1. Making salt – an introduction to some basic techniques – Student Sheet

This practical is intended to help you become confident in some of the skills you may have used in your previous chemistry course. It will also help you revise some concepts and learn new ones.

Intended lesson outcomes

By the end of this practical you should be able to:
- manipulate apparatus successfully;
- read a burette and use a pipette;
- use a Bunsen burner;
- carry out a titration experiment;
- record your titration results in the tables provided;
- use and understand molecular and ionic equations;
- use the mole concept.

Background information

An acid neutralises a base to form a salt and water. Hydrochloric acid and sodium hydroxide are completely ionised in water. We say they are a strong acid or base because they are completely ionised in solution The ions present in hydrochloric acid are $H^+(aq)$ and $Cl^-(aq)$ and in sodium hydroxide are $Na^+(aq)$ and $OH^-(aq)$.

You are going to use the technique of **titration** to produce a sodium chloride solution. Titration is a very accurate way of investigating the reaction of two solutions. It can be used to analyse the amount of a particular substance in a solution. This is known as **quantitative analysis**. In a titration, one solution is placed in a burette and the other is placed in a conical flask using a pipette. The solution in the burette is then run into the conical flask until there is a complete reaction. In this case you will completely neutralise a solution of hydrochloric acid with sodium hydroxide solution. You will use an indicator to tell you when there is complete neutralisation. The indicator changes colour at the exact point of neutralisation. In this case you may use any acid-base indicator because you will be titrating a strong acid with a strong base.

Questions

1 **(a)** Write the word and symbol equation for the reaction of hydrochloric acid with sodium hydroxide.

(b) Write the equation again, this time using the ion symbols.

(c) Only two of the ions you have written in 1(b) take part in the neutralisation reaction. Write the neutralisation equation without the **spectator ions**. The spectator ions are those that appear on both sides of the equation.

(d) Sodium hydroxide is a soluble base. What is the name given to soluble bases?

Safety

(eye protection symbol)	You must wear eye protection throughout this experiment.
(corrosive symbol)	Sodium hydroxide and hydrochloric acid are **corrosive**. (Although the solutions are very dilute, you should mop up any spillages and rinse any solution off your skin or clothing with water).

Always use a pipette filler, or other suitable safety device. Never be tempted to use your mouth to draw liquid into the pipette.

Procedure

1. Rinse two beakers with deionised water. Label one beaker *alkali* and wash this out with a little of the 0.1 mol dm^3 sodium hydroxide solution. Do this twice to ensure that any water used during the rinsing of the beaker will not dilute the sodium hydroxide solution you are about to put in the beaker. Fill the beaker about half full with the sodium hydroxide solution.

    Now label the other beaker *acid* and after rinsing with water, wash it out twice using small volumes of hydrochloric acid before filling this half full with the acid solution. At the moment you do not know the exact concentration of the hydrochloric acid solution. The titration you do will allow you to calculate this.

2. Pipette exactly 25.0 cm^3 sodium hydroxide solution into a clean conical flask (Fig. 1.1). The pipette can be cleaned in a similar way to the beakers, remembering to finish by washing it out with a little of the alkali solution twice to ensure no dilution will take place when using the pipette.

    In this case it does not matter if there is some deionised water left in the flask after rinsing it. Now add one or two drops of acid-base indicator.
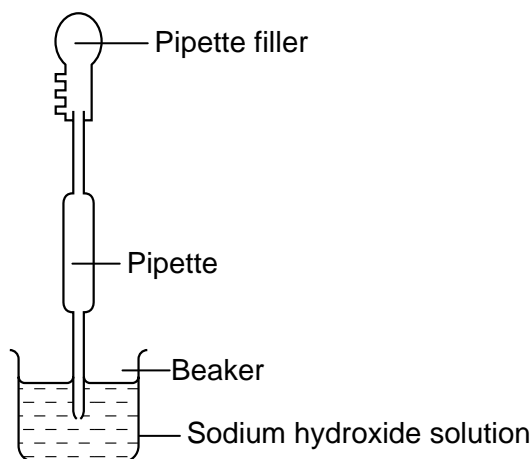

Figure 1.1 A pipette safety filler is used to draw a measured volume of sodium hydroxide solution from the beaker into the pipette.

3. Once the burette has been rinsed and washed out with the acid solution, fill it nearly to the top (Fig. 1.2).

    Clamp the burette carefully and run a little acid through into the beaker until the tip becomes full.
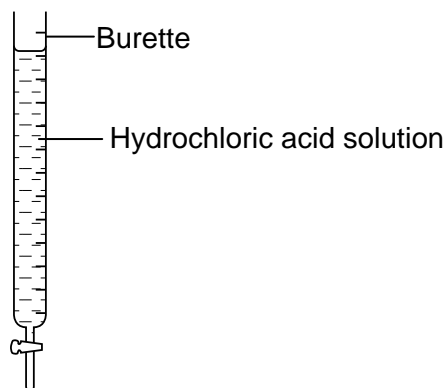

Figure 1.2 Fill the burette with hydrochloric acid solution and ensure the tip is full.

4. Now read the burette and record the reading in the middle row of a table like the one below. Be careful that your eye is level with the bottom of the meniscus or your reading will not be accurate (Fig. 1.3).
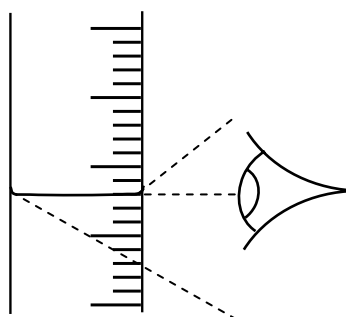

Figure 1.3

	rough titration	1	2
final burette reading/cm^3			
initial burette reading/cm^3			
volume of acid added/cm^3			

5. Place the conical flask below the burette on a piece of white paper. Run the acid into the flask fairly quickly, shaking it all the time. As soon as the colour of the indicator changes, close the tap and note the final burette reading. Record this result in your table above your initial reading. Subtract the initial reading from the final reading to give you the volume of acid added.

6. The first titration is a rough titration to give you an idea of the volume you need to add to exactly neutralise the acid. It is quite likely that you added a slight excess of acid as you were doing the titration quickly. Now repeat steps 2 to 5 but this time run in the acid quickly until you reach about 1 cm^3 less than the volume you added in the rough titration. Swirl the contents of the flask and add one drop of acid at a time from the burette until the indicator just changes colour. Record this volume. This should represent the exact volume you need to add to neutralise 25 cm^3 of 0.1 mol dm^3 of sodium hydroxide solution.

7. To ensure that you have a reliable volume of alkali, you should repeat the whole titration again until you get two readings that agree within 0.05 cm^3.

8. Now you are sure of the volume of sodium hydroxide you require, repeat the titration without using the indicator. This should give you a pure sodium chloride solution, without any excess acid or alkali.

9. Evaporate all the water away from a small sample of your neutral solution using the apparatus in Fig. 1.4. Be sure to adjust your Bunsen burner to give a suitable flame. You will be left with small, white crystals of sodium chloride.
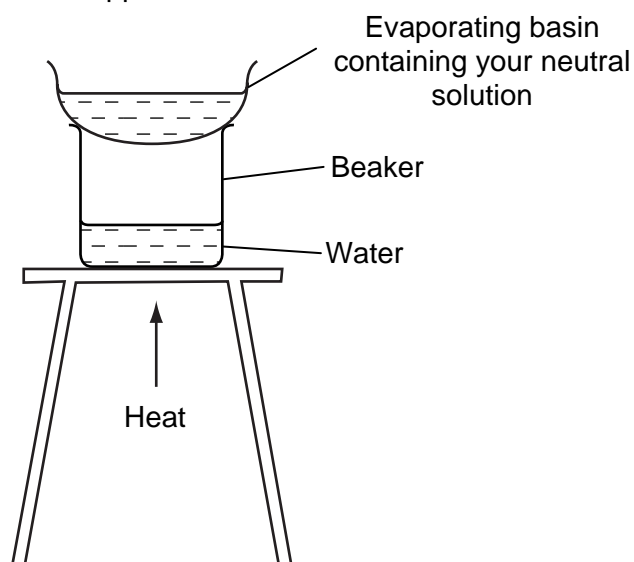
Appendix 2

Figure 1.4

Questions

2   **(a)** How many moles of NaOH were present in 25.0 cm^3 solution?

   (b) How many moles of HCl were present in the volume of acid you used to neutralise the NaOH solution? To calculate this you will need to refer to the equations you wrote in Question 1.

   (c) What was the exact concentration of the hydrochloric acid in mol dm^{-3}?

   (d) How many moles of sodium chloride were produced in the solution in the conical flask?

   (e) What mass of salt was produced in the neutral solution?

3   Look back at the experimental procedure.

   (a) Why was the conical flask placed on a piece of white paper?

   (b) Why were the pipette and the burette washed with the solutions they were going to contain?

   (c) Why was the conical flask not washed with the alkali solution it was going to contain?

   (d) Explain why it does not matter if there is water already in the flask.

   (e) Explain why a conical flask was used and not a beaker.

1. Making salt – an introduction to some basic techniques – Teachers' Notes

This practical focuses on revising some basic techniques, such as the correct use of the Bunsen burner when heating. It gives students the opportunity to set up and manipulate apparatus and provides an introduction to a simple titration. Students will also record their results in a table.

Learning outcomes

These are printed on the Student Sheet.

A suggested approach

If students have not used a burette or pipette before, it would be sensible to spend a few minutes demonstrating the techniques involved in titration.

The Student Sheet gives full information about how to do the practical. Students could prepare for the practical session by reading the sheet in advance and answering Question 1. The students should finish the titration in a 1-hour session. The neutral sodium chloride solutions should be kept if you intend to evaporate off the water. It is motivating and fun for students to actually see the product that they have made and the second lesson could also be used to go through the mole calculations. It is also important to remind students about the correct use of the Bunsen burner.

Technical information

Requirements per student/group:

For the titration
Two 100 cm^3 beakers
25.0 cm^3 pipette
Pipette filler or equivalent safety device
50.0 cm^3 burette
Conical flask
Sheet of white paper or a white tile
Wash bottle and deionised water
0.10 mol dm^{-3} NaOH(aq) – about 80 cm^3 **(corrosive)**
Approximately 0.1 mol dm^{-3} HCl (aq) – about 80 cm^3 **(irritant)**
Stoppered flask to store a sample of the neutral solution prepared

For the evaporation of the neutral solution
Evaporating basin
250 cm^3 beaker
Tripod and gauze
Heating mat
Bunsen burner

Safety

The main points are included on the Student Sheet but it is the teacher's responsibility to ensure that a full risk assessment is carried out prior to the practical session(s). MSDS sheets should be consulted so that the correct action can be taken in event of a spillage and/or accident.

Appendix 2

Answers to questions on the Student Sheet

1 **(a)** sodium hydroxide + hydrochloric acid → sodium chloride + water

$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

(b) $Na^+(aq) + OH^-(aq) + H^+(aq) + Cl^-(aq) \rightarrow Na^+(aq) + Cl^- + H_2O(l)$

(c) $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

(d) Alkalis.

2 **(a)** Moles of NaOH in $25cm^3$ 0.1 mol dm^{-3} = 0.1 x 25/1000 = 2.5 x 10^{-3} mol

(b) This will be the same as **2(a)** above.

(c) This is dependent on the volume of HCl used but it will be approximately 0.1 mol dm^{-3}.

(d) This will be the same numerical answer as in **2(a)** and **2(b)**.

(e) The moles given in **2(d)** are multiplied by M_r NaCl (58.5).

3 **(a)** To see the end point more clearly.

(b) To ensure there was no dilution when they were filled with the solutions to be titrated.

(c) So that no extra moles of NaOH were present in the flask.

(d) The amount of water present does not affect the moles of NaOH that were pipetted into the flask.

(e) The sloping sides prevent splashes, which would lose moles of the acid or alkali and make the titration inaccurate.

44