4 solution = 9.33 x 10^{-4} x 180.0
= 0.168 g
Percentage of glucose in FA 4 = 0.168/1.500 x 100 %
= 11.2 %

percentage of glucose in the mixture = **11.2 %** [1]

[Total: 27]

2 Planning

An electroplating process uses an electric current to coat an object with a thin layer of metal. An example is *copper plating*. In this process, copper metal is made the anode and is allowed to undergo oxidation. The copper is then reformed at the cathode, coating the object to the plated. The mass of copper plated at the end of the process can be calculated by taking relevant mass measurements.

In the following, you will use the *copper plating* process to determine the Faraday constant, which is the charge in coulombs, C, carried by 1 mole of electrons.

(a) Show that the mass of copper plated onto an object is related to the duration of the *copper plating* process by the following expression:

$$m_{Cu} = \frac{I \times M_{Cu}}{2F}\, t$$

where

m_{Cu} : mass of copper plated onto an object in grams, g

M_{Cu} : molar mass of copper, g mol^{-1}

I : current in amperes, A

F : Faraday constant

t : duration of the *copper plating* process in seconds, s

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

$Q = I\,t$ ....(1)

$Q = n_e\,F$ ....(2)

From (1) and (2)

$I\,t = n_e\,F$

$I\,t = (2n_{cu})\,F$

$I\,t = \left(\frac{2m_{Cu}}{M_{Cu}}\right) F$

$m_{Cu} = \left(\frac{I \times M_{Cu}}{2F}\right) t$ (shown)

[2]

(b) A series of experiments can be carried out, in which the duration of each experiment is varied and the corresponding mass of copper plated onto the object is determined. The data obtained is then to be used for plotting a suitable graph to determine the value of the Faraday constant.

You are required to write a plan, describing how such a series of experiments can be carried out for an appreciable mass of copper to be plated. The duration of each experiment should be at least 5 min but not exceeding 50 min.

You may assume that you are provided with:

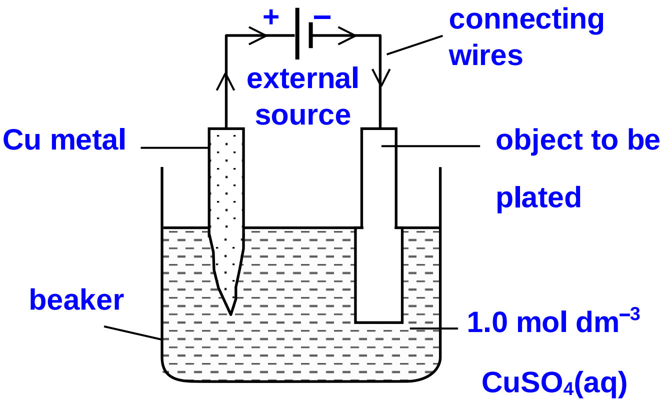
- 1.0 mol dm^{-3} copper(II) sulfate solution;
- five equal pieces of clean, dry copper metals;
- five equal pieces of clean, dry objects to be plated (the object is electrically conductive but is inert during the *copper plating* process);
- an electrical power source that supplies a constant current of 0.50 A (the electrical power source can be switched on and off);

- connecting wires;
- stopwatch;
- the apparatus and chemicals normally found in a school laboratory.

Your plan should include the following:

- a clearly labelled diagram of the *copper plating* setup, including the direction of electron flow;
- the procedure you would follow;
- the measurements you would make and how you would determine the mass of copper plated accurately and reliably to allow a suitable graph to be drawn.

Experimental setup:

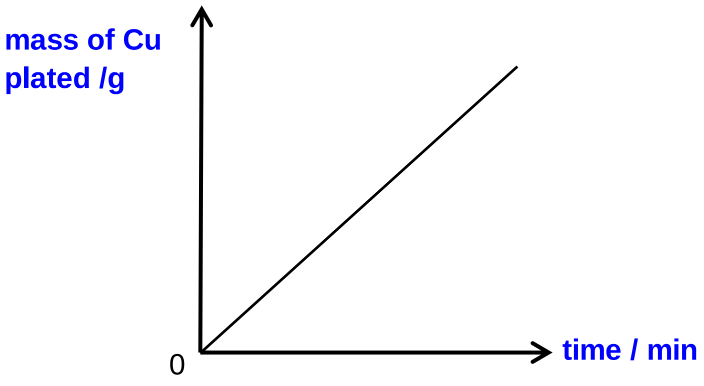


Procedure:

1. Weigh one of the objects to be plated.
2. Set up according to the diagram shown above. Ensure that the electrical power source is switched off.
3. Switch on the electrical power source and start the stopwatch simultaneously.
4. After 5 min, switch off the electrical power source.
5. Remove the plated object, wash the plated object with distilled water and dry it between pieces of filter paper.
6. Weigh the plated object. Then calculate the mass of copper plated by taking the difference between the mass of the plated object and the original mass of the object.
7. Repeat steps 1 to 6 with durations of 15 min, 25 min, 35 min, 45 min.

[6]

(c) Sketch, on the axes below, the graph you would expect and describe how it could be used to determine the value of Faraday constant.



Results analysis:

Let the gradient of the graph be k

$$k = \left(\frac{I \times M_{Cu}}{2F}\right)$$

$$F = \left(\frac{I \times M_{Cu}}{2k}\right)$$

[2]

(d) During the *copper plating* process, the mass lost at the anode should ideally be the same as the mass gained at the cathode. However, the presence of impurities can cause the mass change at one electrode to be numerically larger than the other electrode.

For each of the following, deduce, with explanation, the electrode that will have a numerically larger mass change.

You may find the following information useful:

$Ag^+ + e^- \rightleftharpoons Ag$	$E^\ominus = +0.80$ V
$Cu^{2+} + 2e^- \rightleftharpoons Cu$	$E^\ominus = +0.34$ V
$Pb^{2+} + 2e^- \rightleftharpoons Pb$	$E^\ominus = -0.13$ V

[A_r: Ag, 107.9; Cu, 63.5; Pb, 207.2]

(i) Presence of Pb impurity at the anode.

Electrode with numerically larger mass change: **<u>Anode</u>**

Explanation:

Pb gets preferentially oxidised at the anode but Cu^{2+} gets preferentially reduced at the cathode. For <u>every 1 mol of Pb (1 x 207.2 = 207.2 g) oxidised at the anode</u>, <u>1 mol of Cu (1 x 63.5 g= 63.5 g) is formed at the cathode</u>. Thus the mass change at the anode will be numerically larger.

[1]

(ii) Presence of Ag^+ impurity in the electrolyte.

Electrode with numerically larger mass change: **<u>Cathode</u>**

Explanation:

Cu gets oxidised at the anode but Ag^+ gets preferentially reduced at the cathode. For every 1 mol of Cu (1 x 63.5 = 63.5 g) oxidised at the anode, 2 mol of Ag (2 x 107.9 g= 215.8 g) is formed at the cathode. Thus the mass change at the cathode will be numerically larger.

[1]

(e) It is important to ensure that the cathode is dry at the start and at the end of each experiment.

Explain the effect, if any, on the calculated value of the Faraday constant if the cathode is dry at the beginning but wet at the end of each experiment.

Effect on calculated value of the Faraday constant: **Smaller**

Explanation:

If the cathode is dry at the beginning but wet at the end of each experiment, the mass of copper plated calculated for each experiment will be higher than actual. This will lead to a steeper gradient for the graph plotted. Thus calculated Faraday constant will be smaller than actual. [1]

(f) The Faraday constant obtained from the *copper plating* process can be used to provide an estimate for Avogadro's constant.

A student who conducted the *copper plating* process obtained a value of 96480 C mol^{-1} for the Faraday constant. Using this value, calculate Avogadro's constant.

[Note: The charge of an electron has a numerical value of 1.60×10^{-19} C.]

Avogadro's constant = $96480 / 1.60 \times 10^{-19}$

= 6.03×10^{23} electrons mol^{-1}

[1]

[Total: 14]

3 (a) Investigation of some inorganic reactions

FA 5 is an aqueous solution of $CuSO_4$.

FA 6 is a reducing agent.

FA 7 is a solution of ammonium vanadate(V), NH_4VO_3.

You will also need access to the **FA 2** solution you used earlier.

You are to perform the tests described in **Table 3.1** and record your observations in the table.

Test and identify any gases evolved. If there is no observable change, write **no observable change.**

Table 3.1

	tests	observations
(i)	Using a measuring cylinder, add 10 cm^3 of **FA 5** into a boiling tube. Add 4 spatulas of **FA 6**. Warm the mixture cautiously till boiling. Leave to cool for 5 minutes. Filter the mixture and keep the filtrate for tests **(ii)** and **(iii)**.	**Blue FA 5 turns colourless** **(OR award below for filtrate colour)** **Red-brown/brown/black/ residue** **colourless filtrate**
(ii)	To 1 cm depth of the filtrate from **(i)**, add aqueous ammonia.	**White ppt formed, soluble in excess NH_3 to give a colourless solution.**

(iii)	To another 1 cm depth of the filtrate from **(i)**, add 1 spatula of solid ammonium chloride. Stir to ensure all the solid ammonium chloride dissolves. Then add aqueous ammonia.	**No observable change (OR no ppt)**

Table 3.1

tests		observations
(iv)	Using a measuring cylinder, transfer 2 cm^3 of **FA 7** and 10 cm^3 of **FA 2** into a boiling tube. Swirl the mixture gently. The resultant yellow solution is **FA 8**, which is an acidified solution of VO_2^+.	
(v)	Transfer about 2 spatulas of **FA 6** into a **dry** weighing bottle. From this sample in the weighing bottle, add a **very small** quantity of **FA 6** to the boiling tube containing **FA 8** solution from test **(iv)**. Swirl the mixture gently and record your observations. Continue to add more **FA 6** in **small** quantities with swirling, until all the **FA 6** in the weighing bottle is used up. Record all colour changes observed.	**Yellow solution turns green [mixture of yellow VO_2^+ and blue VO^{2+}]** **Green solution turns blue (OR bluish-green OR greenish blue) [more VO^{2+} formed]** **Blue solution turns green (OR dark green) [V^{3+} formed]** **Green solution turns purple (OR violet OR blue) [V^{2+} formed]** **Effervescence of H_2 gas which** **extinguished lighted splint with a 'pop sound**.
	Filter the mixture and retain the filtrate for test **(vi)**.	

(vi)	To 1 cm depth of the filtrate from test **(v)**, add an equal volume of aqueous hydrogen peroxide.	**Violet solution turns brown (OR orange OR red)** **Effervescence of O_2 gas which** **relighted glowing splint.**

[6]

(b) **(i)** Suggest an identity for **FA 6**. Explain your answer with reference to evidence from relevant tests above.

identity of **FA 6**
Zn

explanation

In test (i), decolourisation of blue $CuSO_4$ suggests reduction of Cu^{2+} to Cu and oxidation of FA 6

In test (ii), formation ofwhite ppt soluble in excess $NH_3(aq)$ suggests presence of Zn^{2+} in filtrate from test (i).

[2]

(ii) Explain the difference in observations between tests **(a)(ii)** and **(a)(iii)**.

$NH_4Cl \rightarrow NH_4^+ + Cl^-$

$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$ ----(1)

Presence of additional NH_4^+ from NH_4Cl (OR NH_4^+ is a common ion) shifts position of equilibrium (1) to the left.

$[OH^-]$ is lowered and ionic product is lesser than K_{sp}. Hence, no ppt is formed in (a)(iii).

[2]

(iii) The half equations that represent the stepwise reduction of VO_2^+ ions are shown below.

Electrode reaction	$E^\ominus$ / V
$VO_2^+ + 2H^+ + e^- \rightleftharpoons VO^{2+} + H_2O$	+1.00
$VO^{2+} + 2H^+ + e^- \rightleftharpoons V^{3+} + H_2O$	+0.34
$V^{3+} + e^- \rightleftharpoons V^{2+}$	–0.26
$V^{2+} + 2e^- \rightleftharpoons V$	–1.20

When excess aqueous ammonia is added to 1 cm depth of the filtrate obtained in test **(a)(v)**, a grey precipitate is observed.

Given that the $E^\ominus$ value for **FA 6** is between −1.00 V to –0.50 V, state the identity of the grey precipitate.

$E^\ominus$ values suggest that FA 6 reduces VO_2^+ to V^{2+}. Hence, the grey ppt is $\underline{V(OH)_2}$.

[1]

(c) **Planning**

There are three unlabelled bottles of aqueous solutions. Each bottle contains a different solution from the other two.

The possible identities of the solutions are:

Na_2CO_3(aq) $Al_2(SO_4)_3$(aq) $Mg(NO_3)_2$(aq)

Plan an investigation, using test–tube reactions, which would allow you to identify each of these three solutions.

Other than the three solutions, the only reagent available is the **FA 5** solution used in **3(a)**, which contains $CuSO_4$.

Your plan should include:

- an outline of the sequence of steps you would follow,
- the expected observations at each step, and
- an explanation of how you would analyse your results in order to identify each solution.

You are **not** required to carry out the plan.

	Na_2CO_3(aq)	**$Al_2(SO_4)_3$(aq)**	**$Mg(NO_3)_2$(aq)**
FA 5 (aq $CuSO_4$)	**Blue (OR green) ppt**	**no ppt**	**no ppt**
Na_2CO_3(aq)	**---**	**White ppt** **CO_2 effervescence**	**White ppt** **No effervescence**

Add <u>FA 5</u> to each of the three solutions in separate test–tubes. The solution that gives a <u>blue (OR green) ppt</u> can be identified as $\underline{Na_2CO_3}$ while the <u>other two solutions</u> give <u>no ppt</u>.

Add $\underline{Na_2CO_3(aq)}$ to fresh samples of the two remaining unidentified solutions in separate test–tubes.

The solution that gives <u>white ppt</u> and CO_2 <u>effervescence</u> can be identified as $\underline{Al_2(SO_4)_3}$ while the solution that gives <u>white ppt</u> with <u>no effervescence</u> is $\underline{Mg(NO_3)_2}$.

[3]

[Total: 14]