Suggested solutions

DUNMAN HIGH SCHOOL
Preliminary Examinations 2018
Year 6

H2 CHEMISTRY — 9729/04

Paper 4 Practical

28 August 2018

2 hour 30 minutes

INSTRUCTIONS TO CANDIDATES

1 Write your **name**, **index number** and **class** on this cover page.
2 Give details of the practical shift and laboratory where appropriate, in the boxes provided.
3 Write in dark blue or black pen.
4 You may use an HB pencil for any diagrams or graphs.
5 Do not use staples, paper clips, 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.
Qualitative Analysis Notes are printed on pages 17 and 18.

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

Shift
Laboratory

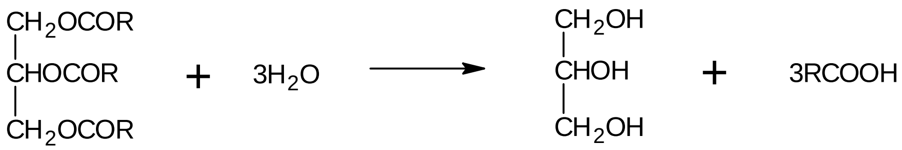
For Examiner's Use	
Question No.	**Marks**
1	**14**
2	**13**
3	**28**
Total	**55**

This question paper consists of **18** printed pages and **0** blank page.

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

1 Investigation of reaction kinetics between triacetin and lipase

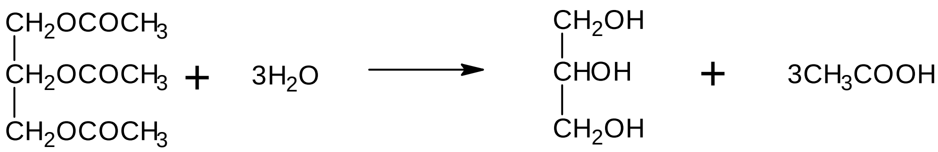
Lipase is an enzyme that breaks down, hydrolyses, the lipids present in animal fats and vegetable oils. It does so by hydrolysing ester bonds. A general reaction scheme for this process is shown below; the size of the R group and its level of saturation will depend on the lipid used.



glycerol fatty acid

Olive oil is often used as a substrate in testing for lipase, with the oleic acid released being titrated with sodium hydroxide solution. Oleic acid, $C_{17}H_{33}COOH$, is a mono–unsaturated fatty acid.

In this experiment, glycerol triethanoate (triacetin) is used as the substrate. Lipase breaks this substrate down into glycerol and ethanoic acid as shown below.



glycerol triethanoate
(triacetin)

The ethanoic acid produced lowers the pH of the solution. This change can be detected by using a suitable acid–base indicator, bromocresol purple which changes from purple at high pH value to yellow at lower pH value. Sodium carbonate is added to the triacetin to ensure that the pH is high enough at the start of the experiment. Only after the sodium carbonate has all reacted with the liberated ethanoic acid will the ethanoic acid concentration rise and the pH value decreases. The reaction time is measured from the initial point of mixing the solutions until the end–point of the indicator is reached.

The following reagents are provided for this experiment.

FA 1 is triacetin
FA 2 is 0.5% sodium carbonate solution
FA 3 is 8% lipase solution
FA 4 is Bromocresol purple solution

Procedure

1. To a test-tube, add 2 drops of **FA 2** followed by 1 drop of **FA 4**.

2. Using a dropper, add 5 drops of **FA 1** to the test tube.

3. Using a 250 cm^3 beaker, collect 200 cm^3 of warm water from the dispenser.

4. **Swirl the bottle** containing **FA 3** before using a syringe to transfer 0.20 cm^3 of **FA 3** into the test tube. Start the stopwatch immediately and place the test tube into the warm water bath.

5. **Shake the test tube continuously**. This is necessary to ensure that **FA 1** and **FA 3** are continually mixed.

6. Record the reaction time, ***t*** (in seconds) when the mixture turns yellow. This will be a gradual process.

7. Repeat steps 1 – 6 for four other test tubes, using different volumes of **FA 3** as shown in the table below.

Test Tube	Volume of **FA 3** / cm^3
1	0.20
2	0.40
3	0.60
4	0.80
5	1.00

8. Calculate the value of reaction rate, $\frac{\mathbf{1}}{\boldsymbol{t}}$, for each experiment.

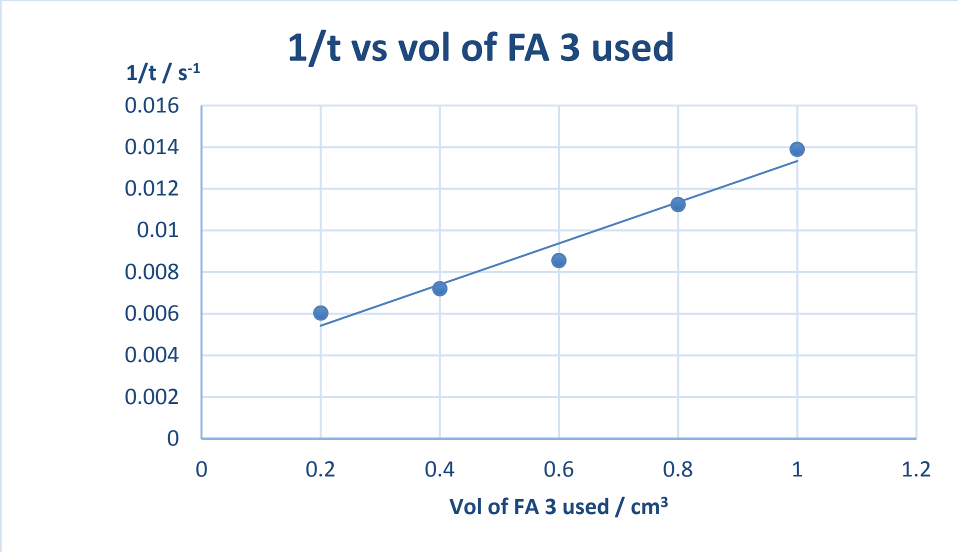
(a) Results:

(i)

Test Tube	Reaction time ‘t’ / s	1/t / s^{-1}
1	166	0.00602
2	139	0.00719
3	117	0.00854
4	89	0.0112
5	72	0.0138

[4]

(ii) Plot the reaction rate, $\frac{1}{t}$, against the volume of **FA 3** used.



[3]

(iii) By considering the graph in **(a)(ii)**, deduce the order of reaction with respect to lipase, **FA 3**. Explain your deduction.

[3]

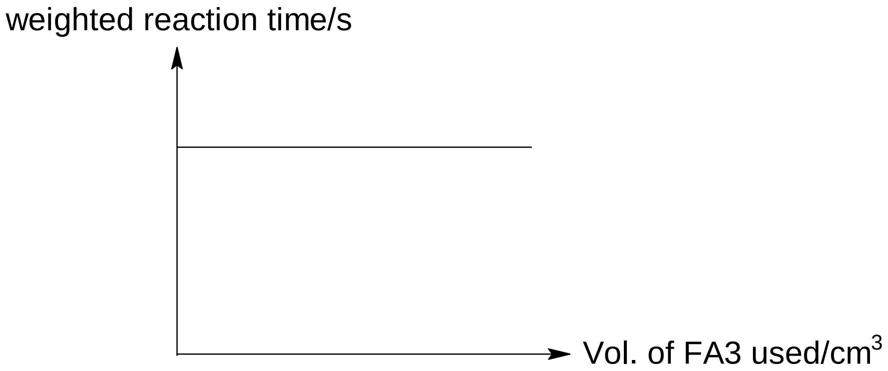
A best fit <u>linear graph</u> was obtained. This suggest that the relative rate as represented by 1/t **<u>increases proportionately</u>** with increasing volume of lipase used. The increasing volume of lipase used can be treated as increasing concentration since volume of all other reagents are maintained the same.
Hence, the **<u>order of reaction with respect to lipase is 1</u>**.

(iv) The **weighted reaction time** can be obtained by multiplying the appropriate conversion factor to the reaction time, ***t*** (in seconds).

Test Tube	Conversion factor	Weighted reaction time / s
1	0.20	
2	0.40	
3	0.60	
4	0.80	
5	1.00	

Based on your deduction in **(a)(iii)**, sketch the expected graph of **weighted reaction time, in seconds,** against the volume of **FA 3** used. Explain your sketch.

[2]



Since the **weighted factor is the same magnitude as the volume of FA3 used** and that **order of reaction with respect to FA3 is one**, a straight line with zero gradient is expected.

(b) Explain how the volumes of **FA 1**, **FA 2** and **FA 3** should change in the procedure in order to investigate the relationship between substrate concentration and the rate of reaction. You are **not** required to give exact figures.

[2]

The **volume of FA 1 (triacetin) must change with each experiment** eg it should increase from test tube 1 to 5 while the **volume of FA 3 (lipase) should be kept constant for all 5 test tubes**. The **volume of FA 2 (sodium carbonate) solution should be kept constant for all 5 test tubes**. Any changes in rate of reaction could then be attributed to the change in substrate concentration.

[Total: 14]

2 Planning

An ideal gas is a theoretical idea – a gas in which there are no attractive forces between the molecules, and in which the molecules take up no space. The ideal gas law is usually stated as $pV = nRT$, where

- p is the pressure
- V is the volume
- n is the number of moles
- T is the temperature in K

of the gas, and

- R is the ideal gas constant.

The value of R is determined experimentally by measuring the other variables in the equation, and solving mathematically to get the value of the constant. R is the same for all gases – provided the gas has ideal behaviour.

In this experiment you will determine the ideal gas constant using H_2 gas. The H_2 will be generated using the reaction between magnesium and hydrochloric acid.

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

You may assume you are provided with:

- a two-necked round bottom flask,
- a rubber tubing,
- a water tub,
- 0.1 mol dm^{-3} hydrochloric acid,
- thermometer,
- barometer (instrument used to measure pressure),
- one piece of magnesium ribbon strip of about 0.06 g,
- apparatus normally found in a school laboratory

(a) Your plan should contain the following:

- the procedure that you would follow and the measurements that you would take
- a tabulation of the experimental data to be collected
- the reactants and conditions that you would use
- the mass of magnesium that you would use and the volume of hydrochloric acid that you would measure. Assume that the molar volume of gas under experimental conditions is 24.0 dm^3 mol^{-1}.
  [A_r: Mg, 24.3]

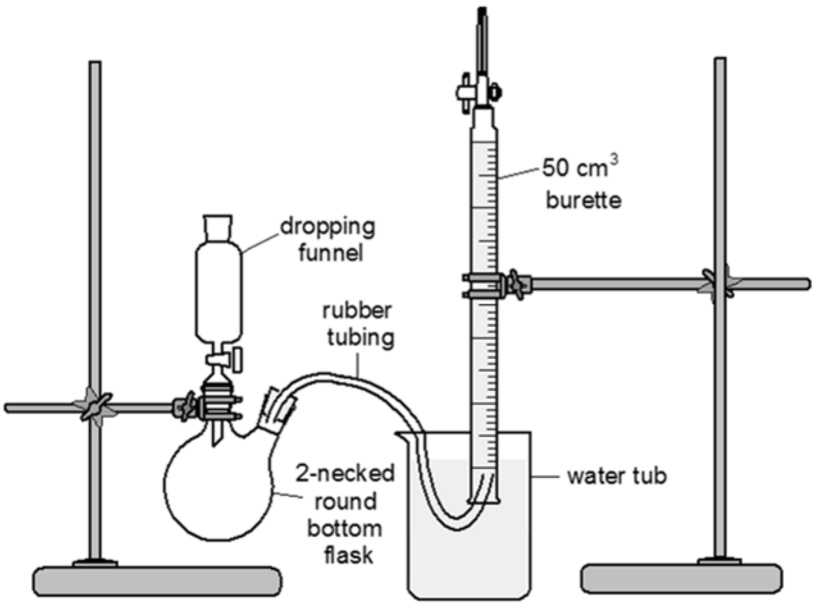
[7]

Pre-Calculations:

If 0.06 g of Mg was used,
No. of moles of HC*l* required = $\frac{0.06}{24.3} \times 2 = 4.938 \times 10^{-3}$ mol
Volume of HC*l* required = 49.38 cm^3 (volume of HC*l* used must be greater than 49.4 cm^3)
Volume of H_2 evolved = $2.469 \times 10^{-3} \times 24000 = 59.3$ cm^3 (assuming r.t.p.)

Maximum mass of Mg allowed to evolve 50 cm^3 of H_2:
No. of moles of Mg = $\frac{0.05}{24} = 2.083 \times 10^{-3}$ mol
Mass of Mg = $2.083 \times 10^{-3} \times 24.3 = 0.0506$ g
Minimum volume of HC*l* required = $\frac{2.083 \times 10^{-3} \times 2}{0.1} = 0.0417$ dm^3 = 41.7 cm^3

Procedure:



1) Cut the magnesium strip to three roughly equal pieces. **Using a weighing balance, weigh each piece of magnesium strip in a weighing bottle**. Each strip should weigh about 0.02 g.

2) Place one piece of cut **magnesium strip into the two-necked round bottom flask.**

3) Secure the two-necked round bottom flask with a retort stand.

4) Using a 50 cm^3 measuring cylinder, **measure out 25 cm^3 of 0.1 mol dm^{-3} hydrochloric acid**. Transfer the 25 cm^3 of 0.1 mol dm^{-3} hydrochloric acid to a dropping funnel.

5) **Connect the dropping funnel** to one of the necks of the two-necked round bottom flask.

6) Fill the water tub with tap water.

7) **Submerge the mouth of an inverted 50.00 cm^3 burette (filled with water) in the water tub** and secure the burette with a second retort stand. Adjust the water level in the burette until it is between 48.00 cm^3 and 50.00 cm^3. Note and record the **initial reading of the burette**.

8) **Connect one end of a rubber tubing** (with stopper) to the other neck of the two-necked round bottom flask and place the other end inside the inverted burette.

9) Measure and record the **ambient temperature using the thermometer**.

10) Measure and record the **pressure of the room using the barometer**.

11) Open the tap of the dropping funnel fully and gently swirl the two-necked round bottom flask. Close the tap immediately when all the hydrochloric acid is added.

12) **Note the final reading of the burette when all effervescence has ceased**.

13) **Repeat Steps 2 to 12 with the other two pieces of magnesium.**

Tabulation of data:

Experiment	Mass of empty weighing bottle /g	Mass of weighing bottle and Mg strip /g	Mass of Mg strip used / g	Final burette reading / cm^3	Initial burette reading / cm^3	Volume of H_2 produced / cm^3	Temperature / °C	Barometer reading / kPa
1								
2								
3								

(b) State two basic assumptions of an ideal gas.

[2]

The gas particles have **<u>negligible volume compared to the volume of the container</u>**.

There are **<u>no intermolecular forces of attraction</u>** between gas particles.

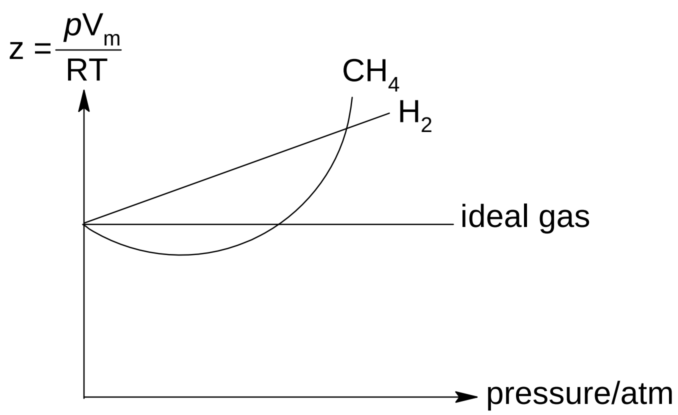
The collisions which occur between gas particles and with walls of the container are **<u>perfectly elastic</u>**.

The gas particles are in **<u>constant random motion</u>.**

(c) Real gases do not obey the ideal gas law exactly. Sketch a **fully labelled** graph representing the deviation from ideality. Your sketch should include:

- ideal gas
- H_2 gas
- CH_4 gas

[4]



[Total: 13]

3 Investigation of chemical reactions of electrolytes

You are provided with the following.
FA 5 is sodium chloride solution.
FA 6 and **FA 9** are aqueous electrolytes, each containing one cation and one anion.

(a) Electrolysis of FA 5

1. Set up the apparatus and circuit as shown in Fig. 3.1.

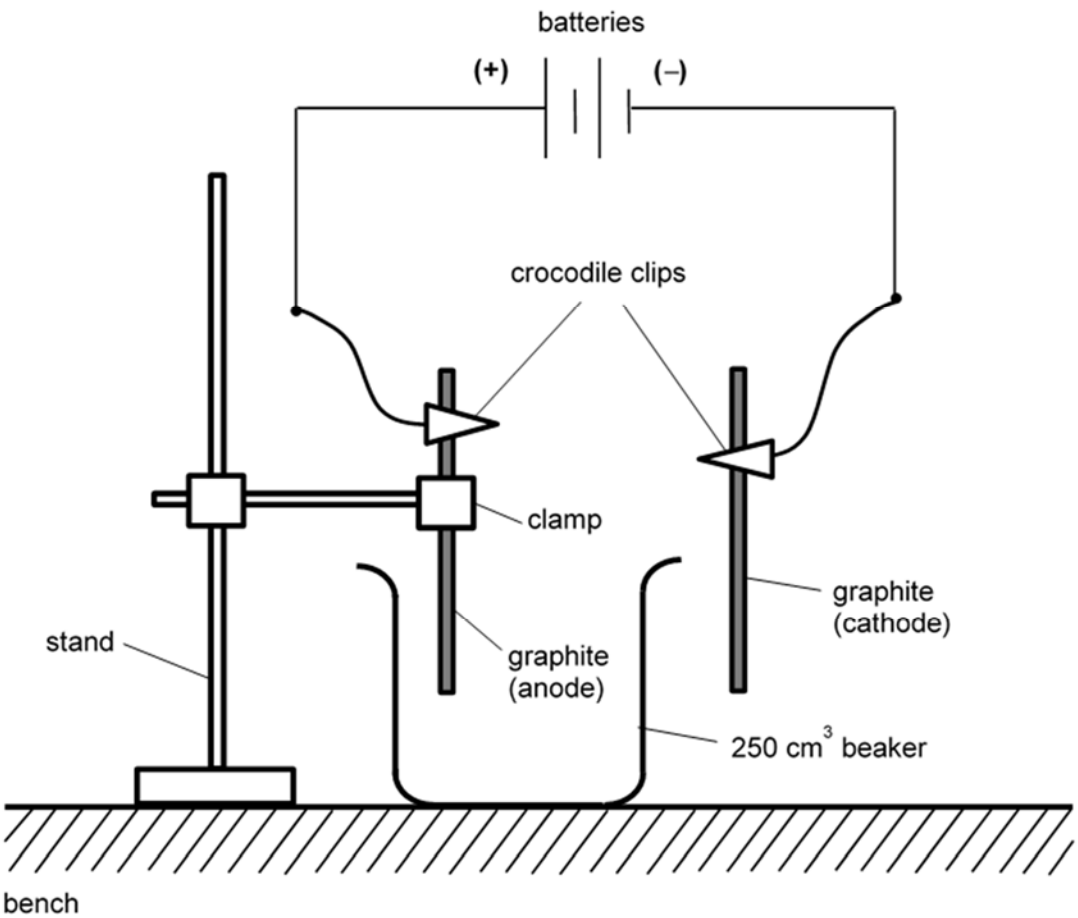


Fig 3.1

2. Transfer about 200 cm^3 of **FA 5** into the glass beaker.
3. Ensure that the graphite (anode) is partially submerged in **FA 5** by adjusting the height of the clamp.
4. Immerse one end of the graphite (cathode) in **FA 5** for a few minutes and record your observations in Table 3.1. You do **not** need to collect and test any gases evolved.

Table 3.1

electrolyte	observations at the electrodes
FA 5	anode: **effervescence / bubbles** of greenish yellow, pungent gas observed
	cathode: vigorous **effervescence /** more **bubbles** of colourless, odourless gas observed

[1]

5. Perform the tests described in Table 3.2, and record your observations in the table.

Table 3.2

tests		observations
1.	Add one drop of Universal Indicator solution to **FA 5** near the **cathode**.	**Green** Universal Indicator solution turned **purple/violet**.
2.	Place a piece of moist blue litmus paper near the **anode**. Observe until no further change is seen.	**Gas evolved** turned moist blue litmus paper red and then **bleached** it.

[2]

6. Remove both electrodes from **FA 5** and rinse them thoroughly with deionised water.

Consider your observations in Table 3.2.

(i) Explain, with the aid of an equation, the observations made in test 1 in Table 3.2. [2]

$$\underline{\mathbf{2H_2O + 2e^- \rightarrow H_2 + 2OH^-}}$$

Water is preferentially reduced at the cathode, producing hydrogen gas and **hydroxide ions** which make the solution near the cathode **basic**.

(ii) With reference to the data below, suggest an explanation for your observations in test 2 in Table 3.2, in terms of *E* values.

electrode reaction	$E^\ominus$ / V
$Cl_2 + 2e^- \rightleftharpoons 2Cl^-$	+1.36
$O_2 + 4H^+ + 4e^- \rightleftharpoons 2H_2O$	+1.23

The observations from test 2 of Table 3.2 suggest that **chlorine gas is produced at the anode**.

Thus **FA 5** could be concentrated sodium chloride solution with a **high concentration of chloride ions** (> 1 mol dm^{-3}) so that by **Le Chatelier's Principle, the position of the $Cl_2 + 2e^- \rightleftharpoons 2Cl^-$ equilibrium will shift to the left** to favour the backward oxidation reaction.

Then **$E(Cl_2/Cl^-)$ becomes less positive than $E^\ominus(O_2/H_2O)$ = +1.23 V** and chloride ion is preferentially oxidised at the anode.

[3]

(b) (i) Qualitative analysis of FA 6

FA 7 is a solution of potassium sodium tartrate tetrahydrate in sodium hydroxide.
FA 8 is an organic compound with only one functional group.

Perform the tests described in Table 3.3, and record your observations in the table. You do **not** need to test and identify any gases evolved. If there is no observable change, write **no observable change**.

Table 3.3

	tests	observations
1.	To 1 cm depth of **FA 6** in a test-tube, add aqueous ammonia dropwise, with shaking, until no further change is seen.	Blue ppt formed, soluble in excess aqueous ammonia to give dark blue solution.
2.	To 1 cm depth of **FA 6** in a test-tube, add an equal volume of **FA 7**.	Blue ppt formed, soluble in excess **FA 7** to give dark blue solution.
	Add 1 cm depth of **FA 8** and shake the mixture thoroughly.	Dark blue solution turned lighter blue (due to dilution).
	Warm the mixture in a water bath. Observe the mixture until no further changes are seen.	Effervescence of gas Blue solution turned purple and then colourless. Brick-red / red-brown precipitate formed.

[4]

Identify the cation present in **FA 6** and the functional group present in **FA 8**.

[2]

FA 6: **<u>Cu^{2+}(aq)</u>**
FA 8: (aliphatic) **<u>Aldehyde</u>**

(ii) Electrolysis of FA 6

1. Repeat steps **1** to **3** of the procedure in **3(a)** using **FA 6**.
2. Immerse one end of the graphite (cathode) in **FA 6** for a few minutes until it is coated with a **thin layer of solid** and record your observations in Table 3.4. You do **not** need to identify any gases evolved.
3. Remove the graphite (cathode) from **FA 6** and rinse it with deionised water.
4. Retain the graphite (cathode) for further tests in **(b)(iii)**.

Table 3.4

electrolyte	observations at the electrodes
FA 6	anode: **bubbles** of colourless, odourless gas observed
	cathode: **pink/red-brown solid** deposited

[1]

(iii) Investigation of some chemical reactions involving the solid deposited

Perform the tests described in Table 3.5, and record your observations in the table. Test and identify any gases evolved. If there is no observable change, write **no observable change**.

Table 3.5

	tests	observations
1.	Put the graphite (cathode) from **(b)(ii)** in a boiling tube.	
	Carry out the following steps in the fume hood.	
	Cautiously add 2 cm depth of **concentrated hydrochloric acid** to the boiling tube and shake the mixture thoroughly to dissolve, as much as possible, the solid coated on the electrode.	Pink/red-brown solid dissolved to give yellow solution.
	Add an equal volume of **FA 6** to the boiling tube.	Yellow solution turned green.
Decant the solution into another boiling tube and retain it for tests 2 and 3.		
2.	Put about 2 cm depth of the solution from test 1 into a test-tube.	
	Add an equal volume of potassium iodide and shake the mixture thoroughly.	Green solution turned brown/orange. White/cream precipitate formed.
	Filter the mixture.	Brown/orange solution as filtrate and White/cream solid as residue.
3.	Put about 2 cm depth of the solution from test 1 into a test-tube.	

	Add 1 cm depth of deionised water to the test-tube. Then add a small piece of aluminium foil and shake the mixture thoroughly. Observe the mixture until no further changes are seen.	Green solution turned pale blue/pale green. Effervescence of a colourless, odourless gas that extinguished a lighted splint with a 'pop' sound. The gas is hydrogen. Aluminium foil dissolved and red-brown solid formed on its surface. Pale blue/ pale green solution decolourised.

[6]

(c) Consider your observations in Table 3.5.

(i) Explain the colour changes you observed in test 1 in Table 3.5, in terms of the ions present in the solutions.

Copper deposited on the electrode dissolved in concentrated HC*l* to form **$[CuCl_4]^{2-}$** ions which gave the yellow solution.

FA 6 is blue due to the presence of **$[Cu(H_2O)_6]^{2+}$** ions in the solution.

On adding **FA 6** to the yellow solution, **a mixture of $[Cu(H_2O)_6]^{2+}$ and $[CuCl_4]^{2-}$** ions gave a green solution.

[2]

(ii) State the role of potassium iodide in test 2 in Table 3.5. Explain your answer in terms of oxidation state changes.

Potassium iodide is a **reducing agent** in test 2. The **oxidation number of iodine increased from –1 in I^- to 0 in I_2** while the oxidation number of copper deceased from +2 in Cu^{2+} to +1 in Cu^+.

[1]

(d) Identification of ions present in FA 9

(i) Perform the tests described in Table 3.6, and record your observations in the table. Test and identify any gases evolved. If there is no observable change, write **no observable change**.

Table 3.6

tests		observations
1.	Test the **FA 9** solution using Universal Indicator paper.	Green UI paper turned dark brown and solution is **<u>pH 1</u>**
2.	To 2 cm depth of **FA 9** in a test-tube, add barium nitrate dropwise, with shaking, until no further change is seen.	white ppt formed

[2]

(ii) Suggest the identity of **FA 9**. Describe one chemical test using a reagent from the Qualitative Analysis Notes provided to confirm the identity of the **cation** in **FA 9**.

[2]

FA 9 is dilute **<u>sulfuric acid</u>**.

Add **<u>CO_3^{2-}</u>**(aq) ions to **FA 9** and **<u>CO_2 will be liberated, forming white ppt with $Ca(OH)_2$(aq)</u>**, if H^+ is present.

[Total: 29]