Presentation of Data & Observation

Question Paper 2

Level	Pre U
Subject	Chemistry
Exam Board	Cambridge International Examinations
Topic	Presentation of data & Observation
Booklet	Question Paper 2

Time Allowed: **75 minutes**

Score: **/65**

Percentage: **/100**

Grade Boundaries:

1 A student suggests that the concentration of sulfuric acid can be determined by measuring the temperature of the solution as the acid is added in small amounts to a known volume of sodium hydroxide solution in a plastic cup.

$$2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$$

The student proposes the following hypothesis.

As the acid is added to the alkali the temperature rise will be directly proportional to the volume of acid added until the end-point of the reaction is reached. Upon further addition of acid there will be a reduction in the temperature of the solution in the cup as the acid added is not reacting and is at a lower temperature than the solution in the plastic cup.

The following reagents are provided.

FA 1 is 2.00 mol dm^{-3} sodium hydroxide, NaOH.
FA 2 is **approximately** 0.75 mol dm^{-3} sulfuric acid, H_2SO_4.

(a) Use the equation for the reaction to estimate the volume of **FA 2** that will neutralise 25.0 cm^3 of **FA 1**.

volume of **FA 2** = ... cm^3 [1]

(b) In the experiment you will add **FA 2** from the burette to 25.0 cm^3 of **FA 1** in a plastic cup. You will measure the temperature of the solution after each addition of a certain volume of acid. You will then plot a graph of the temperature rise against the volume of acid added and use this to determine the end-point. You will then be able to calculate the concentration of H_2SO_4 in **FA 2**.

In order to obtain precise information about the end-point of the reaction, you will need to decide:

- the volume of acid to be added each time (do not use a volume which is less than 2.00 cm^3)
- the total volume of acid to be added.

volume of acid to be added each time = ... cm^3

total volume of acid to be added = ... cm^3 [2]

(c) Method

1. Fill the burette with **FA 2**.
2. Support the plastic cup in the 250 cm^3 beaker.
3. Pipette 25.0 cm^3 of **FA 1** into the plastic cup.
4. Measure and record the temperature of **FA 1** in the plastic cup.
5. Add the first volume of **FA 2** from the burette into the plastic cup. Stir the solution and record the highest temperature that is observed.
6. Continue to add each volume of **FA 2** and record the highest temperature observed.

Record in the space below:

- the initial temperature of **FA 1**
- the total volume of **FA 2** added at each stage in the experiment
- the temperature of the solution in the plastic cup after each addition of acid
- the temperature rise, ΔT, where ΔT = highest temperature of the solution after each addition of acid – initial temperature of **FA 1**.

[6]

(d) On the grid below plot the temperature rise, ΔT, (*y*-axis) against the volume of **FA 2** added (*x*-axis).

[4]

(e) (Use your graph to obtain a value for the volume of **FA 2** added at the end-point of the titration.

volume of **FA 2** at the end-point = ... cm^3 [1]

(ii) Use your answer to **(i)** to calculate the concentration of H_2SO_4 in **FA 2**.
Show your working.

concentration of **FA 2** = ... $mol\,dm^{-3}$ [2]

(f) Explain how the results of your experiment support or do not support each part of the hypothesis proposed by the student.

....

....

....

.... [2]

(g) Calculate the % error in the total volume of **FA 2** added from the burette for the volume which is closest to the end-point.

.... % [2]

(h) A student carrying out the same experiment noticed that each subsequent temperature rise became less as the end-point was approached. Give **two** reasons why this was the case.

reason 1 ....

....

reason 2 ....

.... [2]

(i) Another student put forward the hypothesis that the heat energy produced in the reaction, rather than the temperature rise, is proportional to the volume of acid added.

Calculate the total heat produced by the addition of **FA 2** at the end-point.
Assume that it takes 4.2 J to raise the temperature of 1.0 cm^3 of solution by 1.0 °C.

heat produced = ... J [1]

[Total: 23]

2 Another way to analyse a mixture of sodium carbonate and sodium hydrogencarbonate is to measure the temperature change that occurs when a known mass of the mixture is added to acid.

FA 3 is a mixture of anhydrous sodium carbonate, Na_2CO_3, and sodium hydrogencarbonate, $NaHCO_3$. This mixture is **not** the same as **FA 1**.

FA 4 is 2.00 mol dm^{-3} hydrochloric acid, HCl.

(a) Method

Before starting any practical work, read through all the instructions and prepare a suitable table for your results in the space provided.

1. Support a plastic cup in a 250 cm^3 beaker.
2. Using a 25 cm^3 measuring cylinder, pour 25 cm^3 of **FA 4**, the hydrochloric acid, into the plastic cup.
3. Measure the temperature of the acid in the cup.
4. Weigh the bottle containing **FA 3**.
5. Add the contents of the bottle to the acid in a number of portions to avoid acid spray.
6. Use the thermometer to stir the mixture gently.
7. Measure the temperature that is reached.
8. Reweigh the bottle.

Record all the measurements from your experiment. Include the mass of **FA 3**, M, added to the acid and the change in temperature, ΔT, where ΔT = final temperature – initial temperature.

[4]

(b) In the following calculations you will work out the percentage by mass of sodium hydrogencarbonate in **FA 3**.

You must show your working.

(i) Calculate the change in temperature for each gram of **FA 3** added to the acid.

$\frac{\Delta T}{M}$ = ...°C g^{-1}
sign

(ii) A student carried out similar experiments using separate samples of the two salts.

For pure sodium carbonate the change in temperature per gram was +3.38 °C g^{-1}.

For pure sodium hydrogencarbonate the change in temperature per gram was –2.74 °C g^{-1}.

Use these values and your answer to **(b)(i)** to calculate the mass and the percentage by mass of sodium hydrogencarbonate in **FA 3**.

If your answer to **(b)(i)** does **not** lie between +3.38 and –2.74 °C g^{-1} then assume that the answer to **(b)(i)** is +1.36 °C g^{-1}. This is not the correct value.

mass of sodium hydrogencarbonate = ... g

percentage by mass of sodium hydrogencarbonate = ... %

[5]

(c) **(i)** State the uncertainty in the measurement of **each** mass in this experiment.

uncertainty = ± .... g

(ii) Calculate the percentage error in the mass of **FA 3** that was used.

percentage error = ... %
[1]

(d) The experiment was repeated using the **same** apparatus but this time twice the mass of **FA 3** was added to 50 cm^3 of the hydrochloric acid.

Discuss whether this would determine more accurately the percentage of sodium hydrogencarbonate in **FA 3**.

....

....

....

....[2]

[Total: 12]

3. The acid catalysed reaction between iodine and propanone proceeds relatively slowly at room temperature.

$$CH_3COCH_3(aq) + I_2(aq) \rightarrow CH_3COCH_2I(aq) + H^+(aq) + I^-(aq)$$

The rate of this reaction can be determined by following the change in the concentration of iodine as the reaction proceeds. At set time intervals, a small volume of the reaction mixture is removed and added to a solution of sodium hydrogen carbonate. This effectively stops the reaction by neutralising the acid catalyst. The concentration of iodine can then be determined by titration using sodium thiosulfate solution.

$$I_2(aq) + 2S_2O_3^{2-}(aq) \rightarrow 2I^-(aq) + S_4O_6^{2-}(aq)$$

In the following experiment you will carry out this titration to determine the concentration of iodine.

You are provided with a solution labelled **FA 1**. This solution was obtained by removing 25.00 cm^3 of the reaction mixture and adding excess sodium hydrogen carbonate solution. The mixture was then made up to 150.0 cm^3 using distilled water.

FA 2 contains 0.0500 mol dm^{-3} sodium thiosulfate solution, $Na_2S_2O_3(aq)$.

Read all of the following instructions before you start any experimental work and draw up a table to record your results in part **(a)**.

Method

1. Fill a burette with **FA 2**.
2. Pipette 25.00 cm^3 of **FA 1** into a conical flask.
3. Run the solution from the burette into the conical flask until the red/brown colour of the iodine becomes pale yellow.
4. At this point add approximately 10 drops of starch indicator.
5. Continue to add **FA 2** until the blue/black colour completely disappears.
6. Repeat the titration until you have obtained consistent results.

(a) Record your titration results in the space below. Make sure your recorded results show the precision of your practical work.

[5]

(b) From your titration results obtain a volume of **FA 2** to be used in the following calculations.
Show clearly how you obtained this value.

25.00 cm^3 of **FA 1** required .... cm^3 of **FA 2**. [1]

(c) From the measurements you have made, determine the concentration of the iodine in the reaction mixture at the time when the 25.00 cm^3 sample was removed.
Show your working and give your answer to an appropriate number of significant figures.

.... [5]

(d) In another experiment using different starting concentrations of iodine, propanone and acid, the following values were obtained for the concentration of iodine in the reaction.

time / min	$[I_2(aq)]$ / mol dm^{-3}
0	0.0100
5	0.0096
10	0.0092
15	0.0088

Use the figures shown above to determine the order of the reaction with respect to iodine. Explain your answer.

....

....

....[2]

[Total: 13]

4. Epsom salts occur naturally and are a hydrated form of magnesium sulfate, $MgSO_4.xH_2O$.

In the following experiment you will determine the value of *x*.
Read all of the following instructions before you start any experimental work.

You are provided with the following:
FA 1 hydrated magnesium sulfate, $MgSO_4.xH_2O$

Method

1. Weigh a clean, dry crucible.
2. In the crucible place the entire sample of Epsom salts, **FA 1**.
3. Reweigh the crucible.
4. Place the crucible in a pipe-clay triangle on top of a tripod.
5. Heat the crucible **gently** for about 1 minute and then more strongly for a further 4 minutes.
6. Allow the crucible to cool for about 1 minute and then use a pair of tongs to place the crucible on a heat proof mat.
7. Leave the crucible to cool for approximately three minutes, then reweigh the crucible and its contents.
8. Repeat the cycle of heating and weighing, as described in steps 4 to 7, until consecutive recorded masses do not differ by more than 0.05 g.

(a) In a suitable table, record all masses.
Calculate the mass of the residue and the mass of the water lost. Record both of these masses in the table.

[8]

(b) From the measurements you have made, determine the value of *x* in the formula $MgSO_4.xH_2O$.

Show your working.

....[4]

(c) **(i)** State the uncertainty in the measurement of each mass in this experiment.

uncertainty = ± .... g [1]

(ii) Calculate the percentage error in the mass of water that is lost.

Show your working.

....[2]

(d) Suggest an improvement that a student might make to the experiment and explain why this would lead to the determination of a more accurate value of *x*.

....

....

....[2]

[Total: 17]