Review of H2 Paper 4

1 Determination of the percentage by mass of iron in the wire

Iron wire contains impurities. In this experiment, you will investigate the percentage by mass of iron in a sample of iron wire.

A sample of iron wire is reacted with an excess of sulfuric acid to produce a solution of iron(II) sulfate.

You will titrate the solution of iron(II) sulfate with potassium manganate(VII) of known concentration to determine the amount of iron(II) ions present and hence percentage by mass of iron in the wire. You may assume that impurities do not react with potassium manganate(VII).

Iron(II) ions react with manganate(VII) ions according to the equation shown.

$$5Fe^{2+} + MnO_4^- + 8H^+ \rightarrow 5Fe^{3+} + Mn^{2+} + 4H_2O$$

FA 1 is 0.0200 mol dm^{-3} potassium manganate(VII), $KMnO_4$.

FA 2 is a diluted solution of $FeSO_4$ prepared as follows:

- 48.9 g of iron wire was reacted with sulfuric acid to make 1.00 dm^3 of solution.
- 34.00 cm^3 of the solution was then made up to 250 cm^3 with deionised water.

FA 3 is dilute sulfuric acid, H_2SO_4.

(a) (i) Procedure

1. Fill the burette with **FA 1**.
2. Pipette 25.0 cm^3 of **FA 2** into a 250 cm^3 conical flask.
3. Use a measuring cylinder to add 25 cm^3 of **FA 3** into the conical flask.
4. Add **FA 1** from the burette until the solution in the conical flask turns to a permanent **pale** pink colour.
5. Record your titration results, to an appropriate level of precision in the space on page 3.
6. Repeat steps 2 to 5 until consistent results are obtained.

Titration results

Final Burette Reading / cm^3	**22.80**	**22.80**
Initial Burette Reading / cm^3	**0.00**	**0.00**
Volume of FA1 / cm^3	**22.80**	**22.80**
	√	√

Correct headers with units **[1]**

All burette readings to 0.05 cm^3 **[1]**

Correct calculation of titre volumes i.e. final – initial burette reading and consistent results within 0.10 cm^3. **[1]**

[3]

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

Average titre volume = (22.80 + 22.80)/2 = 22.80 cm^3

Use (at least two) titre values within 0.20 cm^3 to correctly calculate average volume of **FA 1**. Working must be shown or ticks put next to the two (or more) consistent titres selected. **[1]**

Accuracy **[2]**

2 marks if difference is ≤ 0.20 cm^3

1 mark if difference is ≤ 0.40 cm^3

0 mark if difference is > 0.40 cm^3

volume of **FA 1** = ... [3]

(b) **(i)** Calculate the amount of iron(II) ions present in 25.0 cm^3 of **FA 2**.

Amount of MnO_4^- used = $\frac{(22.80)(0.0200)}{(1000)}$ = 4.56×10^{-4} mol

Amount of iron(II) ions present = $4.56 \times 10^{-4} \times 5 = 2.28 \times 10^{-3}$ mol

[1] $\frac{ans\ in\ (a)(ii)}{1000} \times 0.0200 \times 5$

amount of iron(II) ions = ... [1]

(ii) Calculate the mass of iron present in 25.0 cm^3 of **FA 2**.

[A_r: Fe, 55.8]

Mass of iron present = $2.28 \times 10^{-3} \times 55.8 = 0.127$ g

[1] ans in (b)(i) × 55.8

mass of iron = ... [1]

(iii) Calculate the percentage by mass of iron in the sample of iron wire.

Mass of iron in wire sample = $\frac{0.127}{25.0} \times \frac{250}{34.00} \times 1000 = 37.4$ g

% by mass of iron in wire sample = $\frac{37.4}{48.9} \times 100\% = 76.5\%$

[1] Mass of iron in the wire sample = $\frac{ans\ in\ (b)(ii)}{25.0} \times \frac{250}{34.00} \times 1000$

[1] % by mass of iron in the wire sample = $\frac{mass\ of\ iron\ in\ the\ wire}{48.9} \times 100\%$

percentage by mass of iron = ... [2]

(c) A student suggested that when a piece of iron wire was dissolved in a known amount of excess sulfuric acid, the amount of iron that reacted with the acid could be determined by titrating the mixture containing the remaining acid with a known concentration of sodium hydroxide.

Explain whether the student was correct.

Student was incorrect as Fe^{2+} in the mixture can react with NaOH(aq) to form $Fe(OH)_2$. **[1]**. Hence, volume of NaOH(aq) is larger than expected. **[1]**

[Total: 12]

2 Determination of enthalpy change for the reaction between sodium carbonate with water and carbon dioxide

You will determine the enthalpy change, ΔH_r, for the reaction of sodium carbonate, Na_2CO_3, with water and carbon dioxide to form sodium hydrogencarbonate, $NaHCO_3$, via reactions 1 and 2.

$$Na_2CO_3(s) + H_2O(l) + CO_2(g) \longrightarrow 2NaHCO_3(s) \qquad \Delta H_r$$

In reaction 1, you will react sodium hydrogencarbonate with dilute hydrochloric acid and find the temperature change.

$$NaHCO_3(s) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l) + CO_2(g) \qquad \Delta H_1$$

In reaction 2, the reaction of sodium carbonate with dilute hydrochloric acid will be considered.

$$Na_2CO_3(s) + 2HCl(aq) \longrightarrow 2NaCl(aq) + H_2O(l) + CO_2(g) \quad \Delta H_2$$

You will then use the results to calculate the enthalpy change, ΔH_r.

FA 4 is sodium hydrogencarbonate, $NaHCO_3$.

FA 5 is 2.0 mol dm^{-3} hydrochloric acid, HCl.

Handle all chemicals with care and follow all safely precautions at all times.

(a) Reaction 1: Reaction between sodium hydrogencarbonate and dilute hydrochloric acid

Read through the method and prepare two suitable tables in the space provided on page 6 to record:

- all weighings to an appropriate level of precision
- all values of temperature, to an appropriate level of precision
- all values of time, recorded to the nearest minute.

It is important that you measure each temperature at the specified time.

1. Weigh the container with **FA 4** and record the balance reading in your table on page 6.
2. Place the polystyrene cup in the 250 cm^3 beaker.
3. Use the 25 cm^3 measuring cylinder to transfer 25 cm^3 of the acid, **FA 5**, into the polystyrene cup. The acid is in excess.
4. Place the thermometer in the acid and record the initial temperature in your second table on page 6. Tilt the cup if necessary, so that the bulb of the thermometer is fully covered. This is the temperature at t = 0 minute. Start timing.
5. Record the temperature of the acid at 1 minute and at 2 minutes.
6. At time t = 2½ minutes, **carefully** tip all the **FA 4**, in **small portions** to avoid spray, into the acid and stir to dissolve it. ***Avoid inhalation of the acid vapour***.
7. Record the temperature of the solution from t = 3 to 8 minutes at 1 minute intervals.
8. Reweigh the container with any residual **FA 4**. Record the balance reading and the mass of **FA 4** used.

Mass of weighing bottle and FA4 / g	**8.76**
Mass of weighing bottle and residue FA4 / g	**6.33**
Mass of FA4 used / g	**2.43**

time / min	**0.0**	**1.0**	**2.0**	**2.5**	**3.0**	**4.0**	**5.0**	**6.0**	**7.0**	**8.0**
temp / ºC	**33.5**	**33.5**	**33.5**		**27.0**	**27.5**	**28.0**	**28.0**	**28.5**	**28.5**

[✓] mass table with correct headers and units with reweighing

[✓] headers and units for temperature and time. accept if 2.5 min not indicated in table

[✓] mass reading in 2 or 3 dp

[✓] precision of temp to 1 dp (ignore dp for time)

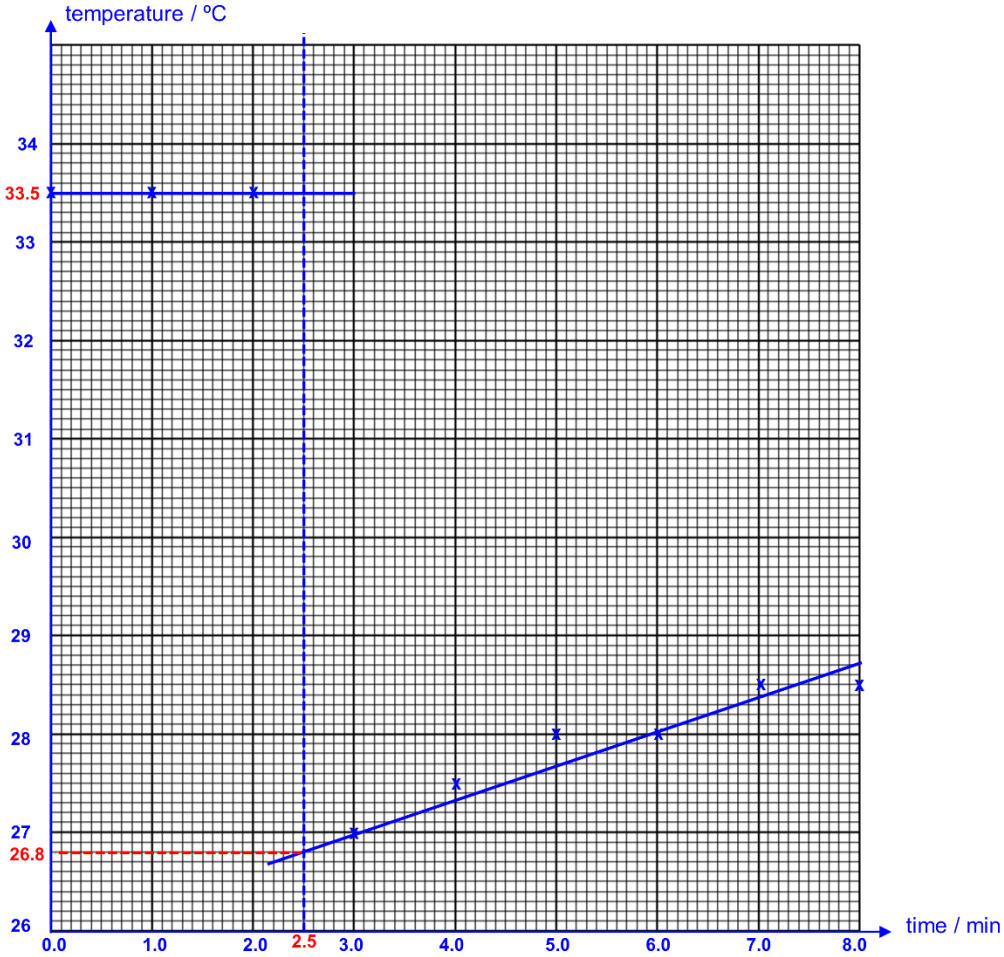
2[✓] = [1]

[1] 3 mass readings + 9 temperature readings but penalise if have temperature reading at 2.5 min + correct trend for endothermic reaction

(b) **(i)** Plot a graph of temperature on the y-axis against time on x-axis on the grid below.

You will use the graph to determine the theoretical temperature change at 2½ minutes.

The scale for temperature should extend at least 1 ºC below your lowest recorded temperature.



Draw two straight lines of best fit on your graph, one for the temperature of the acid before adding **FA 4** and the other for the warming of the solution once the reaction is complete.

Extrapolate the two lines to 2½ minutes. [3]

[1] axis label and scale chosen such that plotted points occupy more than half the available graph area in both x and y directions.

[1] all points plotted to within half a small square

[1] two lines of best fit drawn – one for before adding and the other for warming of solution and both lines extrapolated to 2½ minutes + min temp must be lower than minimum data point

(ii) Determine the change in temperature at 2½ minutes. [1]

ΔT = 26.8 – 33.5 = **– 6.7** °C *(record to half smallest division based on graph)*

[1] correct reading of both initial and lowest temperatures and ΔT calculated

(iii) Use your answer to **(b)(ii)** to calculate the heat energy absorbed when **FA 4** was added to the acid in **reaction 1**. [1]

(Assume that 4.3 J of heat energy changes the temperature of 1.0 cm^3 of the mixture by 1.0 ºC.)

Heat absorbed = $mc\Delta T$

= 25 x 4.3 x 6.7 = **720 J** **[1]**

(iv) Calculate the enthalpy change, ΔH_1, in kJ mol^{-1}, when 1 mole of **FA 4**, $NaHCO_3$, reacts with acid. [1]

[A_r: H, 1.0; C, 12.0; O, 16.0; Na, 23.0]

Amount of $NaHCO_3$ = 2.43 / 84 = 0.0289 mol

ΔH_1 = + 720 / 0.0289 = **+24.9** kJ mol^{-1} **[1]**

(c) Reaction 2: Reaction between sodium carbonate and dilute hydrochloric acid

A student repeats the experiment by adding 2.01 g of sodium carbonate, Na_2CO_3, to
25 cm^3 of dilute hydrochloric acid, HC*l*. The temperature increases from 34.0 °C to 39.0 °C.

(i) Using the above results, calculate the heat energy produced when Na_2CO_3 was added to the acid in **reaction 2**. [1]

(Assume that 4.3 J of heat energy changes the temperature of 1.0 cm^3 of solution by 1.0 ºC.)

Heat produced = $mc\Delta T$

= 25 x 4.3 x 5.0 = **538 J** **[1]**

(ii) Calculate the enthalpy change, ΔH_2, in kJ mol^{-1}, when 1 mole of Na_2CO_3, reacts with acid. [1]

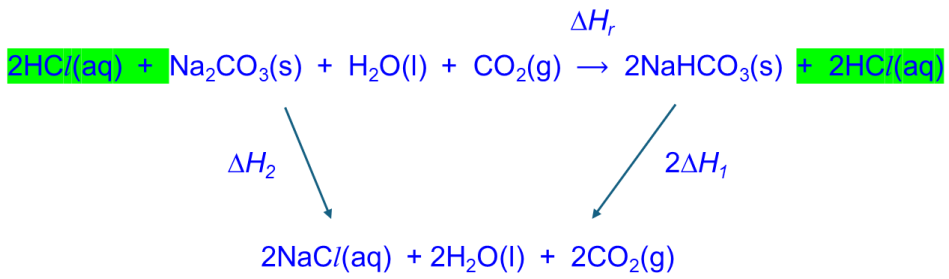
[A_r: C, 12.0; O, 16.0; Na, 23.0]

Amount of Na_2CO_3 = 2.01 / 106 = 0.0190 mol

ΔH_2 = – 538/ (0.01896 x 1000) = **–28.3** kJ mol^{-1} **[1]**

(iii) Use your answers to parts **(b)(iv)** and **(c)(ii)** and the equations for **reactions 1** and **2** to determine the enthalpy change, ΔH_r, for the reaction below. [4]

$$Na_2CO_3(s) + H_2O(l) + CO_2(g) \longrightarrow 2NaHCO_3(s) \qquad \Delta H_r$$



[1] $\Delta H_r = \Delta H_2 - 2\Delta H_1$

$\Delta H_r = -28.3 - 2\Delta H_1 = = -28.3 - 2(24.9) =$ **-78.1 kJ mol^{-1}**

[1] for correct value with sign for ΔH_r

[1] statements for 1(a)(ii), 1(b), 2(b)(iii), 2(b)(iv), 2(c)

[1] all final answers in appropriate s.f., and units for 1(a)(ii), 1(b), 2(b)(iii), 2(b)(iv), 2(c)

(d) **(i)** A student stated that the temperature change determined in **reaction 2** could be made more accurate by using twice the volume of acid in **reaction 2**. Suggest whether the student is correct or incorrect and justify your answer.

[1] The student is **incorrect** because the **acid** is **already in excess**.

[1] **Same amount of heat generated** but now distributed across a **greater volume**. This will result in a **smaller ΔT** and hence a **greater percentage error**

(ii) By considering the calculation of the enthalpy change in part **(c)(ii)**, show that the temperature rise per gram of carbonate used, $\frac{\Delta T}{\text{mass of carbonate}}$, is a constant, when 25 cm^3 of excess dilute HCl is used.

$$\Delta H = -\frac{mc\Delta T}{\text{amount of carbonate}} = -\frac{mc\Delta T}{\text{mass of carbonate / Mr of carbonate}}$$

$$\Delta T = -\frac{\Delta H}{mc(M_r \text{ of carbonate})} \times (\text{mass of carbonate})$$

$= k \times$ mass of carbonate

Hence $\frac{\Delta T}{\text{mass of carbonate}} = k$ **[1]**

(iii) Another student carried out **reaction 2** twice using the carbonate of a different metal and obtained the following results.

	mass of carbonate / g	temperature rise, ΔT / ºC
first result	2.96	4.0
second result	3.65	5.0

Hence, verify by showing appropriate calculations, whether the results obtained are consistent.

$\frac{\Delta T}{\text{mass of carbonate}} = k$

For **first** result, $\frac{\Delta T}{\text{mass of carbonate}}$ = **1.35**;

for **second** result, $\frac{\Delta T}{\text{mass of carbonate}}$ = **1.37** **[1]**

Both values are **close in value** / **approximately constant** / have **small difference**, the results are **consistent**. **[1]**

[Total: 20]

3 Qualitative analysis

At each stage of any test you are to record details of the following:

- colour changes seen
- the formation of any precipitate
- the solubility of such precipitates in an excess of the reagent added.

Where gases are released they should be identified by a test, described in the appropriate place in your observations.

You should indicate clearly at what stage in a test a change occurs.

No additional tests for ions present should be attempted.

Rinse and reuse test-tubes where possible.

Where reagents are selected for use in a test, the name or correct formula of the element or compound must be given.

(a) A sample of limestone was reacted with dilute nitric acid to give solution **FA 6**.

This sample of limestone contained calcium carbonate, $CaCO_3$, and one other salt.

This additional salt contains a single cation and a single anion from those listed on pages 19 and 20. By carrying out the following tests you will be able to suggest identities of the additional ions.

2[✓] = 1 mark

test	observations
(i) To 1 cm depth of **FA 6** in a test-tube, add aqueous ammonia then	[✓] White ppt insoluble in excess $NH_3(aq)$
add a spatula full of ammonium chloride.	[✓] Ppt soluble in ammonium chloride
(ii) To 1 cm depth of **FA 6** in a test-tube, add 1 cm depth of aqueous silver nitrate.	[✓] No ppt
(iii) To 1 cm depth of **FA 6** in a test-tube, add 1 cm depth of aqueous barium nitrate.	[✓] White ppt

[2]

(iv) Identify the anion present in **FA 6**.

[1]

SO_4^{2-} **[1]**

(v) Select a reagent to use in a further test on **FA 6** to confirm the identity of the cation in **FA 6**. Carry out your test and complete the table below.

2[✓] = 1 mark

test	*observations*	*conclusion*
To a 1 cm depth of **FA 6** in a test-tube, add [✓] NaOH(aq)	[✓] White ppt insoluble in excess NaOH(aq)	Mg^{2+} present **[1]**

[2]

(vi) Using your knowledge of Chemical Equilibria and with the aid of equations, explain your observations in part **(a)(i)**.

[2]

[✓] white ppt of $Mg(OH)_2$ is precipitated on addition of aq. ammonia.

$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$ (1) [✓]
$Mg^{2+} + 2OH^- \rightleftharpoons Mg(OH)_2$ (2) [✓]
On adding NH_4Cl, the increase in concentration of NH_4^+ causes equilibrium position of (1) to shift left. The concentration of OH^- decreases and equilibrium position of (2) shifts left and $Mg(OH)_2$ dissolves. [✓]

2[✓] = 1 mark

(b) Organic analysis

In this question, you will deduce the structure of an organic compound, **FA 7**.

FA 7 has the molecular formula $C_4H_8O_2$ with two functional groups present.

Do not carry out the tests for which observations have been recorded.

When organic solutions are to be heated, a hot water bath should be used.

Table 3.1

	tests	observations
(i)	Place about 2 cm depth of **FA 3** in a test-tube. To this test-tube, add 15 drops of **FA 7**, followed by 1 drop of **FA 1**. Warm the mixture in the hot water bath for two minutes.	**[1]** purple $KMnO_4$ decolourised

(ii)	Add 1 cm depth of aqueous silver nitrate to a boiling tube. Then slowly add 1 cm depth of aqueous sodium hydroxide. Add aqueous ammonia slowly, with shaking, until the precipitate just dissolves. You may use a clean glass rod to stir the mixture and help dissolve the precipitate. Add 2 cm depth of **FA 7** to this mixture and shake the boiling tube. Place the boiling tube in the test-tube rack and leave it for 3 minutes.	 brown (or grey) ppt. formed ppt. dissolves no silver mirror
(iii)	Place about 1 cm depth of **FA 7** in a test-tube. To this test-tube, add 2,4–dinitrophenylhydrazine dropwise.	 orange ppt. formed
(iv)	Place about 15 drops of **FA 7** and add 10 drops of aqueous sodium hydroxide in a test-tube. Now add iodine solution dropwise, until a permanent orange/red colour is obtained. Warm the mixture in the hot water bath for two minutes.	 **[1]** pale yellow ppt. formed Accept: yellow ppt
(v)	Place 1 cm depth of **FA 7** in a test-tube. To this test-tube, cautiously add a small piece of sodium metal.	 Effervescence observed. H_2 gas produced extinguishes a lighted splint with a “pop” sound.

[2]

(c) **(i)** Based on the observations in Table 3.1, identify the two functional groups in **FA 7**. Explain your answer.

[2]

Functional group: ✓ ketone

Explanation:
✓ Since FA 7 undergoes condensation with 2,4-DNPH in (b)(iii), FA 7 contains either a ketone or aldehyde (or carbonyl compound).

However, since FA 7 gives a negative test with Tollens' reagent in (b)(ii), it does not contain an aldehyde.

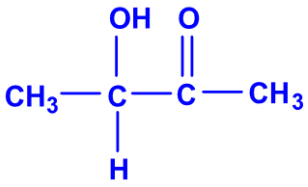
Functional group: ✓ alcohol

Evidence:

✓ Since FA 7 undergoes redox reaction with Na in (b)(v) to give H_2, it contains an alcohol or -OH functional group.

2✓ : 1m

(ii) Suggest a possible structure of **FA 7** that is consistent with all the observations in Table 3.1.



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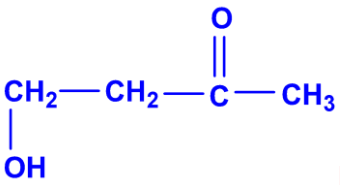

[1]

[1]

[Total: 12]

4 Planning

Calcium iodate(V), $Ca(IO_3)_2$, has low solubility in water. When a sample of this salt is mixed with water, a small amount of solid dissolves.

The amount of iodate(V) ions, IO_3^-, can be determined as described below.

Excess potassium iodide is added to an acidified solution containing IO_3^- ions. Iodine is liberated, as shown in equation 1.

equation 1 $\quad IO_3^-(aq) + 5I^-(aq) + 6H^+(aq) \rightarrow 3I_2(aq) + 3H_2O(l)$

The liberated iodine is then titrated with a standard solution of sodium thiosulfate, as shown in equation 2.

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

The amount of iodate(V) ions found can then be used to determine the solubility product, K_{sp}, of calcium iodate(V).

(a) Write an expression for the solubility product, K_{sp}, of calcium iodate(V). Include units in your answer.

$K_{sp} = [Ca^{2+}][IO_3^-]^2$ units: $mol^3\ dm^{-9}$ **[1]**

[1]

(b) Plan an experiment to determine the solubility product, K_{sp}, of $Ca(IO_3)_2(s)$, at 40 °C.

You may assume that you are provided with:

- deionised water
- solid calcium iodate(V), $Ca(IO_3)_2$
- 1.00 mol dm^{-3} aqueous potassium iodide, KI
- 1.00 mol dm^{-3} dilute sulfuric acid, H_2SO_4
- 0.100 mol dm^{-3} sodium thiosulfate, $Na_2S_2O_3$
- starch indicator
- the equipment normally found in a school or college laboratory.

In your plan, you should include brief details of:

- the apparatus you would use
- the quantities you would use
- the procedure you would follow
- the measurements you would make
- how you would ensure that an accurate and reliable value of K_{sp} is obtained.

[6]

Preparation of saturated solution of $Ca(IO_3)_2$

1. Use a 100 cm^3 measuring cylinder to transfer 100 cm^3 of deionised water to a clean and dry 250 cm^3 beaker.

2. Immerse the beaker in a thermostatically controlled water bath maintained at 40 ^oC.

3. When the temperature is 40 ^oC, add solid $Ca(IO_3)_2$ using a spatula into the beaker. Stir the contents in the beaker using a glass rod until no more solid can dissolve.

4. Leave the mixture to stand for several minutes to allow equilibrium to be reached.

5. Filter the mixture through a dry filter paper and dry filter funnel into a dry conical flask to collect the filtrate for titration.

Titration of the filtrate

6. Fill a burette with 0.100 mol dm^{-3} sodium thiosulfate, $Na_2S_2O_3$. Record the initial burette reading.

7. Pipette 25.0 cm^3 of the filtrate into a 250 cm^3 conical flask.

8. Use a 10 cm^3 measuring cylinder to add about 10 cm^3 of excess KI(aq) to the conical flask.

9. Use another 10 cm^3 measuring cylinder to add about 10 cm^3 of excess H_2SO_4(aq) to the conical flask.

10. Add $Na_2S_2O_3$ from the burette into the conical flask until a pale yellow solution is obtained.

11. Add 1 cm^3 / 10 drops of starch solution to the conical flask. Continue adding $Na_2S_2O_3$ until the solution turns from blue-black to colourless. Record the final burette reading.

12. Repeat steps 7-11 until at least two consistent titre results within 0.10 cm^3 are obtained.

M1: coherent plan i.e preparation of saturated solution and correct titration of iodine generated.

M2: correct preparation of saturated solution: appropriate volume of DI water (more than 50 cm^3 for sufficient titrations), add solid and stir until no more solid can dissolve, allow solution to stand/ attain equilibrium

M3: maintain constant temperature at 40 ^0C using a thermostatically controlled water bath (do not accept Bunsen burner).

M4: filtration using **dry** apparatus + apparatus of filter paper, filter funnel and conical flask (do not accept beaker instead of conical flask)

M5: Add *excess* KI (no need actual values) and H_2SO_4 to the filtrate (infer from excess or look at the reasonable quantities stated) + use of measuring cylinders

M6: titration with end-point colour change of blue-black to colourless + starch indicator used (Not penalised if didn't state number of drops of starch. Penalise if starch is added at the start of titration.)

M7: at least two consistent results with criteria of within 0.10 cm^3 specified

7 max 6 marks

(c) A student claims that the K_{sp} of calcium iodate(V) will change if $Ca(IO_3)_2$ solid is dissolved in a solution of KIO_3 instead of water at the same temperature.

State and explain whether you agree with the student's claim.

Do not agree. K_{sp} only changes with temperature. **[1]**

(d) Given that the average titre value obtained from the titrations is **V** cm^3, determine in terms of **V**, the solubility product, K_{sp}, of $Ca(IO_3)_2(s)$ at 40 ºC. Show your working clearly.

[1] amount of iodate = $\frac{V}{1000} \times 0.100 \times \frac{1}{6}$

[1] amount of Ca^{2+} = $\frac{V}{1000} \times 0.100 \times \frac{1}{6} \times \frac{1}{2}$

[1] divide both amount by 0.025, ecf K_{sp} expression in **(a)**

Final expression will be

$K_{sp} = \frac{1.6 \times 10^{-5} \times V^2}{36} \times \frac{2 \times 10^{-3} \times V}{6} = (4.44 \times 10^{-7} \times V^2)\ (3.33 \times 10^{-4} \times V)$

Note : calculation must be consistent with plan in (b).

[3]

[Total: 11]

Qualitative Analysis Notes

[ppt. = precipitate]

(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

(b) Reactions of aqueous 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)

(c) Test 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

(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