

## 18

## Chemical Analysis



## Introduction

The study of chemistry includes the isolation and purification of substances from the Earth and the analysis of these substances using various tests. In Chapter 2, students learnt how substances are purified. In this chapter, they learn how substances are analysed using qualitative analysis to identify the elements and compounds present in substances and quantitative analysis to determine the concentrations of these elements and compounds.

## Chapter Opener (page 268)

1. Begin the chapter by discussing the following questions. Precise answers are not needed at this stage.

***What physical tests can be used to identify substances?***

**Answer:** Refer to Section 18.2 in the Textbook.

***What tests can be used to identify hydrogen and oxygen gas?***

**Answer:** Refer to Table 18.4 in the Textbook.

***What kind of experiment can be done to determine the concentration of an acid?***

**Answer:** Carry out a titration using a solution of sodium hydroxide of known concentration. Refer to Section 18.5 on pages 277–280 of the Textbook.

2. Carry out an 'Inquiry Preview.'

## Learning Outcomes

**After completing this chapter, you should be able to:**

- ▶ describe, with examples, physical tests that can be used to help in the identification of elements and compounds
- ▶ describe tests to identify common cations using aqueous sodium hydroxide and aqueous ammonia
- ▶ describe tests to identify the common anions
- ▶ describe tests to identify ammonia, carbon dioxide, chlorine, hydrogen, oxygen and sulfur dioxide
- ▶ carry out volumetric experiments and apply the concept of concentration to process the results of these experiments

### Teaching pointers

## 18.1 What are the Uses of Chemical Analysis? (page 269)

### Stimulation

Begin by asking the class if they have read or heard of any recent reports about the discovery of harmful substances in foods or commercial products. (Refer back to Chapter 2 page 13 in the Textbook for examples.) If possible, the teacher could show newspaper or magazine clippings of these. Discuss the role of chemical analysis in detecting and identifying the harmful substances in the products.

To introduce quantitative analysis, show a bottle of vinegar or some antacid tablets, identify the main ingredient in each and ask a rhetorical question as to how the amounts of these ingredients might be found.

### Notes for Teachers

#### Additional examples of the use of chemical analysis

- Detecting the impurities in drugs and medicines in general.
- Detecting poisonous pesticides in vegetables.
- Detecting illegal drugs used by athletes.
- Detecting and measuring the concentrations of air pollutants.
- Detecting fertilisers and other pollutants in streams and rivers.
- Determining the composition of vehicle exhaust gases.
- Identifying the drugs carried by traffickers.
- Identifying coloured dyes in food and drinks.

## Teaching pointers

## 18.2 What Physical Tests are Used to Identify Substances? (page 269)

- Most of the material in this section has been covered in earlier chapters. Emphasise the role of *elimination* when trying to identify substances.
- The information strip about Robert Bunsen in 'Chemistry in Society' on page 270 of the Textbook shows how he devised flame tests, use them to identify substances and how he discovered new elements. Use the example of Bunsen to remind students of what can happen if they do not wear safety goggles when doing experiments (refer to point 3 of the strip).

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### Mystery Clue

Iron(III) oxide. (About 40%-45% of the waste was this oxide, which gave it the reddish colour.)

## Chemistry in Society (page 270)

### Robert Bunsen

#### Exercise

- He showed perseverance and willingness to contribute to science through his research. He also showed socio-cultural sensitivity and awareness by being able to empathise through understanding, acceptance and respect.
- The fireworks contain metal compounds. When these are heated, they produce colours. (For example, sodium compounds produce a yellow colour; copper(II) compounds produce a green colour.)

#### Skills Practice (page 271)

- Examples:
  - Is ink a single substance?
  - How many substances does this dye contain?
  - Is this water pure? What (harmful) substances could it contain?
  - What pollutants are in the air?
  - Do these slimming tablets contain harmful compounds?
  - Is this food safe to eat?
  - Is this medicine pure enough to treat people with?
  - Does the alloy for the wings of this new aircraft contain the correct amount of aluminium?
  - Does this paint used in a child's toy contain harmful metals?
  - What elements does this fertiliser contain?
- Quantitative refers to quantities and measurements of quantities. Qualitative refers to qualities and does not involve measurements.
- $K_2CO_3$  – potassium carbonate, white, soluble in water.
    - $(NH_4)_2SO_4$  – ammonium sulfate, white, soluble in water.
    - $CuCO_3$  – copper(II) carbonate, green, insoluble in water.
    - $FeCl_3$  – iron(III) chloride, reddish-brown, soluble in water.
    - $Zn(OH)_2$  – zinc hydroxide, white, insoluble in water.
    - $Na_2SO_3$  – sodium sulfite, white, soluble in water.
    - $KMnO_4$  – potassium manganate(VII), purple, soluble in water.
  - All coloured compounds are eliminated, e.g. copper(II), iron(II), iron(III), nickel(II) compounds. All insoluble substances are eliminated, e.g. most carbonates and hydroxides.
    - Possible substances for **S** include:
      - All sodium, potassium and ammonium compounds.
      - All white nitrates, e.g. calcium, magnesium, aluminium, lead(II).
      - Chlorides of calcium, magnesium, aluminium.
      - Sulfates of magnesium, aluminium.

## Teaching pointers

# 18.3 How Can Ions be Identified? (page 271)

- In Chapter 15, students carried out precipitation reactions with alkalis but would probably have been unaware that these reactions can be used to identify metal cations.
- Discuss the flow charts in Figures 18.6 and 18.7 on page 273 of the Textbook to identify cations.

Demonstrate their use with

- an ammonium compound,
- a coloured solution of metal ions, e.g. iron(II) ions, and
- a colourless solution of metal ions, e.g. calcium or aluminium ions.

**Note:** These flow charts show that knowledge of whether or not a precipitate dissolves in excess alkali is not needed to identify most cations.

Exercise 18.2 of the Theory Workbook gets students to devise a flow chart to identify anions.

- Aluminium and lead(II) ions cannot be fully identified using solutions of sodium hydroxide and ammonia. To distinguish them, add potassium iodide solution to solutions of aluminium and lead(II) ions. Lead gives a yellow precipitate (refer to Table 18.2 on page 274 of the Textbook for the formation of lead(II) iodide in the test for the iodide ion).
- Barium nitrate gives a white precipitate with sulfites as well as with sulfates. However, unlike barium sulfate, barium sulfite dissolves in dilute acids to produce sulfur dioxide gas. That is the reason why barium nitrate solution is acidified with nitric acid in the test for a sulfate ion.
- Experiments 18.2 and 18.3 of the Practical Workbook involve testing for cations and anions. Exercise 18.2 involves the identification of ions in a sports drink. (See 'Notes for Teachers' on the next page on the units for the concentration of ions in drinks.)

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## Mystery Clue

From the previous clues, the sludge contains the cations aluminium  $Al^{3+}$ , calcium  $Ca^{2+}$ , lead(II)  $Pb^{2+}$  and iron(III)  $Fe^{3+}$ . To test for these cations, refer to Table 18.1 and Figures 18.6 and 18.7 in the Textbook. The aluminium and lead(II) ions would be more difficult to confirm as they cannot be distinguished using just sodium hydroxide solution or ammonia solution.

## Chemistry Inquiry (page 275)

### What Ions are Present in the Unknown Salt?

#### Group Discussion

- Refer to Note 2 on page 274 of the Textbook.
- Sulfate ion.
  - Copper, iron(II), iron(III), aluminium, calcium, lead, zinc and ammonium ions.
  - Ammonium sulfate.

#### Skills Practice (page 275)

- A green precipitate, which slowly turns brown in air.
  - A white precipitate that dissolves in excess dilute sodium hydroxide to produce a colourless solution.
- A brown precipitate.
  - A blue precipitate that dissolves in excess ammonia solution to produce a dark blue solution.
  - No reaction.
- Add excess dilute ammonia solution to some of the liquid. If it is a zinc salt, a white precipitate will form. The white precipitate then dissolves in excess ammonia solution to give a colourless solution.

## Notes for Teachers

### Concentration of ions in sports drinks

Sports drinks, like mineral water, contain several cations and anions. The labels often identify the ions present and their concentrations. The units for the concentration of ions in drinks may not be those used in this course. For example, concentration of ions in drinks may be expressed in mg/l or mEq/l (milliequivalents per litre). The latter uses 'equivalents' which are almost never taught in general Chemistry courses nowadays. However, it is still commonly used in medicine and in the pharmaceutical industry.

To convert mEq/l to mg/l:

In Chemistry, number of equivalents = number of moles  $\times$  valency (charge on ion)  
= mass/molar mass  $\times$  valency

So, mass = no. equivalents  $\times$  molar mass / valency

So, mg/L = mEq/l  $\times$  molar mass / valency

For example, take a concentration of calcium (ions) of 1.5 mEq/l.

Calcium:  $A_r = 40$

valency = 2

Concentration in mg/l =  $1.5 \times 40 / 2$   
= 30 mg/l

It is not necessary for students to know how to do these conversions.

## 18.4 How Do We Test for Gases and Water? (page 276)

### Skills Practice (page 276)

1. (i) Test: Add a piece of magnesium to the colourless liquid.

Explanation: Bubbles of hydrogen gas will be produced from the reaction of acid with magnesium.

Test the gas with a burning splint. It will burn with a 'pop' sound.

- (ii) Test: Add nitrate acid followed by barium nitrate/barium chloride solution.

Explanation: A white precipitate of barium sulfate will be produced from the reaction between sulfate and barium ions. The acid has sulfate ions and so is sulfuric acid.

2. A is sodium hydroxide.  
B is calcium hydroxide.  
C is ammonia.

### Teaching pointers

## 18.5 What is Volumetric Analysis? (page 277)

### Note on the use of Indicators

**Due to toxicity concerns with phenolphthalein, this indicator is no longer used in Volumetric Analysis experiments in the Practical Workbook. It has been replaced, in most cases, by methyl orange.**

- Show the class a bottle of vinegar again and discuss in general terms how the concentration of the (ethanoic) acid in the vinegar can be determined.
- To detect the end point of a titration reaction, indicator *solutions* are used rather than indicator paper. This is because the colour change using indicator paper is more difficult to observe. This makes it difficult to indicate that the end point has been reached throughout the entire solution.

- In the typical acid-alkali titration shown in Figure 18.12 on page 278 of the Textbook, the sodium hydroxide solution is placed in the burette. Solutions of sodium hydroxide slowly attack glass and cause the glass stopcock to be stuck. To prevent this from happening, wash the burettes thoroughly immediately after use. Alternatively, do the titration with hydrochloric acid in the burette (as in Experiment 18.6 of the Practical Workbook) or use a burette without a glass stopcock.
- Research has shown that many students solve volumetric analysis problems with little understanding of what they are doing. The research also shows that a conceptual understanding of what is happening in the experiment enhances the ability to solve numerical problems, especially the less routine questions. Therefore, teach this topic slowly and ensure that students acquire a good conceptual understanding. Further, working out the steps of a solution qualitatively *before* solving the problem quantitatively also enhances problem-solving ability.
- Avoid using relationships such as  $M_1V_1 = M_2V_2$  at this stage. This will ensure that students understand the calculations they perform rather than doing them mechanically.
- In Experiments 18.4 and 18.5 of the Practical Workbook, phenolphthalein is used as the indicator as its colour change, from pink to colourless, is simple and clear. While methyl orange can be used, it is more complicated as the colour change involves three different colours.
- The titrations in Experiments 18.8 and 18.9 of the Practical Workbook involve redox reactions. Although redox reactions are not studied until Chapter 21, students can still use the same procedure as the same steps apply to all types of titrations (acid-alkali and oxidising agent-reducing agent). These experiments can be used as an introduction to redox reactions. To help students with this kind of titration, additional notes to help with Experiment 18.9 are provided at the end of this chapter. These notes may be photocopied and distributed to the class. The question with the notes involves the reaction of iodine solution with sodium thiosulfate solution which is similar to the titration in Experiment 18.13 of the Practical Workbook.
- To complete the section on Volumetric Analysis, an optional Practical Test is included at the end of this chapter.

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**Mystery Clue**

Temperature is the main factor that would affect the speed of the reaction.

**Notes for Teachers****Cleaning glassware for titrations**

To ensure that a titration gives accurate results, the glassware must always be clean.

- Before using the pipette and burette, wash them with distilled water or deionised water to remove any impurities. Then rinse them with the solution to be used. This prevents traces of water that remain in the glassware from affecting the concentration of the solutions used.
- The conical flask should *not* be rinsed with the solution to be used as traces of solution that remain in the flask will increase the volume of solution and the amount of solute present, thus affecting the result of the titration. Therefore, wash the conical flask with water only. Traces of water do not affect the amount of solute.

### Experiment 18.14: Analysis of Brands of Vinegar on Sale in Supermarkets (PWB page 128)

This experiment is a project in which students are required to plan and carry out an analysis a commercial brand of vinegar. The time required for the experiment is estimated to be between three to four periods, with additional time needed for planning. Instead of asking students to purchase the vinegar themselves, an alternative is for you to purchase various brands of vinegar and distribute them to groups of students. Below is a suggested teaching and learning sequence of this experiment.

#### Suggested teaching and learning sequence

Lesson 1: An introduction to the project (1 period)

- Introduce the project and give a time schedule.
- Discuss the general approach with the class and the hints given in the worksheet. (Many commercial brands of vinegar are about  $1 \text{ mol/dm}^3$  solutions of ethanoic acid.)
- Discuss with students what should be included in the plan.
- Discuss how groups should write up the report of their findings.
- (Distribute the vinegar to the groups.)
- Get groups of students to begin planning for their investigation.

Preparation before lesson 2

- Get the various groups of students to complete their plan, discuss it with you and make any necessary changes.
- The various groups are to hand their apparatus and materials lists to the teacher who passes them to the laboratory technician for preparation.

Lesson 2: Performing experiments (2–3 periods)

- In the laboratory, get the various groups of students to carry out their procedures to analyse the different brands of vinegar.

Follow-up activity

Groups are to prepare a written group report.

#### Suggestions for the practical investigation

Safety precautions

Dilute sodium hydroxide solution is corrosive/an irritant. Wear plastic/rubber gloves. If the solution splashes onto the skin, wash it off immediately under running water from the tap.

Procedure

1. Note the volume of vinegar in the bottle. (Perhaps measure this volume to check whether or not it is the same as that written on the label.)
2. With a pipette, transfer  $25 \text{ cm}^3$  of the vinegar to a volumetric flask. Fill to the  $250 \text{ cm}^3$  mark with distilled/deionised water. Stopper the flask and mix.
3. Place  $25 \text{ cm}^3$  of the  $0.1 \text{ mol/dm}^3$  sodium hydroxide solution in a conical flask. Add 2 drops of phenolphthalein solution.
4. Fill a burette with the diluted vinegar and titrate until the indicator changes colour.
5. Repeat the titration until consistent results are obtained.
6. Calculate the mass of ethanoic acid in (i) the diluted solution, (ii) the initial  $25 \text{ cm}^3$  sample of vinegar and (iii) the bottle of vinegar.
7. Calculate the price per gram of ethanoic acid for the brand of vinegar.
8. Pool findings for all brands of vinegar analysed.
9. Decide which vinegar is the 'best buy'.

## Skills Practice (page 280)

1. (a)  $\text{HCl}(\text{aq}) + \text{NaOH}(\text{aq}) \longrightarrow \text{NaCl}(\text{aq}) + \text{H}_2\text{O}(\text{l})$   
 Number of moles of HCl = Number of moles of NaOH  
 $= 0.3 \div 1000 \times 12.5$   
 $= 3.75 \times 10^{-3}$  (0.00375)  
 Concentration of HCl =  $3.75 \times 10^{-3} \div 25 \times 1000$   
 $= 0.15 \text{ mol/dm}^3$
- (b) Number of moles of HCl = Number of moles of NaOH  
 $= 1.2 \div 1000 \times 20$   
 $= 0.024$   
 Concentration of HCl =  $0.024 \div 25 \times 1000$   
 $= 0.96 \text{ mol/dm}^3$
- (c) Number of moles of HCl = Number of moles of NaOH  
 $= 0.5 \div 1000 \times 35$   
 $= 0.0175$   
 Concentration of HCl =  $0.0175 \div 25 \times 1000$   
 $= 0.7 \text{ mol/dm}^3$
2. (a)  $\text{H}_2\text{SO}_4(\text{aq}) + 2\text{NaOH}(\text{aq}) \longrightarrow \text{Na}_2\text{SO}_4(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$   
 Number of moles of  $\text{H}_2\text{SO}_4 = \frac{1}{2} \times$  Number of moles of NaOH  
 $= \frac{1}{2} \times 0.3 \div 1000 \times 12.5$   
 $= 1.875 \times 10^{-3}$  (0.001875)  
 Concentration of  $\text{H}_2\text{SO}_4 = 1.875 \times 10^{-3} \div 25 \times 1000$   
 $= 0.075 \text{ mol/dm}^3$
- (b) Number of moles of  $\text{H}_2\text{SO}_4 = \frac{1}{2} \times$  Number of moles of NaOH  
 $= \frac{1}{2} \times 1.2 \div 1000 \times 20$   
 $= 0.012$   
 Concentration of  $\text{H}_2\text{SO}_4 = 0.012 \div 25 \times 1000$   
 $= 0.48 \text{ mol/dm}^3$
- (c) Number of moles of  $\text{H}_2\text{SO}_4 = \frac{1}{2} \times$  Number of moles of NaOH  
 $= \frac{1}{2} \times 0.5 \div 1000 \times 35$   
 $= 8.75 \times 10^{-3}$  (0.00875)  
 Concentration of  $\text{H}_2\text{SO}_4 = 8.75 \times 10^{-3} \div 25 \times 1000$   
 $= 0.35 \text{ mol/dm}^3$
3. (a) Number of moles of KOH =  $2 \times$  Number of moles of  $\text{H}_2\text{SO}_4$   
 $= 2 \times 0.2 \div 1000 \times 18$   
 $= 7.2 \times 10^{-3}$  (0.0072)  
 Concentration of KOH in  $\text{mol/dm}^3 = 7.2 \times 10^{-3} \div 25 \times 1000$   
 $= 0.288 \text{ mol/dm}^3$
- (b) Concentration of KOH in  $\text{g/dm}^3 = 0.288 \times (39 + 16 + 1)$   
 $= 16.1 \text{ g/dm}^3$
4. Number of moles of NaOH = Number of moles of HCl  
 $= 0.1 \div 1000 \times 22.5$   
 $= 2.25 \times 10^{-3}$  (0.00225)  
 Concentration of the impure alkali solution =  $2.25 \times 10^{-3} \div 25 \times 1000$   
 $= 0.09 \text{ mol/dm}^3$   
 Concentration of the impure alkali solution =  $0.09 \times (23 + 16 + 1)$   
 $= 3.6 \text{ g/dm}^3$   
 Percentage purity of the alkali =  $3.6 \div 4 \times 100$   
 $= 90\%$
5. Number of moles of phosphoric acid =  $\frac{1}{2} \times$  Number of moles of NaOH  
 $= \frac{1}{2} \times 0.12 \div 1000 \times 20.8$   
 $= 1.248 \times 10^{-3}$  (0.001248)  
 Concentration of phosphoric acid =  $1.248 \times 10^{-3} \div 25 \times 1000$   
 $= 0.05 \text{ mol/dm}^3$   
 Concentration of phosphoric acid =  $0.05 \times [3 + 31 + 4(16)]$   
 $= 4.9 \text{ g/dm}^3$   
 Percentage purity of the alkali =  $4.9 \div 5 \times 100$   
 $= 98\%$



# Solving the **Mystery** (page 281)

## What is in the toxic red sludge?



As part of the process at an aluminium plant, bauxite, the raw material, is taken out of the ground and washed with sodium hydroxide. This produces alumina, which is processed further, and waste, which is composed of solid impurities, heavy metals, and the chemicals used as processing agents.

About 40%–45% of the waste was iron-oxide, which gave the sludge its red colour. Another 10%–15% was aluminium oxide, a further 10%–15% silicon dioxide. There were smaller quantities of calcium oxide, titanium dioxide and sodium oxide.

Some studies have linked the inhalation of titanium dioxide and silicon dioxide dust to cancer in animals, but it has not been proven in humans. However, the quantity of these substances was far too small to cause any harm.

### Infer

The acid could also burn the skin and damage the eyes. Also neutralisation of alkalis with acids produces a lot of heat which could also burn the skin.

### Connect

A heavy metal is one with a high relative atomic mass / high density (greater than about  $5.0 \text{ g/cm}^3$ ), especially one that is poisonous. The three heavy metals of greatest public health concern are lead, mercury and cadmium. All three are toxic. Lead affects the kidneys and nervous system (especially in children). Mercury damages the brain and nervous system, causes hearing loss and birth defects. Cadmium damages the kidney, causes heart disease and perhaps cancer.

### Further Thought

Accept any reasonable points 'for' and 'against'. Students are not expected to reach any definite answers to the questions.

# 18 Chapter Review



## Self-Management

### Misconception Analysis (page 282)

- True** Both sodium hydroxide and ammonia solutions can be used to identify many, but not all, cations.
- False** Most metal ions do form precipitates with sodium hydroxide solution, but a few, such as potassium and sodium, do not.
- False** Some precipitates are soluble in excess alkali while others are not. For example, a precipitate of copper(II) hydroxide is insoluble in excess sodium hydroxide solution but is soluble in excess ammonia solution.
- False** Nitrates and carbonates are identified from gases produced in reactions.
- False** Chlorides form white precipitates with silver nitrate solutions whereas iodides form yellow precipitates. This can be used to distinguish a chloride from an iodide.
- False** Many gases, such as hydrogen, oxygen and carbon dioxide, are odourless.
- True** Note that volumetric analysis can also be used to find the concentrations of solutions of oxidising and reducing agents.

## Practice

### Structured Questions (page 282–284)

1. (a) (i) Cations: Hydrogen ions,  $H^+$ , iron(III) ions,  $Fe^{3+}$   
 (ii) Anions: Chloride ions,  $Cl^-$   
 (b) Concentrated hydrochloric acid is hazardous as it is corrosive.  
 (c) To test for hydrogen ions: Place blue litmus paper in the chemical waste. Expected result: Blue litmus paper turns red.

To test for iron(III) ions: Add sodium hydroxide solution to the chemical waste. Expected result: Reddish-brown precipitate of iron(III) hydroxide is formed. The precipitate is insoluble in excess sodium hydroxide.

To test for chloride ions: Add dilute nitric acid, then add silver nitrate solution. Expected result: A white precipitate of silver chloride is formed.

2. (a) For example, **S** could not be sodium chloride or iron(III) hydroxide. Sodium chloride is white and is soluble in water. Iron(III) hydroxide is a green salt insoluble in water but it does not give off a gas when dilute hydrochloric acid is added.  
 (b) (i) It could contain the copper(II) ions as some copper(II) compounds are green.  
 (ii) Dissolve a small sample of **S** in excess dilute hydrochloric acid. To 1  $cm^3$  of the solution, add 2 drops of aqueous sodium hydroxide. A dirty-green precipitate will indicate the present of the  $Fe^{2+}$  ion.  
 (c) (i) **G** is carbon dioxide gas.  
 (ii) Bubble the gas through limewater. A white precipitate is formed with limewater.  
 (d) **S** is probably copper(II) carbonate which is a green salt insoluble in water.
3. (a) **M** is iron.  
 (b) (i) **W** is iron(II) sulfate,  $FeSO_4$ .  
 (ii) **X** is iron(II) hydroxide,  $Fe(OH)_2$ .  
 (iii) **Y** is hydrogen,  $H_2$ .  
 (c) Put a lighted wooden splint near the gas. The gas burns with a 'pop' sound.  
 (d) **Z** is  $Fe(OH)_3$ .  
 (e)  $Fe^{2+}(aq) + 2OH^-(aq) \longrightarrow Fe(OH)_2(s)$
4. (a) (i) Iodide ion  
 (ii) Calcium ion  
 (iii) **P** is calcium iodide.  
 (b) (i) **B** is carbon dioxide gas.  
 (ii) **C** is calcium hydroxide.  
 (iii) **E** is lead(II) iodide.  
 (iv) **D** is calcium nitrate.  
 (c) **A** is calcium carbonate.

5. (a) **A** is ammonia solution.  
**B** is nitric acid.  
**C** is ammonia gas.  
**D** is ammonium nitrate solution.  
**A** must be an alkaline solution (Universal Indicator turns blue) and contains ammonium ion (gives off ammonia gas when boiled). A is therefore ammonia solution.  
**B** must contain nitrate ion since it gives off ammonia gas when warmed with dilute sodium hydroxide and aluminium foil. It must be an acid since it turns Universal Indicator red. B must be nitric acid.  
**C** must be ammonia gas since it turns damp red litmus paper blue.  
**D** is the product of the reaction between A (ammonia solution) and B (nitric acid). It must therefore be ammonium nitrate.
- (b) **B** contains  $H^+$  and  $NO_3^-$ .
6. (a) Add dilute nitric acid followed by silver nitrate solution to aqueous solutions of the two compounds. The chloride ion forms a white precipitate of silver chloride. The iodide ion forms a yellow precipitate of silver iodide. Alternatively, add lead(II) nitrate solution to the solution of iodide ions; a yellow precipitate of lead(II) iodide forms.  
 (b) Add a few drops of sodium hydroxide to solutions of the two compounds. Iron(II) sulfate solution forms a green precipitate of iron(II) hydroxide which slowly turns brown in air (due to oxidation to iron(III)). Iron(III) sulfate solution forms a brown precipitate of iron(III) hydroxide. Both precipitates are insoluble in excess sodium hydroxide.  
 (c) Add a few drops of sodium hydroxide to solutions of the two compounds. Calcium nitrate solution forms a white precipitate of calcium hydroxide, *insoluble* in excess sodium hydroxide. Aluminium nitrate solution forms a white precipitate *soluble* in excess sodium hydroxide.  
 (d) Add sodium hydroxide (solid or aqueous solution) to the compounds. On warming, ammonium chloride produces ammonia gas (which turns red litmus paper blue). There is no reaction with sodium chloride.
7. (a) A suitable indicator is phenolphthalein. It changes colour from pink to colourless.

(b)

Titration	Initial burette reading ( $cm^3$ )	Final burette reading ( $cm^3$ )	Volume of acid added ( $cm^3$ )
1	3.0	25.8	22.8
2	25.8	47.3	21.5
3	3.6	25.1	21.5

- (c) The first reading should be discarded. The reading is too high, probably because in the first titration, too much acid was added and the end point was overshot.  
 (d) 21.5  $cm^3$  of acid is required for neutralisation.

- (e) Number of moles of sodium hydroxide  
 = Number of moles of hydrochloric acid  
 =  $0.1 \div 1000 \times 21.5$   
 =  $2.15 \times 10^{-3}$

Concentration of sodium hydroxide  
 =  $2.15 \times 10^{-3} \div 25 \times 1000$   
 =  $0.086 \text{ mol/dm}^3$

8. (a) Concentration of NaOH =  $6 \div (23 + 16 + 1)$   
 =  $0.15 \text{ mol/dm}^3$
- (b) Number of moles of NaOH used in the reaction  
 =  $0.15 \div 1000 \times 20$   
 =  $0.003 \text{ mol}$
- (c) Number of moles of  $\text{HNO}_3$  = Number of moles of NaOH  
 =  $0.003 \text{ mol}$
- (d) (i) Concentration of  $\text{HNO}_3$  in  $\text{mol/dm}^3$   
 =  $0.003 \div 4 \times 1000$   
 =  $0.75 \text{ mol/dm}^3$
- (ii) Concentration of  $\text{HNO}_3$  in  $\text{g/dm}^3$   
 =  $0.75 \times [1 + 14 + 3(16)]$   
 =  $47.25 \text{ g/dm}^3$
9. (a)  $\text{CH}_3\text{COOH}(\text{aq}) + \text{NaOH}(\text{aq}) \longrightarrow \text{CH}_3\text{COO}^-\text{Na}^+(\text{aq}) + \text{H}_2\text{O}(\text{l})$
- (b) (i) Number of moles of  $\text{CH}_3\text{COOH}$   
 = Number of moles of NaOH  
 =  $0.05 \div 1000 \times 30$   
 =  $1.5 \times 10^{-3}$  (0.0015)  
 Concentration of NaOH in  $\text{mol/dm}^3$   
 =  $1.5 \times 10^{-3} \div 25 \times 1000$   
 =  $0.06 \text{ mol/dm}^3$
- (ii) Concentration of NaOH in  $\text{g/dm}^3$   
 =  $0.06 \times (23 + 16 + 1)$   
 =  $2.4 \text{ g/dm}^3$
10. (a) Pipette  
 (b) Volumetric flask  
 (c)  $\text{H}_2\text{SO}_4(\text{aq}) + 2\text{NaOH}(\text{aq}) \longrightarrow \text{Na}_2\text{SO}_4(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$
- (d) (i) Number of moles of NaOH  
 =  $2 \times$  Number of moles of  $\text{H}_2\text{SO}_4$   
 =  $2 \times 0.01 \div 1000 \times 12.8$   
 =  $2.56 \times 10^{-4}$   
 Concentration of diluted NaOH  
 =  $2.56 \times 10^{-4} \div 25 \times 1000$   
 =  $0.01 \text{ mol/dm}^3$
- (ii) Since the oven cleaner is diluted 10 times, the concentration of the NaOH solution is 10 times more concentrated, i.e.  $0.01 \times 10 = 0.10 \text{ mol/dm}^3$
- (e) As the original solution is concentrated, an unnecessarily large volume ( $250 \text{ cm}^3$ ) of the  $0.01 \text{ mol/dm}^3$  sulfuric acid would be needed in the titration.
11. (a)  $\text{K}_2\text{CO}_3(\text{aq}) + 2\text{HCl}(\text{aq}) \longrightarrow 2\text{KCl}(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$
- (b) Number of moles of  $\text{K}_2\text{CO}_3$  in  $25 \text{ cm}^3$   
 =  $\frac{1}{2} \times$  Number of moles of HCl  
 =  $\frac{1}{2} \times 0.1 \div 1000 \times 30$   
 =  $1.5 \times 10^{-3}$   
 Number of moles of  $\text{K}_2\text{CO}_3$  in  $100 \text{ cm}^3$  =  $1.5 \times 10^{-3} \times 4$   
 =  $6 \times 10^{-3}$

Mass of  $\text{K}_2\text{CO}_3$  in 2.5 g of potash  
 =  $6 \times 10^{-3} \times [2(39) + 12 + 3(16)]$   
 =  $0.828 \text{ g}$

(c) Percentage purity of potash =  $0.828 \div 2.5 \times 100$   
 =  $33.12\%$

### Free Response Questions (page 284)

1. Responses to this question may include the following points:
- To separate ammonium sulfate from the mixture: Add the mixture to water to dissolve the ammonium sulfate. Filter the mixture to remove the insoluble sand. Heat the filtrate slowly until crystals of ammonium sulfate appear. Leave the hot saturated solution to cool. Filter the crystals and wash them with a little water.
  - To test for the presence of ammonium ions: Add a little solid sodium hydroxide to the mixture in a test tube and heat gently. Ammonia is produced:  

$$(\text{NH}_4)_2\text{SO}_4(\text{s}) + 2\text{NaOH}(\text{s}) \longrightarrow \text{Na}_2\text{SO}_4(\text{s}) + 2\text{NH}_3(\text{g}) + 2\text{H}_2\text{O}(\text{l})$$
 Hold a piece of moist red litmus paper at the mouth of the test tube. The red litmus paper turns blue.
  - To test for the presence of sulfate ions: Dissolve a few of the crystals in water. Add dilute nitric acid followed by aqueous barium nitrate. The formation of a white precipitate of barium sulfate identifies a sulfate ion.  

$$\text{Ba}(\text{NO}_3)_2(\text{aq}) + (\text{NH}_4)_2\text{SO}_4(\text{aq}) \longrightarrow \text{BaSO}_4(\text{s}) + 2\text{NH}_4\text{NO}_3(\text{aq})$$
2. Responses to this question may include the following points:
- Use a pipette and a pipette filler to transfer  $25.0 \text{ cm}^3$  sodium hydroxide solution to a conical/titration flask.
  - Add a few drops of a suitable indicator to the flask, e.g. phenolphthalein.
  - Fill the burette with the sulfuric acid. Record the initial volume.
  - Add the acid to the alkali. The indicator changes colour at the end point, e.g. from pink to colourless for phenolphthalein. Record the final burette reading and the volume of acid used.
  - Repeat the titration several times to obtain consistent results.
  - Use the best results to obtain the average volume of acid required. Then calculate the concentration of the sulfuric acid.

## Extension (page 284)

### 1. What is the Question?

(a) Possible questions:

- What will you observe when hydrochloric acid is added to marble chips?
- What will you observe when sulfuric acid is added to sodium carbonate solution?

(b) Possible questions:

- A solution of salt X gives a yellow precipitate with aqueous potassium iodide and ammonia gas when boiled with aluminium foil and aqueous sodium hydroxide. What is salt X?
- Name one salt that has the following properties:
  - A solution of the salt gives a white precipitate with dilute sodium hydroxide and the precipitate dissolves in excess sodium hydroxide.
  - A solution of the salt gives an alkaline gas when heated with aluminium foil and excess dilute sodium hydroxide.
- Name a nitrate salt whose solution gives a white precipitate with dilute ammonia solution and the precipitate does not dissolve in excess sodium hydroxide.

### 2. Titration Calculations with a Spreadsheet

It is assumed that students have had some experience with spreadsheet exercises. Thus they have an understanding of how to write and enter simple formulae into the spreadsheet cells.

#### Titration of sulfuric acid with sodium hydroxide

Using Microsoft Excel, the spreadsheet should look as follows:

A	B	C	D	E	F	G	H
Vol. NaOH	Vol. NaOH	Conc. NaOH	Moles NaOH	Moles H <sub>2</sub> SO <sub>4</sub>	Vol. H <sub>2</sub> SO <sub>4</sub>	Vol. H <sub>2</sub> SO <sub>4</sub>	Conc. H <sub>2</sub> SO <sub>4</sub>
cm <sup>3</sup>	dm <sup>3</sup>	mol/dm <sup>3</sup>	moles	moles	cm <sup>3</sup>	dm <sup>3</sup>	mol/dm <sup>3</sup>
30	0.03	0.1	0.003	0.0015	25	0.025	0.06
15	0.015	0.4	0.006	0.003	25	0.025	0.12
24.5	0.0245	0.12	0.00294	0.00147	20	0.02	0.0735

# Additional Teaching Material



## Additional Exercise 1: Analysis of Chemical Waste

### Objective

- To suggest tests for ions present in a mixture of waste chemicals

### Key Competency

**CIT:** sound reasoning [*predicting*]

A barrel of chemical waste has the following label on it.

**Warning**  
**Hazardous Chemicals**

This drum contains the following:

nitric acid  
 iron(III) chloride  
 sulfuric acid

- You have to check the contents of this barrel by carrying out a test for each ion that is believed to be in the waste. Write out a plan of the tests you would carry out by filling in the table below.

Test number	Name and formula of ion	Test to be carried out	Expected result if the test is positive
1			
2			
3			
4			
5			

- Suggest a suitable chemical to add to the barrel that would neutralise the contents so that it could be safely disposed of.

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# Additional Teaching Material

## Additional Exercise 2: Other Titrations

### Objective

- To calculate the concentration of iodine solution used

### Key Competency

**CIT:** sound reasoning [*calculating*]

Most of the titrations you will carry out involve acids and alkalis. However, titrations can also be carried out with solutions containing substances other than acids and alkalis. The method of calculation is always the same. The following steps below show how a titration can be used to find the concentration of a solution of iron(II) sulfate,  $\text{FeSO}_4$ , using a standard solution of potassium manganate(VII),  $\text{KMnO}_4$ .

Titrations can also be carried out with solutions containing substances other than acids and alkalis. The method of calculation is always the same. The following steps below show how a titration can be used to find the concentration of a solution of iron(II) sulfate,  $\text{FeSO}_4$ , using a standard solution of potassium manganate(VII),  $\text{KMnO}_4$  (Figure 1).

Typical results using the above procedure are given in the example after Figure 1.

using a pipette filler

$\text{FeSO}_4(\text{aq})$

1. Pipette  $25.0 \text{ cm}^3$  of  $\text{FeSO}_4(\text{aq})$  into a titration flask.

dilute  $\text{H}_2\text{SO}_4$

2. Add approximately  $20 \text{ cm}^3$  of dilute sulfuric acid to the titration flask [this is to enable  $\text{KMnO}_4(\text{aq})$  to react with  $\text{FeSO}_4(\text{aq})$ ].

$\text{KMnO}_4(\text{aq})$

3. Fill the burette with the standard  $\text{KMnO}_4(\text{aq})$ . Then run it into the titration flask. The  $\text{KMnO}_4(\text{aq})$  is purple in colour and is immediately decolourised by the  $\text{FeSO}_4(\text{aq})$ . As  $\text{KMnO}_4$  is run in, the purple colour disappears until, at the end point, one drop of  $\text{KMnO}_4$  causes a permanent pink colour. This completes the titration. Record the volume of  $\text{KMnO}_4(\text{aq})$  used.

**Figure 1** Titration of iron(II) sulfate with potassium manganate(VII)

### Example

$25.0 \text{ cm}^3$  of  $\text{FeSO}_4(\text{aq})$ , acidified with sulfuric acid, required  $27.5 \text{ cm}^3$  of  $0.020 \text{ mol/dm}^3$   $\text{KMnO}_4(\text{aq})$  for reaction in a titration. Calculate the concentration of the  $\text{FeSO}_4(\text{aq})$ .

### Solution

Start with the standard solution,  $\text{KMnO}_4$ .

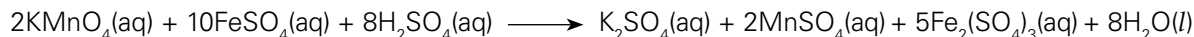
**Step 1**

Find the number of moles of  $\text{KMnO}_4$  used in the titration.

$$\begin{aligned}\text{Number of moles of } \text{KMnO}_4 &= \text{Concentration} \times \text{Volume in dm}^3 \\ &= 0.020 \times \frac{27.5}{1000} \text{ mol}\end{aligned}$$

**Step 2**

Write the chemical equation for the reaction.



(You would always be given this equation or be told how many moles of  $\text{FeSO}_4$  react with 1 mole of  $\text{KMnO}_4$ , which you need for the next step.)

**Step 3**

From the equation, find the ratio of the number of moles of  $\text{FeSO}_4$  to the number of moles of  $\text{KMnO}_4$ .

$$\frac{\text{Number of moles of } \text{FeSO}_4}{\text{Number of moles of } \text{KMnO}_4} = \frac{10}{2} = \frac{5}{1}$$

**Step 4**

Use the ratio to find the number of moles of  $\text{FeSO}_4$  that reacted in the titration.

$$\begin{aligned}\text{Number of moles of } \text{FeSO}_4 &= 5 \times \text{Number of moles of } \text{KMnO}_4 \\ &= 5 \times 0.020 \times \frac{27.5}{1000} \text{ mol}\end{aligned}$$

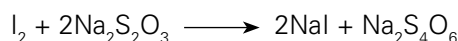
**Step 5**

Find the concentration of  $\text{FeSO}_4(\text{aq})$  in  $\text{mol/dm}^3$ .

$$\begin{aligned}\text{Concentration} &= \frac{\text{Number of moles of } \text{FeSO}_4}{\text{Volume of } \text{FeSO}_4 \text{ used in dm}^3} \\ &= (5 \times 0.020 \times \frac{27.5}{1000} \text{ mol}) \div \frac{25.0}{1000} \\ &= 0.11 \text{ mol/dm}^3\end{aligned}$$

**Question**

Iodine solution reacts with sodium thiosulfate solution,  $\text{Na}_2\text{S}_2\text{O}_3(\text{aq})$ , according to the equation:



In a titration,  $25.0 \text{ dm}^3$  of iodine solution reacted with  $32.0 \text{ cm}^3$  of  $0.12 \text{ mol/dm}^3$  sodium thiosulfate solution.

Calculate the concentration of the iodine solution in

**(a)** moles of  $\text{I}_2$  per  $\text{dm}^3$ , and

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**(b)** grams of  $\text{I}_2$  per  $\text{dm}^3$ .

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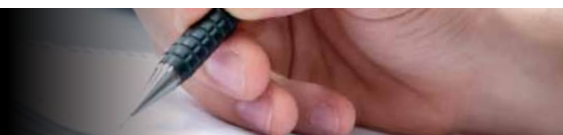


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# Additional Teaching Material



## Practical Test (50 minutes)

### Apparatus and materials

- titration flask (250 cm<sup>3</sup>)
- pipette (25 cm<sup>3</sup>)
- pipette filler
- burette (50 cm<sup>3</sup>)
- glass rod
- filter funnel
- wash bottle
- retort stand, boss and clamp
- white tile
- Solution **P**: hydrogen peroxide (unknown concentration)
- Solution **Q**: sodium thiosulfate (0.100 mol/dm<sup>3</sup>)
- dilute sulfuric acid
- aqueous potassium iodide
- starch indicator
- Solid **U**
- usual reagents for qualitative analysis
- concentrated hydrochloric acid

### Safety warnings



- dilute sulfuric acid
- solution **P**, solution **Q**, solid **U**

1. You are given a Solution **P**, which contains hydrogen peroxide. It was prepared by adding dilute sulfuric acid to 8.50 g of barium peroxide, filtering off the insoluble barium sulfate and diluting the resulting solution to 1.00 dm<sup>3</sup>.



The concentration of hydrogen peroxide in **P** can be determined by adding acidified aqueous potassium iodide and titrating the liberated iodine with aqueous sodium thiosulfate.

You are to determine the concentration of hydrogen peroxide in **P** and use this to determine the relative formula mass of barium peroxide.

Solution **Q** is 0.100 mol/dm<sup>3</sup> sodium thiosulfate.

### Procedure

- (a) Put solution **Q** into the burette.
- (b) Pipette a 25.0 cm<sup>3</sup> of solution **P** into a flask and add about a test-tubeful of dilute sulfuric acid followed by about a test-tubeful of aqueous potassium iodide. The solution should turn red-brown. Do not add the starch indicator at this stage.
- (c) Add solution **Q** from the burette until the red-brown colour fades to pale yellow, then add a few drops of the starch indicator. This will give a dark blue solution.
- (d) Continue adding solution **Q** slowly from the burette until one drop of **Q** causes the blue colour to disappear, leaving a colourless solution. Record your results in the table.
- (e) Repeat the titration as many a number of times to achieve consistent results.



## Results

Burette readings	1	2	3
Titration number			
Final reading (cm <sup>3</sup> )			
Initial reading (cm <sup>3</sup> )			
Volume of <b>Q</b> used (cm <sup>3</sup> )			
Best titration results (✓)			

## Summary

Tick (✓) the best titration results.

Using these results, the average volume of solution **Q** required was \_\_\_\_\_ cm<sup>3</sup>.

Volume of solution **P** used was \_\_\_\_\_ cm<sup>3</sup>.

## Calculations

(a) **Q** is 0.100 mol/dm<sup>3</sup> sodium thiosulfate.

One mole of hydrogen peroxide reacts with potassium iodide to produce iodine.

The iodine produced reacts with two moles of sodium thiosulfate.

Calculate the concentration, in mol/dm<sup>3</sup>, of hydrogen peroxide in **P**.

Concentration of hydrogen peroxide in **P** is \_\_\_\_\_ mol/dm<sup>3</sup>.

(b) Solution **P** was prepared by adding dilute sulfuric acid to 8.50 g of barium peroxide and diluting the resulting solution to 1.00 dm<sup>3</sup>.

Using your answer in Step 1, calculate the relative formula mass of barium peroxide.

Relative formula mass of barium peroxide is \_\_\_\_\_.

2. Carry out the following experiments on solid **U**, recording your observations in the table. Identify any gases produced. You are not expected to identify solid **U**.

Test no.	Test	Observations
1	Heat a portion of <b>U</b> in a hard glass test tube.	
2	(a) Heat a portion of <b>U</b> with excess dilute nitric acid in a small beaker. Boil the mixture, with stirring, for a few seconds.	
	(b) Filter the mixture from (a). Use the filtrate for Test no. 3 and the residue for Test no. 4.	
3	(a) To a portion of the filtrate from Test no. 2(b) add aqueous potassium iodide.	
	(b) To a portion of the filtrate from Test no. 2(b) add dilute hydrochloric acid.	
	(c) To a portion of filtrate from Test no. 2(b) add an equal volume of dilute sodium hydroxide.  Then add dilute sodium hydroxide until there is no further change.	
4	Wash the residue from Test no. 2(b) with a little water.	
	(a) Place a portion of the residue into a test tube. Add aqueous potassium iodide and a little dilute sulfuric acid. Cork the test tube and shake it vigorously. Allow the mixture to settle.	
	(b) Place a portion of the residue from Test no. 2(b) into a test tube. Add <i>a little concentrated</i> hydrochloric acid.	

[19]

## Conclusions

- Name one element that must be present in **U**. \_\_\_\_\_  
Give the Test no. that shows this. \_\_\_\_\_
- What deduction can be made from the result of Test no. 4(b) about the nature of the residue produced in Test no. 2?

# Answers

## Additional Exercise 1:

Test number	Name and formula of ion	Test to be carried out	Expected result if the test is positive
1	hydrogen ion, $H^+$	add Universal Indicator	red colour due to the presence of acids
2	iron(III) ion, $Fe^{3+}$	add excess dilute sodium hydroxide	brown precipitate
3	nitrate ion, $NO_3^-$	add dilute sodium hydroxide and then add aluminium powder and warm	ammonia gas evolved turns damp red litmus paper blue
4	chloride ion, $Cl^-$	add silver nitrate solution (with nitric acid)	white precipitate
5	sulfate ion, $SO_4^{2-}$	add barium nitrate solution (with nitric acid)	white precipitate

2. Slaked lime or calcium carbonate

## Additional Exercise 2:

(a) Number of moles of  $I_2 = \frac{1}{2} \times$  Number of moles of  $Na_2S_2O_3$   
 $= \frac{1}{2} \times 0.12 \div 1000 \times 32$   
 $= 1.92 \times 10^{-3}$  (0.00192)

Concentration of  $I_2$  in  $mol/dm^3 = 1.92 \times 10^{-3} \div 25 \times 1000$   
 $= 0.0768 mol/dm^3$

(b) Concentration of  $I_2$  in  $g/dm^3 = 0.0768 \times (127 + 127)$   
 $= 19.5 g/dm^3$

## Practical Test:

Observations
1. Oxygen gas is produced, which relights a glowing splint. The solid changes colour from orange to brown.
2. (a) The orange solid turns dark brown. (b) The filtrate is colourless. The residue is dark brown.
3. (a) A yellow ppt is produced. (b) A white ppt is produced. (c) A white ppt is produced. The ppt dissolves to produce a colourless solution.
4. (a) The potassium iodide solution turns down. (b) A pale yellow gas is produced. Chlorine gas is produced, which bleaches damp blue litmus paper. A yellow or brown solid is left.

## Conclusions

- Oxygen, 1
- It is an oxidising agent.

**Notes:** **S** is  $0.1250 mol/dm^3$  potassium manganate(VII); **T** is  $0.0600 mol/dm^3 H_2O_2$  in  $1 mol/dm^3 H_2SO_4$ ; **U** is  $Pb_3O_4$  ('red lead').

The titration results must give the correct 5 : 2 ratio. To ensure the correct answer is obtained, carry out the experiment in advance. Assuming **S** is  $0.025 mol/dm^3$ , calculate the concentration of **T**. This value can be used instead of  $0.0600 mol/dm^3$ .

**Blank**