RIVER VALLEY HIGH SCHOOL

YEAR 6 PRACTICAL EXAMINATION II

H2 CHEMISTRY 9729

23 AUGUST 2018

2 HOURS 30 MINUTES

NAME _____

CLASS 6 ( )

INDEX NO. _____

INSTRUCTIONS TO CANDIDATES

DO NOT OPEN THIS BOOKLET UNTIL YOU ARE TOLD TO DO SO.

Read these notes carefully.

Write your name, class and index number in the spaces at the top of this page.
Give details of the practical shift and laboratory where appropriate, in the boxes provided.
Write in dark blue or black pen.
You may use a 2B pencil for any diagrams or graph.
Do not use staples, paper clips, highlighters, glue or correction fluid.

*Answer **all** questions in the spaces provided on the Question Paper.*

The use of an approved scientific calculator is expected, where appropriate.

You may lose marks if you do not show your working or if you do not use appropriate units.

Shift
Laboratory

For Examiner's Use	
3 s.f.	
Units	
Total	**55**

This Question Paper consists of **14** printed pages and **1** blank page.

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

For Examiner's Use

1 To investigate the kinetics of the reaction between iron(III) ions and iodide ions

You will investigate rate of oxidation of iodide by iron(III) ions in **reaction 2**. The chemical changes in this experiment can be presented by the following equations:

Reaction 1: $Fe^{3+}(aq) + S_2O_3^{2-}(aq) \rightleftharpoons [Fe(S_2O_3)]^+(aq)$

Reaction 2: $2Fe^{3+}(aq) + 2I^-(aq) \rightarrow 2Fe^{2+}(aq) + I_2(aq)$

Reaction 3: $I_2(aq) + 2S_2O_3^{2-}(aq) \rightarrow 2I^-(aq) + S_4O_6^{2-}(aq)$

The reaction is started by mixing acidic solution of iron(III) chloride with sodium thiosulfate, potassium iodide, and starch. The iodine, I_2, produced in **reaction 2** reacts immediately with thiosulfate ions, $S_2O_3^{2-}$ in **reaction 3**.

When all the thiosulfate has been used, the iodine produced will turn starch indicator blue-black. The rate of the reaction can therefore be determined by finding the time for the blue-black colour to appear.

FA 1 is 0.030 mol dm^{-3} iron(III) chloride, $FeCl_3$.

FA 2 is 0.060 mol dm^{-3} potassium iodide, KI.

FA 3 is 0.0060 mol dm^{-3} sodium thiosulfate, $Na_2S_2O_3$.

starch indicator

You are advised to read the instructions before starting any practical work and draw a table, in an appropriate format, for your results in the space on page 4.

You will need to include volume of **FA 1**, volume of water, reaction time and calculated rate of reaction for each of the five experiments. Record all calculated values to 3 significant figures.

(a) Method

Experiment 1

1. Use the measuring cylinders to place the following in a 100 cm^3 beaker.
    - 10 cm^3 of **FA 2**
    - 20 cm^3 of **FA 3**
    - 10 cm^3 of starch indicator
2. Using a measuring cylinder, transfer 20 cm^3 of **FA 1** rapidly into the same 100 cm^3 beaker. Start the stopwatch during this addition.
3. Stir the mixture and place the beaker on a white tile.
4. The mixture turns brown and then yellow before turning a blue-black colour. Stop timing when this blue-black colour appears. Record the time to the nearest 0.1 second.
5. Wash the beaker thoroughly with water and carefully dry the beaker.

Experiment 2

6. Repeat step **1** in **Experiment 1**.
7. Using a measuring cylinder, add 12 cm^3 of **FA 1**, and make up the volume to 20 cm^3 using deionised water. Transfer the solution rapidly into the same 100 cm^3 beaker. Start the stopwatch during this addition.
8. Stir the mixture and place the beaker on the white tile.
9. Stop timing when a blue-black colour appears.
10. Wash the beaker thoroughly with water and carefully dry the beaker.

Experiments 3−5

Carry out three further experiments to investigate the effect of changing the concentration of $Fe^{3+}(aq)$ by altering the volume of aqueous $FeCl_3$, **FA 1**, used.

You should use a volume of **FA 1** that is at least 12 cm^3 and the total volume of the reaction mixture must always be 60 cm^3.

(b) Show, by means of calculation, that the change in the concentration of Fe^{3+}, $\Delta[Fe^{3+}]$, which occurred when the blue-black colour appeared was 2.00×10^{-3} mol dm^{-3}.

For Examiner's Use

$$\text{Amount of } S_2O_3^{2-} = \frac{20}{1000} \times 0.006 = 1.20 \times 10^{-4} \text{ mol}$$

$$\Delta[Fe^{3+}] = \frac{1.20 \times 10^{-4}}{\frac{60}{1000}} = 2.00 \times 10^{-3} \text{ mol dm}^{-3}$$

1	

(c) Results

The rate of the reaction can be calculated as shown.

$$\text{rate} = \frac{\Delta[Fe^{3+}]}{\text{reaction time}} \times 10^6$$

Calculate the rate of reaction for each experiment and complete your table.

Expt	V_{FA1}/ cm^3	V_{H_2O} /cm^3	Reaction time, ***t*** /s	Rate/ mol dm^{-3} s^{-1}
1	20.00	0.00	12.9	155
2	12.00	8.00	28.8	69.4
3	14.00	6.00	20.7	96.6
4	16.00	4.00	17.3	115
5	18.00	2.00	14.7	136

2	
3	
4	
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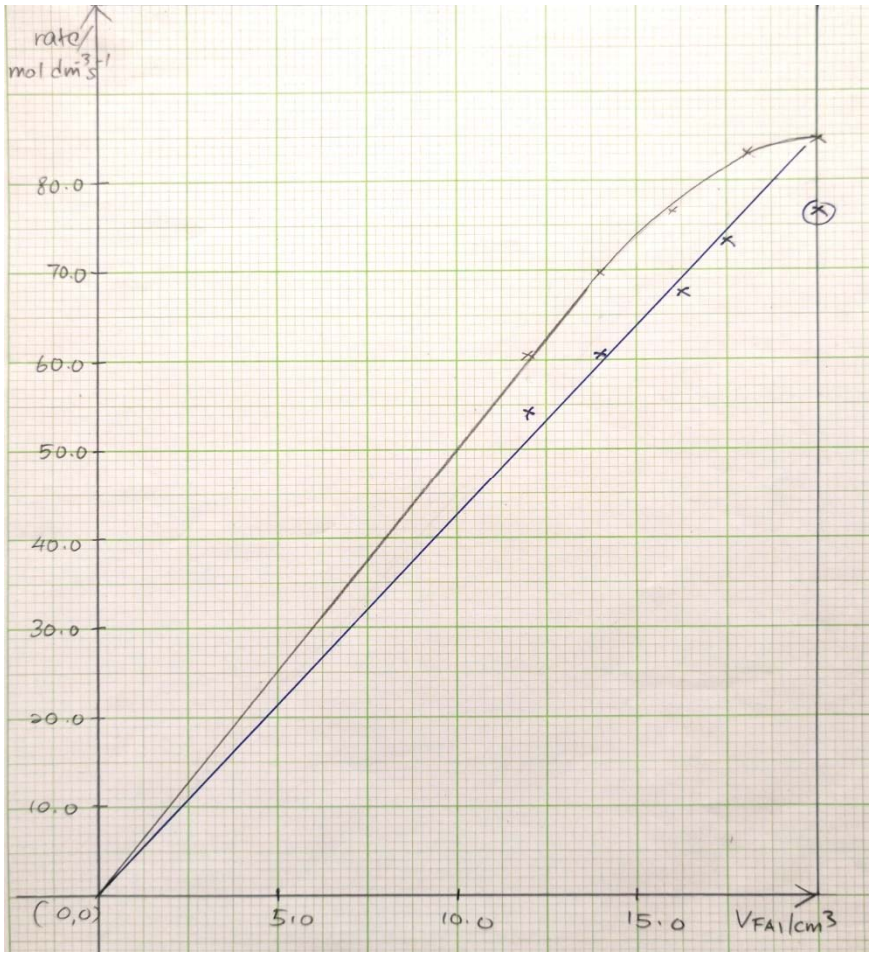
(d) (i) Plot the rate (y-axis) against the volume of **FA 1** (x-axis) on the grid. Draw a line of best fit through the points.

Your chosen scales should allow you to extrapolate this line to volume of **FA 1** = 0 cm^3.

You should also identify any points you consider anomalous by drawing a circle around it.



6	
7	
8	

(ii) Explain why the volume of **FA 1** in each experiment can be used as the [Fe^{3+}].

Since total volume is kept constant by adding deionised water, volume of **FA 1** used is proportional to [Fe^{3+}].

9	

(iii) Deduce the order of reaction with respect to Fe^{3+}. Use evidence from your graph to support your deduction.

✓ Observe from the graph, any one of the following:
- the rate increase by xxx for every additional xxx cm^3 of **FA 1** used
- when volume of **FA 1** double, rate double (or other multiples)
- the graph is linear/gradient is a constant, showing that rate ∝ volume/ concentration of **FA 1** / state that it is a graph of rate against [Fe^{3+}]

✓ Therefore, the reaction is first order with respect to Fe^{3+}.

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(e) **(i)** Use your graph to calculate the time that the reaction would have taken if 4.0 cm^3 of **FA 1** had been used. Show your working clearly.

If 4.0 cm^3 of **FA 1** is used, rate = xxx mol dm^{-3} s^{-1}
Time = $\Delta[Fe^{3+}] \times 10^6$/ rate = xxx s (to 3 sf or nearest seconds).

time = ... | 11 |

(ii) Calculate the initial concentration of iron(III) ions and the initial concentration of iodide ions if 4.0 cm^3 of **FA 1** had been used.

$[Fe^{3+}] = \frac{0.030 \times 4.0}{60.0}$ = 0.00200 mol dm^{-3}

$[I^-] = \frac{0.060 \times 10.0}{60.0}$ = 0.0100 mol dm^{-3}

initial $[Fe^{3+}]$ = ...
initial $[I^-]$ = ... | 12 |

(iii) Given that the reaction is second order with respect to iodide, calculate the rate constant.

rate = $k[Fe^{3+}][I^-]^2$

$k = \frac{rate}{[Fe^{3+}][I^-]^2}$ = xxx mol^{-2} dm^6 s^{-1}

rate constant = ... | 13 |

(f) The complex $[Fe(S_2O_3)]^+$ formed in **reaction 1** is purple.

Consider the colour change observed when **FA 1** was added to the solution prepared in **Step 1** before the blue-black colour appeared. Suggest an explanation for the sequence of colour change in terms of the chemistry involved.

The brown solution is due to a mixture of purple complex $[Fe(S_2O_3)]^+$ and orange-brown $Fe^{3+}(aq)$.

Since the I_2 formed reacted immediately with thiosulfate, $[S_2O_3^{2-}(aq)]$ decreased, and the position of equilibrium in (1) shifts left.

Thus, concentration of $[Fe(S_2O_3)]^+$decreased. Hence, the colour of the solution gradually fades/ until the solution became pale yellow due to Fe^{3+}. | 14 | 15 |

[Total: 15]

For Examiner's Use

2 Determining the percentage purity of calcium carbonate.

In this question, you will determine the percentage purity of an industrial grade calcium carbonate, $CaCO_3$, by two different methods.

(a) Method 1

The percentage purity of calcium carbonate can be determined by measuring the change in mass when the sample of calcium carbonate reacts with hydrochloric acid.

$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$

FA 4 is industrial grade calcium carbonate, $CaCO_3$.

FA 5 is 2.00 mol dm^{-3} dilute hydrochloric acid.

You do not need to carry out **Method 1**.

Procedure:

1. Weigh accurately about 2.0 g of **FA 4** in a small weighing bottle.
2. Weigh accurately about 25 cm^3 of **FA 5** in a small conical flask.
3. Add **FA 4** to **FA 5** and swirl the mixture continuously until no more bubbles is observed.
4. Reweigh the weighing bottle with its residual **FA 4**.
5. Reweigh the conical flask and its content.

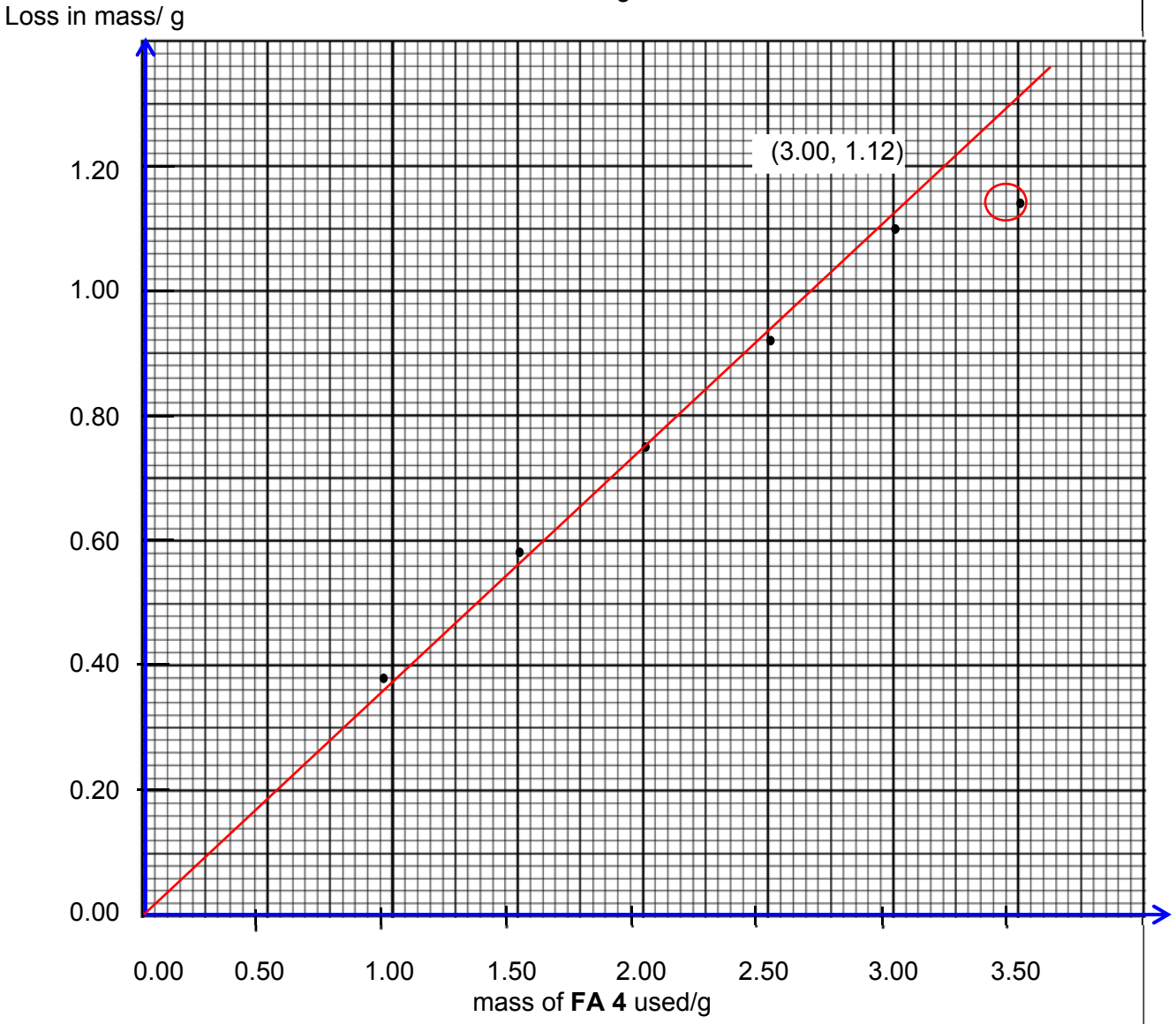
A student performed a series of experiments by repeating the above procedure with different masses of **FA 4**.

The results from her series of experiment are plotted on the grid in Fig 2.1.

Identify the anomalous data point by drawing a circle around it. Draw the most appropriate best-fit line taking into account all of the plotted points, except for the anomaly. Extrapolate (extend) this line to the mass of **FA 4** used = 0.00 g.

For Examiner's Use

Fig 2.1.



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Calculations

For Examiner's Use

(b) **(i)** Determine the gradient of this line, showing clearly how you did this.

Hence, determine the mass of carbon dioxide evolved per gram of **FA 4** used.

gradient = $\frac{1.12-0.00}{3.00-0.00}$ =0.3733

mass lost = mass of carbon dioxide evolved

= 0.3733 g per gram of **FA 4** used

mass of carbon dioxide evolved = 0.3733 g .....

17	
18	

(ii) Using your answer in **b(i)**, calculate a value for the percentage purity of the sample of industrial grade calcium carbonate, **FA 4**.

[A_r: Ca, 40.1; C, 12.0; O,16.0]

For every gram of **FA 4** used:

Amount of CO_2 evolved = $\frac{0.3733}{44}$ = 0.008484 mol

= amount of $CaCO_3$

Mass of $CaCO_3$ = 0.008484 × 100.1 = 0.8492

percentage purity of **FA 4** = $\frac{0.8492}{1.00} \times 100\%$ = 84.9 %

percentage purity of **FA 4** = ...

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(c) **Method 2**

You will determine the amount of hydrochloric acid remaining in flask **X** after the reaction with another sample of industrial grade calcium carbonate. You will do this by titration with sodium hydroxide of known concentration.

NaOH(aq) + HC*l*(aq) → NaC*l*(aq) + H_2O(l)

FA 6 is 0.140 mol dm^{-3} sodium hydroxide, NaOH.

Thymolphthalein indicator

Procedure:

1. Pipette 25.0 cm^3 of **FA 5** into flask **X**.
2. Weigh accurately about 0.6 g of **FA 4** in a dry weighing bottle.
3. Add the **FA 4** into flask **X**.
4. Reweigh the weighing bottle with its residual **FA 4**.
5. Swirl the mixture until no more bubbles are observed.
6. Transfer **all** the contents of flask **X** and the washings into a 250 cm^3 volumetric flask.
7. Add deionised water up to the mark, stopper the volumetric flask and mix the contents thoroughly. Label this solution **FA 7**.

8. Fill the burette with **FA 6**.
9. Wash and rinse the pipette then use it to transfer 25.0 cm^3 of **FA 7** into a conical flask.
10. Add about 5 drops of thymolphthalein indicator.
11. Titrate **FA 7** with **FA 6**.
12. Repeat the titration as many times as you think necessary to obtain consistent results.
13. Record in the space below, all of your mass readings, burette readings and the volume of **FA 6** added.

Results

Mass of weighing bottle and **FA 4**/ g	4.433
Mass of weighing bottle and residual **FA 4**/g Or Mass of weighing bottle and residue/g	3.823
Mass of **FA 4** used/ g	0.610

Titration number	1	2
Final burette reading /cm^3	29.30	29.30
Initial burette reading /cm^3	0.00	0.00
Volume of **FA 6** (used) /cm^3	29.30	29.25

20	
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25	

Supervisor expected titre
28.53
Student expected titre
Difference

(d) From your accurate titration results, obtain a suitable value for the volume of **FA 6** to be used in your calculations. Show clearly how you obtained this value.

$$\frac{29.30 + 29.25}{2} = 29.28\ cm^3$$

Volume of **FA 6** used = ……

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Calculations

For Examiner's Use

(e) **(i)** Calculate the number of moles of sodium hydroxide, NaOH, present in volume of **FA 6** calculated in **(d)**.

amount of NaOH in titre = $\frac{titre}{1000} \times 0.140$ = _____ mol

amount of NaOH = ... **27**

(ii) Calculate the number of moles of hydrochloric acid remaining in flask **X** after the reaction with **FA 4** has completed.

amount of NaOH in titre = amount of HC*l* in 25.0cm^3 **FA 7**

amount of HC*l* remaining in flask **X**

= amount of HC*l* in 250 cm^3 of **FA 7**

= $\frac{titre}{1000} \times 0.140 \times \frac{250}{25.0}$

amount of HC*l* remaining in flask **X** = ... **28**

(iii) Calculate the number of moles of hydrochloric acid that reacted with the **FA 4** used.

Total amount of HC*l* used = $2.00 \times 0.0250 = 0.0500$ mol

Amount of HC*l* reacted = $0.0500 - (\frac{titre}{1000} \times 0.140 \times \frac{250}{25.0})$

amount of HC*l* reacted = ... **29**

(vi) Calculate the mass of $CaCO_3$ in the sample of **FA 4** used, and hence the percentage purity of the industrial grade $CaCO_3$.

[Ar: Ca, 40.1; C, 12.0; O,16.0]

Mass of $CaCO_3$ in the sample = $\frac{n_{HCl\ reacted}}{2} \times 100.1$

Percentage purity = $\frac{mass\ of\ CaCO_3 in\ the\ sample}{mass\ of\ sample\ used} \times 100\%$

percentage purity of the industrial grade $CaCO_3$ = ... **30**

(f) The two different methods were used to find the percentage purity of industrial grade calcium carbonate.

It is found that the samples of **FA 4** used contain calcium oxide which reacts with dilute hydrochloric acid.

How would this affect the percentage purity calculated for each method? Explain your answers.

Method 1: Percentage purity of **FA 4** calculated is not affected.

Reaction of calcium oxide with hydrochloric acid does not result in loss in mass/ does not produce any gases.

Method 2: Percentage purity of **FA 4** calculated is higher than expected.

Less HC*l* is left after reaction with sample. Based on the lower titre obtained/ less NaOH is needed to react with the remaining HC*l* , a greater amount of $CaCO_3$ is calculated to be present.

31	
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(g) Explain why sulfuric acid cannot be used in both methods.

The insoluble $CaSO_4$/ salt formed coats the calcium carbonate particles, which results in incomplete reaction/ owtte.

33	

(h) Planning

Unlike the acid-base titration in **Part (c)** that uses acid-base indicator to determine the end-point, thermometric titration uses heat evolved or temperature change of solution to determine the end-point of the reaction. In thermometric titration, the titre value can be determined from the graph of energy evolved against volume of titrant (solution from the burette) added.

Since a neutralisation reaction is exothermic, thermometric titration can be used to determine the amount of excess hydrochloric acid in flask **X** from **Part (c)**.

(i) Given that the concentration of undiluted excess hydrochloric acid in flask **X** is about 1.7 mol dm^{-3}, and the desired titre is around 24 cm^3, calculate the concentration of sodium hydroxide required for the thermometric titration.

$[NaOH] = \frac{0.025 \times 1.7}{0.024} = 1.77$ mol dm^{-3}

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(ii) You are to plan a thermometric titration to determine the concentration of hydrochloric acid. The experimental results obtained should allow the titre value to be determined graphically.

You may assume that you are provided with

- 50 cm^3 of hydrochloric acid, HC*l*, of concentration around 1.7 mol dm^{-3}
- standard solution of sodium hydroxide solution, NaOH, of the same concentration calculated in **(h)(i)**
- apparatus commonly found in a school or college laboratory

In your plan you should include the following:

- a diagram of the apparatus set-up that may be used
- the procedure that you would follow and the measurement you would take
- the precautions and measures that you would take to minimise heat loss to the surroundings

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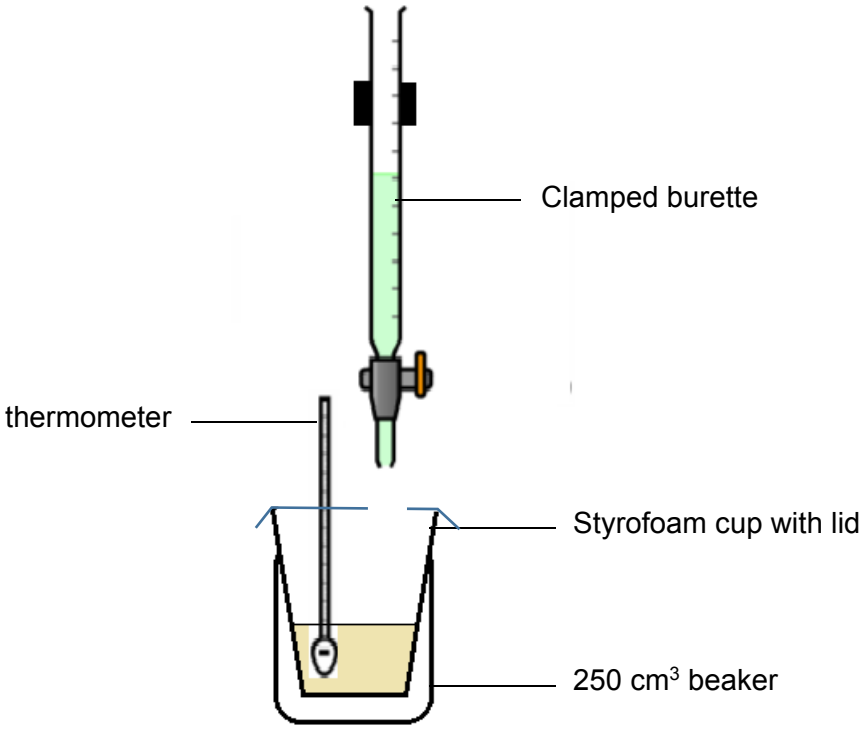

Procedure

1. Set up the apparatus as shown in the diagram and fill the burette with 1.77 mol dm^{-3} aqueous NaOH.
2. Using a burette/pipette, transfer 25.0 cm^3 of the HC*l* solution into the styrofoam cup, and measure its initial temperature.
3. Using the burette, add 5.00 cm^3 of NaOH solution to the HC*l* solution in the cup. Stir by swirling the cup/ stir using the thermometer and measure the highest temperature reached.
4. Immediately add another 5.00 cm^3 of the NaOH solution, stir, and again measure the highest temperature.
5. Repeat step 4 until the total volume of NaOH added is 45.00 cm^3.

Precaution: This experiment should be carried out in a draught-free room with the fans switched off.

(iii) Your procedure in **(h)(ii)** should allow the graph of total energy evolved against volume of titrant to be plotted. Briefly describe how you would use the results obtained to determine the total energy evolved since the start of the experiment for each data point.

Energy evolved = $mc\Delta T$

Where m = (Volume of HC*l* + total volume of NaOH added/ total volume of reaction mixture) × density , c = specific heat capacity of the solution/ 4.18 $Jg^{-1}K^{-1}$, ΔT = highest temperature – initial temperature of HC*l* in step x.

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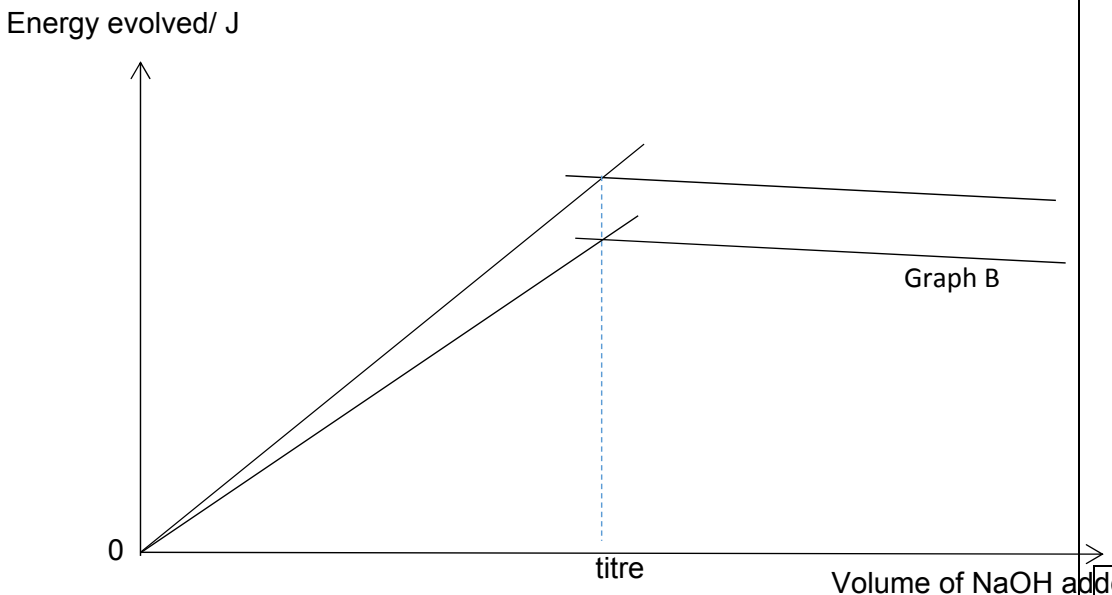
(iv) Sketch, on the axes in Fig. 2.2, the graph you would expect to obtain and show how you would use this graph to determine the titre value.

Explain the shape of your graph.

Fig. 2.2



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Explanation:

As NaOH is added, more neutralisation occurs which evolves more heat. The total amount of heat evolved reaches maximum at the equivalence point when all HC/ just reacted. Past equivalence point, the graph is horizontal because no more heat is evolved/Graph is downward sloping due to heat lost to the surroundings'.

(v) Sketch, on the same axes in Fig. 2.2, the graph you would expect to obtain if aqueous ammonia of the same concentration as aqueous NaOH is used as the titrant instead.

Label this graph clearly as **Graph B**.

Explain the shape of **Graph B**.

Graph B

Ammonia is a **weak base** which dissociates partially to produce OH^-.

Heat is absorbed to **complete the dissociation/ break the O-H in water** to produce OH^-**.** Thus, the reaction is **less exothermic/ less energy is evolved.**

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45	

3 Inorganic qualitative analysis

In this experiment you are to explore the chemistry of some compounds of an unknown transition element **W**. You are provided with 4 samples, **FA 8**, **FA 9**, **FA 10** and **FA 11**.

FA 8 is a solid sample of a common dioxide of the unknown transition element **W**.

FA 9 is a dilute sulfuric acid, H_2SO_4.

FA 10 is a solid sample of sodium ethanedioate, $Na_2C_2O_4$.

FA 11 is a solution of pure compound **X**, which is the product formed in **(a)(i)**.

Carry out the following experiments. Carefully record your observations. Test any gases produced.

	test	observations
(a)(i)	Transfer all of the solid sample of **FA 10** into a 100 cm^3 beaker. Add 25 cm^3 of **FA 9** to this beaker. Gently heat the beaker until the temperature of the mixture reaches about 60°C. Stir the mixture carefully. Place the beaker on the white tile. Add 2 – 3 spatula of **FA 8** to the mixture, stir the mixture carefully with thermometer and observe any changes in the temperature of the mixture. Filter the mixture into a boiling tube when you think that the reaction is completed. Leave the filtrate to stand. The filtrate contains compound **X**. Retain this filtrate for use **(a)(ii)**.	White solid ($Na_2C_2O_4$) dissolves to give a colourless solution. [accept: partially dissolves] Effervescence/ bubbles after addition (not gas evolved) Colourless, odourless gas gave white ppt with limewater Temperature of mixture rises (even though no heat is applied) Temperature rise stops/ temperature falls/ mixture starts to cool as the effervescence much less vigorous (Pale pink) filtrate turns yellow-green/pale yellow/pale orange

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47	

(a) **(ii)** You are to investigate the effect of the addition of aqueous sodium hydroxide, and the addition of aqueous ammonia, to separate portions of the filtrate from **(a)(i)** and **FA 11**.

In the space below, record the details of the tests performed and the observations made in a suitable table.

	Add sodium hydroxide, dropwise with shaking, followed by in excess, to separate samples.	Add aqueous ammonia, dropwise with shaking, followed by in excess, to separate samples.
Filtrate from **(a)(ii)**	(Light) brown precipitates insoluble in excess *ppt darken (on standing)	(Light) brown precipitates insoluble in excess *ppt darken (on standing)
FA 11	Off-white/ very light brown precipitates insoluble in excess *ppt darken (on standing)	Off-white/ very light brown precipitates insoluble in excess *ppt darken (on standing)

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49	
50	

(b) Explain how you determined when the reaction is completed in **4(a)(i)**. You should support your answer by referring to your observations.

Cessation of effervescence/ bubbles (on addition of more **FA 8**)

or

temperature stops rising/ starts to fall (on addition of more **FA 8**)

51	

(c) **(i)** Consider your observations in **(a)(ii)**.

Identify the transition metal ion formed in **(a)(i)**. Justify your choice by reference to your observations in **(a)(ii)**.

ion present is Mn^{2+}

justification off-white ppt, $Mn(OH)_2$, formed with both NaOH and NH_3 which are insoluble in excess and darkens/turned darker brown on standing.

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(ii) In **(a)(i)**, the reaction between **FA 8** and **FA 10** occurs under acidic conditions.

Write a balanced equation for this reaction.

$\mathbf{Mn}O_2 + 4H^+ + C_2O_4^{2-} \rightarrow \mathbf{Mn}^{2+} + 2H_2O + 2CO_2$

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(iii) State the role of **FA 8** in the reaction and support your answer by referring to your observations.

FA 8 is an oxidising agent. Colourless, odourless gas evolved gave white ppt in limewater showed that CO_2 was evolved because $C_2O_4^{2-}$ was oxidised/ Dark grey/ black MnO_2 solid is reduced to pale pink Mn^{2+} ions.

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(d) Deduce the nature of reaction, other than the type of reaction identified in **(c)**, occurring in **(a)(i)**.

Justify by referring to your observations.

The reaction is exothermic. The temperature increased, without further heating, as **FA 8** was added.

Accept neutralisation between MnO_2 and H_2SO_4 with correct evidence and conclusion.

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END OF PAPER

9 Qualitative Analysis Notes

[ppt. = precipitate]

9(a) Reactions of aqueous cations

cation	*reaction with*	
	NaOH(aq)	NH_3(aq)
aluminium, Al^{3+}(aq)	white ppt. soluble in excess	white ppt. insoluble in excess
ammonium, NH_4^+ (aq)	ammonia produced on heating	–
barium, Ba^{2+}(aq)	no ppt. (if reagents are pure)	no ppt.
calcium, Ca^{2+}(aq)	white. ppt. with high $[Ca^{2+}(aq)]$	no ppt.
chromium(III), Cr^{3+}(aq)	grey-green ppt. soluble in excess giving dark green solution	grey-green ppt. insoluble in excess
copper(II), Cu^{2+}(aq),	pale blue ppt. insoluble in excess	blue ppt. soluble in excess giving dark blue solution
iron(II), Fe^{2+}(aq)	green ppt., turning brown on contact with air insoluble in excess	green ppt., turning brown on contact with air insoluble in excess
iron(III), Fe^{3+}(aq)	red-brown ppt. insoluble in excess	red-brown ppt. insoluble in excess
magnesium, Mg^{2+}(aq)	white ppt. insoluble in excess	white ppt. insoluble in excess
manganese(II), Mn^{2+}(aq)	off-white ppt., rapidly turning brown on contact with air insoluble in excess	off-white ppt., rapidly turning brown on contact with air insoluble in excess
zinc, Zn^{2+}(aq)	white ppt. soluble in excess	white ppt. soluble in excess

9(b) Reactions of anions

anion	*reaction*
carbonate, CO_3^{2-}	CO_2 liberated by dilute acids
chloride, Cl^-(aq)	gives white ppt. with Ag^+(aq) (soluble in NH_3(aq))
bromide, Br^-(aq)	gives pale cream ppt. with Ag^+(aq) (partially soluble in NH_3(aq))
iodide, I^-(aq)	gives yellow ppt. with Ag^+(aq) (insoluble in NH_3(aq))
nitrate, NO_3^- (aq)	NH_3 liberated on heating with OH^-(aq) and Al foil
nitrite, NO_2^- (aq)	NH_3 liberated on heating with OH^-(aq) and Al foil; NO liberated by dilute acids (colourless NO $\rightarrow$ (pale) brown NO_2 in air)
sulfate, SO_4^{2-} (aq)	gives white ppt. with Ba^{2+}(aq) (insoluble in excess dilute strong acids)
sulfite, SO_3^{2-} (aq)	SO_2 liberated with dilute acids; gives white ppt. with Ba^{2+}(aq) (soluble in dilute strong acids)

9(c) Tests for gases

gas	*test and test result*
ammonia, NH_3	turns damp red litmus paper blue
carbon dioxide, CO_2	gives a white ppt. with limewater (ppt. dissolves with excess CO_2)
chlorine, Cl_2	bleaches damp litmus paper
hydrogen, H_2	"pops" with a lighted splint
oxygen, O_2	relights a glowing splint
sulfur dioxide, SO_2	turns aqueous acidified potassium manganate(VII) from purple to colourless

9(d) Colour of halogens

halogen	*colour of element*	*colour in aqueous solution*	*colour in hexane*
chlorine, Cl_2	greenish yellow gas	pale yellow	pale yellow
bromine, Br_2	reddish brown gas / liquid	orange	orange-red
iodine, I_2	black solid / purple gas	brown	purple