

Titrating sodium hydroxide with hydrochloric acid

This investigation is part of the **Nuffield practical collection**, developed by the Nuffield Foundation and the Royal Society of Chemistry. Delve into a wide range of chemical concepts and processes with this collection of over 200 step-by-step practicals: rsc.li/43bjGqI

Learning objectives

1. Set up and carry out a neutralisation reaction by titration.
2. Calculate the concentration of hydrochloric acid using experimental results.
3. Produce a pure sample of sodium chloride crystals from sodium chloride solution.
4. Accurately record experimental observations and data.
5. Calculate the percentage yield of sodium chloride (challenge).

Success criteria

During stage 1 of the practical, learners set up and carry out a neutralisation reaction by titration, LO1. Judge how successfully this LO has been met by observing how learners carry out the practical and the quality of their answers to the stage 1 questions. Learners then go on to calculate the actual concentration of the acid used, meeting LO2.

During stage 2 and 3 of the practical, learners go on to produce a pure sample of chloride crystals (LO3). Judge how successfully this LO has been met by observing how learners carry out the practical and the quality of their answers to the stage 2 questions and stage 3 questions 2 and 3.

To achieve LO4, learners must carefully record the volume of acid that is neutralised by the sodium hydroxide (stage 1 Q5) and the mass of the dry crystals (stage 3, method point 3).

Learners must apply their knowledge of moles, concentration and percentage yield calculations to answer stage 3 question 5 and achieve LO5, which will provide extra challenge.

Introduction

In this experiment learners neutralise sodium hydroxide with hydrochloric acid to produce the soluble salt sodium chloride in solution. They then concentrate the solution and allow it to crystallise to produce sodium chloride crystals

Scaffolding

Decide if this experiment is suitable to use with different classes and if learners need preliminary training in using the techniques involved in titration. The following

practical assumes that you have judged the class to be capable of doing this experiment using a burette with reasonable expectation of success.

There are three stages to this practical. In the accompanying presentation, each stage has three slides including a written method, integrated instructions and a set of questions. Depending on the class and the time available, learners can either work through all three stages of the practical or just carry out part of it. For example, you can do stage 1 as a teacher demonstration and once the required volume of hydrochloric acid had been determined, learners can carry out stages 2 and 3 for themselves.

Alternatively, carry out the practical over two lessons. In the first lesson, learners concentrate on the titration and producing concordant results. They will use this result in the second lesson, to produce a pure sample of sodium chloride.

There are two versions of the student worksheet available to download from rsc.li/3O2L8TL. Only the first four LOs appear on the scaffolded sheet (★) whereas all five LOs appear on the unscaffolded worksheet (★★).

Technician notes

Please read our standard health and safety guidance (rsc.li/3zyJLkx) and carry out a risk assessment before running any live practical.

Start the practical with the apparatus ready at each work space in the laboratory to avoid fragile and expensive glassware (the burette) being collected from an overcrowded central location.

Time required

Stage 1

Filling the burette, measuring out the alkali into the flask, and titrating it until it is neutralised takes about 20 minutes, with false starts being likely for many groups. In practice it does not matter if the endpoint is overshoot, even by several cubic centimetres, but the aim is to find the proportions for a roughly neutral solution.

Stage 2

Producing a neutral solution free of indicator, should take no more than 10 minutes.

Stage 3

Evaporating the solution may take the rest of the lesson to the point at which the solution can be left to crystallise for the next lesson. Watching solutions evaporate can be tedious for learners, so give them another task to keep them occupied, e.g. rinsing and draining the burettes with purified water.

Equipment

Apparatus (per group)

- Safety glasses
- Burette, 30 or 50 cm³ (note 1)
- Conical flask, 100 cm³
- Beaker, 100 cm³
- Pipette, 20 or 25 cm³, with pipette filter
- Stirring rod
- Small (filter) funnel, about 4 cm diameter
- Burette stand and clamp (note 2)
- White tile or sheet of white paper
- Bunsen burner
- Tripod
- Pipe clay triangle (note 3)
- Evaporating basin, at least 50 cm³ capacity
- Crystallising dish (note 4)
- Microscope or hand lens suitable for examining crystals in the crystallising dish

Apparatus notes

1. If your school still uses burettes with glass stopcocks, consult the CLEAPSS Laboratory Handbook, section 10.10.1, for their care and maintenance, or consult your local safety authority body. This experiment will not be successful if the burettes used have stiff, blocked or leaky stopcocks. Modern burettes with PTFE stopcocks are much easier to use, require no greasing, and do not get blocked. Burettes with pinchcocks of any type are not recommended; while cheap, they also are prone to leakage, especially in the hands of student beginners.
2. Burette stands and clamps are designed to prevent crushing of the burette by over-tightening, which may happen if standard jaw clamps are used.
3. You can use ceramic gauzes instead of pipe clay triangles, but the evaporation then takes longer.
4. The evaporation and crystallisation stages may not complete in the lesson time. The crystallisation dishes need to be set aside for crystallisation to take place slowly. However, do not allow the dishes to dry out completely, as this spoils the quality of the crystals. With occasional checks, decide when to decant surplus solution from each dish to leave good crystals for the learners to inspect in the following lessons.

Chemicals (per group)

- Sodium hydroxide solution, 0.4 M (IRRITANT), about 100 cm³ in a labelled and stoppered bottle
- Dilute hydrochloric acid, 0.4 M, about 100 cm³ in a labelled and stoppered bottle
- Methyl orange indicator solution (or alternative such as phenolphthalein) in small dropper bottle

**Safety and hazards**

- Wear eye protection throughout.
- Sodium hydroxide solution, NaOH(aq), (IRRITANT at concentration used) – see CLEAPSS Hazcard [HC091a](#) and CLEAPSS Recipe Book RB085, refer to [SSERC](#) or consult your local safety advisory body. You do not need to make the concentration of the solution up to a high degree of accuracy, but it should be reasonably close to the concentration of the dilute hydrochloric acid and less than 0.5 M.
- Dilute hydrochloric acid, HCl(aq) – see CLEAPSS Hazcard [HC047a](#) and CLEAPSS Recipe Book RB043 refer to [SSERC](#) or consult your local safety advisory body. You do not need to make the concentration of the solution up to a high degree of accuracy, but it should be reasonably close to the concentration of the sodium hydroxide solution and less than 0.5 M.
- Methyl orange indicator solution (the solid is TOXIC but not the solution) – see CLEAPSS Hazcard [HC032](#) and CLEAPSS Recipe Book RB000, refer to [SSERC](#) or consult your local safety advisory body.

**Method**

A full method is provided in the student worksheet (rsc.li/3O2L8TL).

Teaching notes

This experiment is best suited to learners aged 14–16. Titration using a burette, to measure volumes of solution accurately, requires careful and organised methods of working, manipulative skills allied to mental concentration and attention to detail. All of these are desirable traits to be developed in learners, but all learners need to be able to use a burette with basic competence and reliability before attempting this practical.

Train learners in using burettes correctly, including how to clamp them securely and fill them safely. Consider demonstrating burette technique and give learners the opportunity to practise this. In this experiment, a pipette is not necessary, as the aim is to neutralise whatever volume of alkali is used and that can be measured roughly using a measuring cylinder.

It is not the intention here to do quantitative measurements leading to calculations but for some learners this can provide an extra challenge or extension work (see the challenge question). The aim is to introduce learners to the titration technique only to produce a neutral solution.

Alternative indicators you can use include screened methyl orange (green in alkali, violet in acid) and phenolphthalein (pink in alkali, colourless in acid).

Leaving the concentrated solutions to crystallise slowly helps produce larger crystals. The solubility of sodium chloride does not change much with temperature, so simply cooling the solution is unlikely to form crystals.

Under the microscope (if possible, a stereomicroscope is best), you can see the cubic nature of the crystals. If crystallisation has occurred in a shallow solution, with the crystals only partly submerged, 'hopper-shaped' crystals may be seen. In these crystals, each cube face becomes a hollow, stepped pyramid shape.

Answers

Stage 1 – both sheets

1. Sodium chloride and water
2. hydrochloric acid + sodium hydroxide → sodium chloride + water
3. $\text{HCl(aq)} + \text{NaOH(aq)} \rightarrow \text{NaCl(aq)} + \text{H}_2\text{O(l)}$
4. To indicate/show when the reaction was over by turning from yellow–orange to red
5. This should be the same as the volume of NaOH used
- 6.

Un scaffolded sheet

The concentration should be around 0.4 mol/dm^3

Scaffolded sheet

	Volume (cm^3)	Volume (dm^3)	Concentration (mol/dm^3)
Sodium hydroxide	25	0.025	0.4
Hydrochloric acid	answer to Q5	answer to Q5 divided by 1000	

- (a) Number of moles = concentration x volume
 (b) $0.4 \times 0.025 = 0.01 \text{ mol/dm}^3$
 (c) 0.01 mol/dm^3
 (d) Concentration of HCl = $0.01/\text{volume from experiment}$

Stage 2*Unscaffolded sheet*

1. The solution from stage 1 also contained indicator, so pure salt crystals could not be obtained. So, the reaction must be done again using fresh sodium hydroxide and hydrochloric acid.
2. The required volume of acid is already known from stage 1. We don't want to add any impurities into the solution.
3. The first reading may not be correct (for example, the end point colour change of the indicator may have been missed) or to improve the accuracy of the readings.
4. Test the pH using a pH meter. It should be pH 7. Dip a clean glass rod into the product and then place a drop onto some universal indicator paper.

Stage 2*Scaffolded sheet*

1. Sodium chloride and water
2. Sodium chloride
3. The indicator would make the salt crystals impure
4. To make sure all the reactants, hydrochloric acid and sodium chloride were used up in the reaction
5. Use a pH meter

Stage 3 – both sheets

1. As the water molecules **evaporate**, a **more** concentrated solution of sodium chloride **ions** forms in the basin. Eventually the solution becomes **saturated** and sodium chloride crystals start to form.
2. Mass in grams of crystals formed.
3. Accept any sketch resembling a cube
- 4.

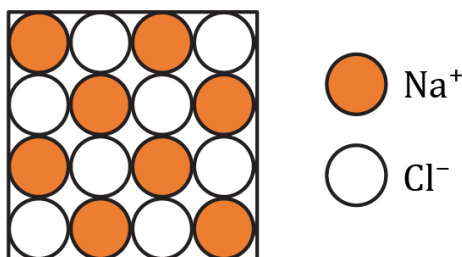
Unscaffolded sheet

Cubic

Scaffolded sheet

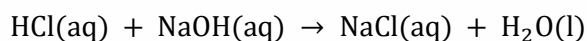
C because it shows a regular lattice structure formed from alternating particles

5. The Na^+ and Cl^- ions are attracted to each other in a giant 3D lattice structure consisting of regular rows of alternating Na^+ and Cl^- ions. The diagram should include a key or individual particles should be labelled. For example:



Challenge question

Two different approaches to doing this calculation are included here.

Approach 1 – using a table

From the balanced equation we know the number of moles of NaCl produced in the reaction will be the same as the number of moles of the starting reactants.

	Mathematical formula	Sodium hydroxide	Sodium chloride
Volume (cm ³)		25	
Concentration (mol/dm ³)		0.4	
Moles	Number of moles = concentration (mol/dm ³) x volume (dm ³)	0.4 x 0.025 = 0.01	0.01
M_r	M_r = sum of A_r		23 + 35.5 = 58.5
Theoretic mass (g)	Mass = moles x M_r		0.01 x 58.5 = 0.585 g

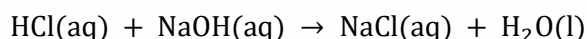
Note this could also be worked out using the volume of hydrochloric acid used

$$\begin{aligned} \% \text{ yield} &= \frac{\text{actual mass}}{\text{theoretical mass}} \times 100 \\ &= \frac{\text{actual mass}}{0.585} \times 100 \end{aligned}$$

Approach 2 – writing out in full

Number of moles of sodium hydroxide = concentration (mol/dm³) x volume (dm³)

$$0.4 \times (25/1000) = 0.01 \text{ mol/dm}^3$$



From the balanced equation we know the number of moles of NaCl produced in the reaction will be the same as the number of moles of the starting reactants.

So, moles of NaCl produced = 0.01 mol

$$M_r \text{ of NaCl} = 23 + 35.5 = 58.5$$

$$\text{Theoretical mass} = \text{moles of NaCl} \times M_r = 0.01 \times 58.5 = 0.585 \text{ g}$$

$$\% \text{ yield} = \frac{\text{actual mass}}{\text{theoretical mass}} \times 100 = \frac{\text{actual mass}}{0.585} \times 100 = \text{your answer}$$

Note, learners will need to substitute their own mass.