ST ANDREW'S SECONDARY SCHOOL

NAME | TEACHING GROUP

Chemistry 6092

PRACTICAL Theory Practice

Volumetric Analysis

1 White vinegar is a colourless solution containing ethanoic acid.
You are provided with two samples of white vinegar labelled **A** and **B**.
You are going to investigate the reaction between these samples and 0.400 mol/dm^3 aqueous sodium hydroxide.

Instructions
You are going to do four titration experiments.
Rinse and fill a burette with **A**.

(a) Experiment 1

- Use a volumetric pipette to add 25.0 cm^3 of 0.400 mol/dm^3 aqueous sodium hydroxide to a conical flask.
- Add five drops of thymolphthalein indicator to the conical flask.
- Record the initial burette reading in Table 1.1.
- Add **A** from the burette white swirling the flask, adding drop by drop near the end-point, until the solution changes colour.
- Record the final burette reading in Table 1.1.

Experiment 2

- Empty the conical flask and rinse it with distilled water.
- Refill the burette if necessary.
- Repeat Experiment 1.

Complete Table 1.1 with the volume used in each experiment.

Table 1.1

	Experiment 1	Experiment 2
Final burette reading/ cm^3	**21.40**	**42.90**
Initial burette reading/ cm^3	**0.00**	**21.40**
Volume of **A** used/ cm^3		

(b) Experiments 3 and 4

- Pour away any **A** remaining in the burette.
- Rinse and fill the burette with **B**.
- Empty the conical flask and rinse it with distilled water.
- Repeat Experiments 1 and 2 using **B**.
- Complete Table 1.2 for Experiments 3 and 4.

Table 1.2

	Experiment 3	Experiment 4
Final burette reading/ cm^3	**33.10**	**42.90**
Initial burette reading/ cm^3	**0.00**	**10.00**
Volume of **B** used/ cm^3		

(c) Calculate the mean volume of **A** and **B** needed to neutralise 25.0 cm^3 of 0.400 mol/dm^3 aqueous sodium hydroxide. [2]

(d) Suggest why the titrations using **A** and **B** are repeated.

..... [1]

(e) The equation for the reaction between ethanoic acid and sodium hydroxide is shown.

$$CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$$

The answer to **(c)** shows the mean volume of **B** used to neutralise 25.0 cm^3 of 0.400 mol/dm^3 aqueous sodium hydroxide.

Calculate the concentration (in mol/dm^3) of ethanoic acid in **B**. Give your answer to an appropriate number of significant figures. [2]

(f) Use your answer to **(e)** to calculate the mass of ethanoic acid in 500 cm^3 of **B**. [2]
[A_r: C, 12; O, 16; H, 1]

(g) Use your answers to **(c)** and **(f)** to calculate the mass of ethanoic acid in 500 cm^3 of **A**. [1]

(h) Suggest why the conical flask should be rinsed with distilled water and **not** 0.400 mol/dm^3 aqueous sodium hydroxide between titrations.

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(i) Some vinegars are brown instead of colourless.
Suggest why this titration method would **not** be suitable for finding the concentration of ethanoic acid in brown vinegar.

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2 Milk of magnesia is a suspension of insoluble magnesium hydroxide in water. It is taken by people who have stomach pain caused by indigestion and works by neutralising acid in the stomach.

P is an aqueous solution of volume 110 cm^3 prepared by reacting 10.0 cm^3 of milk of magnesia with an excess of hydrochloric acid, HC*l*. In preparing **P**, all the magnesium hydroxide in the 10.0 cm^3 of suspension reacted when it was added to 100 cm^3 of 1.00 mol/dm^3 hydrochloric acid, an excess.

$$Mg(OH)_2 + 2\ HCl \rightarrow MgCl_2 + 2\ H_2O$$

You are to determine by titration the amount of acid remaining in **P**.
Q is 0.527 mol/dm^3 sodium hydroxide, NaOH.

(a) Put **P** into the burette.
Pipette a 25.0 cm^3 portion of **Q** into a conical flask and titrate with **P**, using an indicator.
Record your results in the table, repeating the titration as many times as you consider necessary to achieve consistent results.

Results

Titration number	1	2	3	4	5
Final burette reading/ cm^3	19.70	41.60	20.80	40.70	19.80
Initial burette reading/ cm^3	0.00	19.70	1.00	20.80	0.00
Volume of **P** used/ cm^3					
Best titration results (✓)					

Use the best titration results in the table to calculate the average volume of **P** used. [1]

(b) **Q** is 0.527 mol/dm^3 sodium hydroxide, NaOH.
Calculate the number of moles of sodium hydroxide in the volume of **Q** used. [1]

(c) Using the equation shown and your answer to **(b)**, deduce the number of moles of hydrochloric acid that reacted with the volume of **Q** used. [1]

(d) Using your answer to **(c)** and the average volume of **P** calculated from the titration results, calculate the number of moles of hydrochloric acid in 110 cm^3 of **P**. [1]

(e) Calculate the number of moles of hydrochloric acid in 100 cm^3 of 1.00 mol/dm^3 hydrochloric acid. [1]

(f) Using your answers from **(d)** and **(e)**, calculate the number of moles of hydrochloric acid that reacted with the magnesium hydroxide in the milk of magnesia. [1]

(g) Using your answer to **(f)**, calculate the concentration of magnesium hydroxide in milk of magnesia in g/dm^3. [2]
[A_r: Mg, 24; O, 16; H, 1]

3 Brass is an alloy of copper used for making screws.
You are going to use titration to determine the percentage by mass of copper in a brass screw.

The brass screw has been dissolved in nitric acid to form a solution containing aqueous Cu^{2+} ions.
When excess aqueous potassium iodide is added to the aqueous Cu^{2+} ions, iodine and a white precipitate of copper(I) iodide are produced.

$$2\ Cu^{2+}\ (aq) + 4\ I^-\ (aq) \rightarrow 2\ CuI\ (s) + I_2\ (aq)$$

The amount of iodine produced is found by titration with aqueous thiosulfate, $S_2O_3^{2-}$.

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

P is 0.120 mol/dm^3 aqueous thiosulfate ions.
Q contains aqueous Cu^{2+} ions made from dissolving a 9.45 g brass screw in nitric acid and the solution is made up to 1.00 dm^3.
R is 0.5 mol/dm^3 aqueous potassium iodide.
S is starch solution which you will use as an indicator.

(a) (i) Instructions

- Put **P** into the burette.
- Pipette 25.0 cm^3 of **Q** into a conical flask.
- Use a measuring cylinder to add 20 cm^3 of **R**, which is an excess, to **Q**. A white precipitate forms in a dark brown solution.
- Add **P** from the burette to the flask until the solution turns light brown. **Then** use a measuring cylinder to add about 1 cm^3 of **S** to the flask. A dark-coloured solution forms.
- Continue adding **P** slowly from the burette until one drop of **P** causes the dark colour to disappear leaving a white precipitate.
- Record your titration results in an appropriate format in the space provided. Repeat the titration as many times as you consider necessary.

Results

Titration number	1	2	3	4
Final burette reading/ cm^3	24.90	48.80	47.50	24.50
Initial burette reading/ cm^3	0.00	24.90	22.90	0.00
Volume of **P** used/ cm^3				
Best titration results (✓)				

(ii) Use your titration results to obtain the average volume of **P** used. [1]

(b) **(i)** Calculate the amount in moles, of thiosulfate ions present in the average volume of **P**. [1]

(ii) Use the equations to determine the amount in moles, of Cu^{2+} ions in 1.00 dm^3 of **Q**. [2]

(iii) Use your answer to **(b)(ii)** to calculate the mass of copper in the brass screw. [1]
[A_r: Cu, 64]

(iv) Calculate the percentage by mass of copper in the brass screw. [3]

(c) A student repeats the experiment but adds a smaller volume of aqueous potassium iodide which is not in excess.

Describe and explain the effect this has on the calculated value for the percentage by mass of copper in the brass screw.

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BONUS!

4 Calcium carbonate reacts with hydrochloric acid. This reaction is exothermic.

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

The student investigates the amount of heat released during this reaction by using the following method.

1. Weigh the container with the calcium carbonate.
2. Place 50.0 cm^3 of hydrochloric acid, which is an excess, in an insulated plastic cup and measure the temperature of the acid. Record the temperature in the table.
3. Start a stopwatch and measure the temperature of the acid every minute for 2 minutes. Record the temperatures.
4. At 3 minutes transfer the calcium carbonate from the container into the cup. Stir the contents and continue to measure, and record, the temperature every minute up to 8 minutes.
5. Reweigh the container with any residual solid.

The results the student obtains are shown.

Time / min	0	1	2	3	4	5	6	7	8
Temperature / °C	22.0	22.0	22.0		26.0	25.5	25.5	24.5	24.0

mass of container with calcium carbonate = 24.8 g
mass of container after adding calcium carbonate to the cup = 21.3 g

(i) Plot the student's results on the grid with temperature on the *y*-axis and time on the *x*-axis.

Complete the graph by drawing **two** straight lines of best fit. One line should be drawn using data from 0 to 2 minutes and the other line should be drawn using data from 4 to 8 minutes.

Extend both of these lines to 3 minutes.

Draw a vertical line at 3 minutes to connect these lines.

[4]

(ii) Use your lines from **(i)** to calculate the maximum temperature rise in the experiment. [1]

(iii) The amount of heat released in the experiment can be calculated using the expression shown.

heat released (in J) = mass of solution (in g) × maximum temperature rise (in °C) × 4.2

Use this expression to calculate, in J, the amount of heat released.

Assume that 1.0 cm^3 of solution has a mass of 1.0 g. [1]

(iv) The molar mass of calcium carbonate is 100 g mol^{-1}.

Use this information to calculate the number of moles of calcium carbonate in the experiment. [1]

(v) Calculate the amount of heat released, in kJ/mol, when one mole of calcium carbonate reacts with hydrochloric acid. [1]

(vi) Another student carries out the same experiment but only measures the initial temperature of the solution and the highest temperature reached in order to calculate the temperature rise.

Suggest why the student who uses a graph to calculate the temperature rise is able to obtain a more accurate value for the heat released.

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