THE UNITED REPUBLIC OF TANZANIA

NATIONAL EXAMINATIONS COUNCIL OF TANZANIA

CERTIFICATE OF SECONDARY EDUCATION EXAMINATION

032/2 **CHEMISTRY 2**

ALTERNATIVE TO PRACTICAL

(For Both School and Private Candidates)

Time: 3 Hours **ANSWERS** **Year: 2012**

Instructions

1. This paper consists of five questions. Answer all questions.

2. Each question carries 10 marks

1. The diagram below represents the apparatus used in demonstrating the preparation of oxygen gas. Study it very carefully and then answer the questions that follow.

(a)(i) Name the parts labelled A, B, C, D, E, and F.

A – Thistle funnel, used to introduce liquid reactants into the reaction flask.
B – Round-bottom flask, where the reaction occurs to produce oxygen gas.
C – Delivery tube, used to transport the oxygen gas to the collection setup.
D – Water trough, which holds liquid X to facilitate gas collection.
E – Gas jar, where oxygen gas is collected.
F – Gas jar cover, used to prevent gas escape.

(ii) Identify liquids X, Z, and solid Y.

Liquid X – Water, used for gas displacement.
Liquid Z – Hydrogen peroxide (H_2O_2), which decomposes to produce oxygen.
Solid Y – Manganese dioxide (MnO_2), which acts as a catalyst.

(b)(i) Both liquids X, Z, and solid Y contain oxygen. Which one can decompose to give out oxygen gas?

Hydrogen peroxide (H_2O_2) can decompose to give oxygen gas.

(ii) Write a balanced chemical equation for the decomposition of the compound you have named in (b) (i).

$2H_2O_2(l) \text{ ---> } 2H_2O(l) + O_2(g)$
(c) What is the name of the method of collecting oxygen gas as shown in the diagram? Give two reasons to support your answer.

The method used is the downward displacement of water.

Reasons:
1. Oxygen gas is slightly soluble in water, so it does not dissolve significantly during collection.
2. Oxygen gas is less dense than water, allowing it to displace the water in the gas jar.

2. In an experiment, solution R containing 4 g of pure sodium hydroxide in 1 dm^3 was titrated against solution Q containing 5.32 g of impure sulphuric acid in 1 dm^3. Methyl orange indicator was used. The volume of pipette used was 25 cm^3 and the results were recorded in Table 1.

Table 1

Titration	Pilot	1	2	3
Final reading (cm^3)	25.00	25.50	24.50	24.50
Initial reading (cm^3)	0.00	0.50	0.00	0.00

| Volume used (cm^3) | 25.00 | 25.00 | 24.50 | 24.50 |

(a) Complete Table 1 by filling the values of the missing volumes.
The completed table is shown above.

(b) Write a balanced chemical equation for the reaction at the neutral point.
$H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$

(c) Showing your procedures clearly, calculate the percentage purity of sulphuric acid.

Moles of NaOH in 25 cm^3 = (4 g/dm^3 ÷ 40 g/mol) × (25/1000)
= 0.0025 moles
Since the reaction ratio is 1:2, moles of H_2SO_4 = 0.0025 ÷ 2 = 0.00125 moles

Mass of H_2SO_4 = 0.00125 × 98 = 0.1225 g

Percentage purity = (0.1225 / 5.32) × 100
= 23.03%

3. A student investigated the electrolysis of copper sulfate solution as follows:
Two pieces of copper were weighed. One piece was used as the positive electrode and the other as the negative electrode. After electrolysis:
(i) The positive electrode was dipped in propanone.
(ii) Washed with distilled water.
(iii) Dried by evaporation.

The electrodes were weighed and the data were tabulated in Table 2.

Table 2

Electrode	Before electrolysis	After electrolysis
Positive electrode	16.41 g	16.10 g
Negative electrode	15.46 g	15.75 g

(a) Explain why the electrode would dry faster when washed with propanone instead of water.

Propanone is volatile and evaporates quickly, carrying away moisture and allowing the electrode to dry faster than with water.

(b) What is the change in mass of the electrodes?

Mass lost at anode = 16.41 - 16.10 = 0.31 g
Mass gained at cathode = 15.75 - 15.46 = 0.29 g

(c) The mass lost by the positive electrode should have been equal to the mass gained by the negative electrode. Suggest six reasons why the results were not as expected.

1. Some copper ions remained in solution instead of depositing.
2. Impurities in the copper electrodes affected deposition.
3. Copper ions may have undergone secondary reactions.
4. Incomplete drying of the electrodes led to weighing errors.
5. Loss of copper particles during washing.
6. Experimental errors in measuring mass.

(d) If the electrodes were replaced by platinum, what changes, if any, are observed at the cathode, at the anode, and in the solution?

At the cathode – No deposition occurs since platinum is inert.
At the anode – Oxygen gas is evolved instead of copper dissolution.
In the solution – The blue color of Cu^{2+} remains unchanged.

(e) Describe how electrolysis is used to make pure copper from a lump of impure copper.

Impure copper is used as the anode, and pure copper is used as the cathode in a copper sulfate solution. On electrolysis, Cu^{2+} ions from the anode dissolve into solution and migrate to the cathode, where they deposit as pure copper. Impurities fall as sludge below the anode.

4. Mary investigated the rate of reaction between marble ($CaCO_3$) and hydrochloric acid. She used an excess of marble and measured the mass of the flask and contents every half minute for five and a half minutes. The results are shown in Table 3.

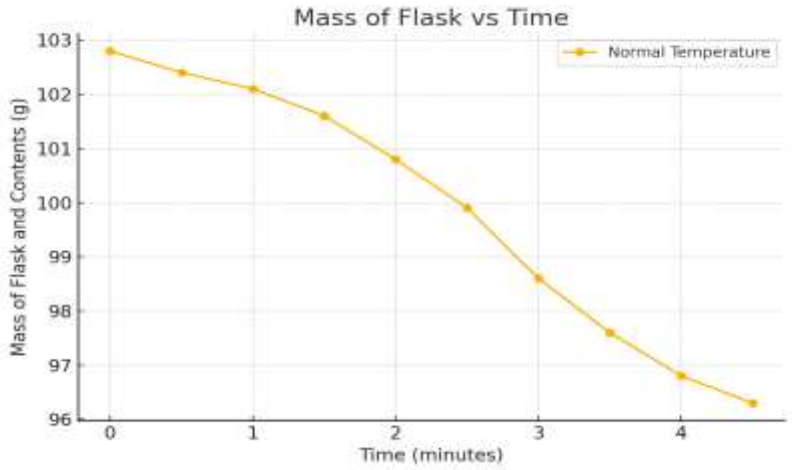
Table 3

Time (min)	Mass of flask and contents (g)	Mass lost (g)
0.0	102.8	0.00
0.5	102.4	0.30
1.0	102.1	0.70
1.5	101.6	1.30
2.0	100.8	2.30
2.5	99.9	3.40
3.0	98.6	4.20
3.5	97.6	4.80
4.0	96.8	5.30
4.5	96.3	5.30

(a) With state symbols, write a balanced equation for reaction between marble and hydrochloric acid.

$CaCO_3(s) + 2HCl(aq) \text{ ----> } CaCl_2(aq) + CO_2(g) + H_2O(l)$

(b) Plot the graph of mass of flask and contents (grams) against time (minutes).



(c) Why did the mass of the contents in the flask decrease with time?

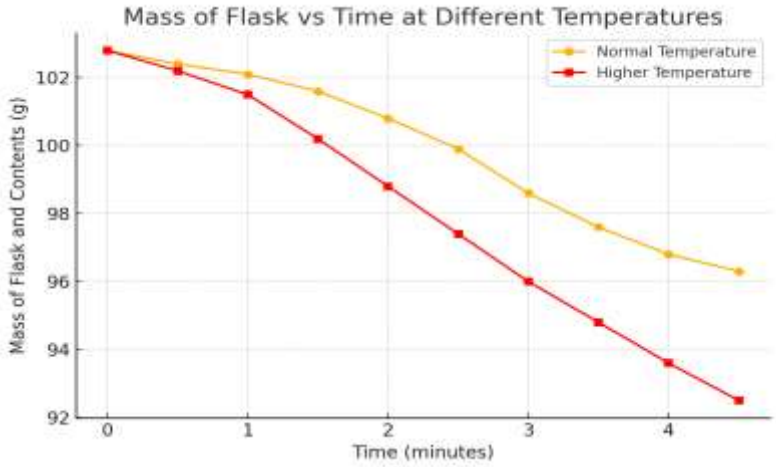
The mass decreased due to the escape of carbon dioxide gas (CO_2) produced in the reaction.

(d) After how many minutes had all the acid been used?

From the table, the mass loss remains constant after 4.5 minutes, indicating the reaction had stopped.

(e) Mary repeated the experiment at high temperature. All other variables were kept the same as in the first experiment. The rate of reaction was much faster.

(i) Sketch on the same graph to show the results for the second experiment. Comment on the shape of the two graphs.



(ii) Why does an increase in temperature increase the rate of reaction? Give three reasons.

1. Increases kinetic energy of molecules, leading to more frequent collisions.
2. Increases the proportion of molecules with energy equal to or greater than activation energy.
3. Enhances effective collisions, increasing reaction rate.

5. An unknown sample N was analyzed and found to contain one cation and one anion. Complete Table 4 and identify the cation, anion, and write the formula and name of the compound.

Table 4

S/n	Experiment	Observations	Inferences
(a)	Appearance of sample N.	White crystalline solid.	SO_4^{2-} or Fe^{2+} may be present.
(b)	Sample N was heated strongly in a dry test tube.	No gas evolved.	Absence of HCO_3^- or CO_3^{2-}.
(c)	Dilute hydrochloric acid was added to the solid sample N in a test tube.	Solution formed without effervescence.	Fe^{2+} present.
(d)	Sample N was dissolved in distilled water and divided into three portions.	Clear solution formed.	NO_3^-, SO_4^{2-}, or Cl^- may be present.
(e)	To the first portion of the solution, sodium hydroxide solution was added.	A green precipitate was formed.	Fe^{2+} present.
(f)	To the second portion, dilute hydrochloric acid was added followed by barium chloride.	White precipitate was formed.	SO_4^{2-} present.
(g)	To the third portion, potassium ferrocyanide solution was added.	Light blue precipitate was formed.	Fe^{2+} confirmed.

Conclusion

(i) The cation in sample N is Fe^{2+}, and the anion is SO_4^{2-}.

(ii) The formula of compound N is $FeSO_4$.

(iii) The chemical name of compound N is iron(II) sulfate.

(iv) With state symbols write a balanced chemical equation taking place at (b).

$FeSO_4(s) \text{ ---> } Fe^{2+}(aq) + SO_4^{2-}(aq)$