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Executive Preview

Chemistry

for Cambridge International AS & A Level

MULTI-COMPONENT SAMPLE

Lawrie Ryan & Roger Norris

Third edition

Digital Access

Brighter Thinking
Better Learning

At Cambridge University Press, we put you at the heart of our teaching and learning resources. This new series has been developed using extensive research with our exclusive teacher community (the Cambridge Science Panel), as well as teacher interviews and lesson observations around the world. It meets the real needs that we have discovered in our research – solving and supporting the biggest classroom challenges that you have told us about. We want to help you deliver engaging lessons that use the best practical pedagogies to enable your students to achieve their learning goals. In essence, we want to make your teaching time easier and more effective.

At the heart of this new series, our completely revised and expanded teacher's resource helps you to use each of the resources in the series effectively. This includes teaching activity, assessment and homework ideas, suggestions on how to tackle common misconceptions, and support with running practical activities. This resource will inspire and support you while saving much-needed time.

For this new edition of the coursebook, we have added new features. These include reflection opportunities and self-evaluation checklists that develop responsible learners, a broader range of enquiry questions that support practical activities, as well as group work and debate questions that develop 21st century skills. The 'Science in Context' features now include open-ended discussion questions that enable students to practise their English skills, interpret ideas in a real-world context and debate concepts with other learners. There is also extra support to help English as a second language learners successfully engage with their learning (including improved and expanded support for learning the all-important scientific vocabulary) alongside simple definitions of key terms and command words. Active lesson ideas and multi-part exam-style questions ensure student engagement and helps them feel confident approaching assessment.

The workbook is the perfect companion for the coursebook. You can use it to reinforce learning, promote application of theory and help students practise the essential skills of handling data, evaluating information and problem solving. The workbook now includes frequent tips to support students' understanding, alongside a range of formative exercises that map directly onto, and build on, coursebook topics and concepts. Multi-part exam-style questions also provide students with practice in a familiar format.

To support the syllabus focus on practical work and the scientific method, the practical workbook contains step-by-step guided investigations and practice questions. These give students the chance to test their knowledge and help build confidence in preparation for assessment. Practical investigation helps to develop key skills – such as planning, identifying equipment, creating hypotheses, recording results, and analysing and evaluating data. This workbook is ideal for teachers who find running practical experiments difficult due to lack of time, resources or support. It contains help and guidance on setting up and running practical investigations in the classroom, as well as sample data for when students can't do the experiments themselves.

We're very pleased to share with you draft chapters from our forthcoming coursebook, teacher's resource, workbook and practical workbook. We hope you enjoy looking through them and considering how they will support you and your students.

If you would like more information or have any questions, please contact your local sales representative: **cambridge.org/education/find-your-sales-consultant**

Steve Temblett
Head of Publishing – Science, Technology & Maths, Cambridge University Press

Hello. I am Roger Norris. I am part of the team involved in writing the books in this series which supports the revised Cambridge International AS & A Level Chemistry syllabus (9701). I am an experienced teacher, author and examiner. I am very pleased to be able to introduce myself to you and give you some information about the content in the new series. As you may be aware, there have been some revisions to the syllabus for first examination in 2022. You will find the full syllabus document online at cambridgeinternational.org.

The series has four components – a coursebook, workbook, practical workbook and teacher resource. We have made sure that they work together to give you and your students full support in every aspect of the Cambridge International AS & A Level Chemistry course.

- We have revised the **coursebook** so that it covers all of the learning objectives in the syllabus. We have reviewed the language level, to make it more accessible for students whose first language is not English.
- Each chapter begins with a context, to stimulate discussion. Within the text, questions encourage students to deepen their understanding of the topics covered. At the end of chapters, there are exam-style questions to build learner confidence.
- The number of technical terms in science can be challenging. We have fully explained these when they first appear. They are highlighted in 'Key Words' boxes and can also be found in the glossary. Summaries, a self-evaluation table and a reflection feature, encourage learners to reflect and improve.
- The **workbook** helps learners to develop the many *skills* that they need in order to prepare for examination questions. These include the Assessment Objective 2 skills and some of the skills that are used in practical work (AO3).
- Practical work can be a challenge, but we also know its importance to help learners reach their full potential. We have therefore provided a **practical workbook** to give really detailed guidance in doing practical work. We have trialed all of the experiments in a school laboratory, and provide comprehensive step-by-step instructions.
- We have completely revised the **teacher's resource**, to ensure that it provides the teacher with extensive support for all aspects of the course.

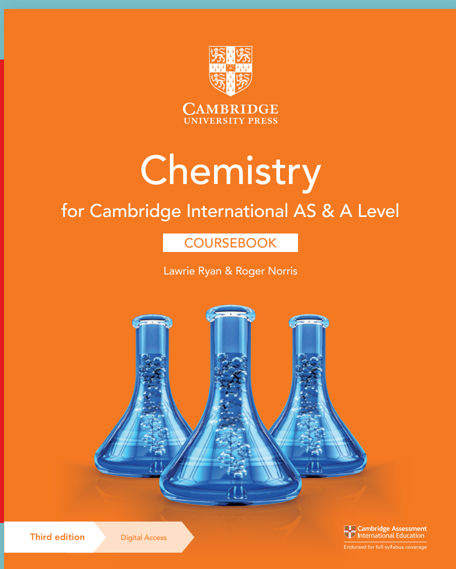
All the authors for this series are experienced teachers of chemistry. I sincerely hope that you and your students will enjoy using these new editions and wish you every success.

Kind regards,

Roger Norris

How to use this series

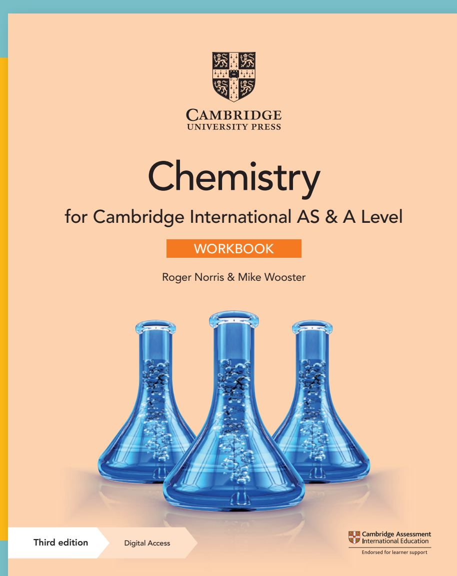
This suite of resources supports students and teachers following the Cambridge International AS & A Level Chemistry syllabus (9701). All of the books in the series work together to help students develop the necessary knowledge and scientific skills required for this subject.

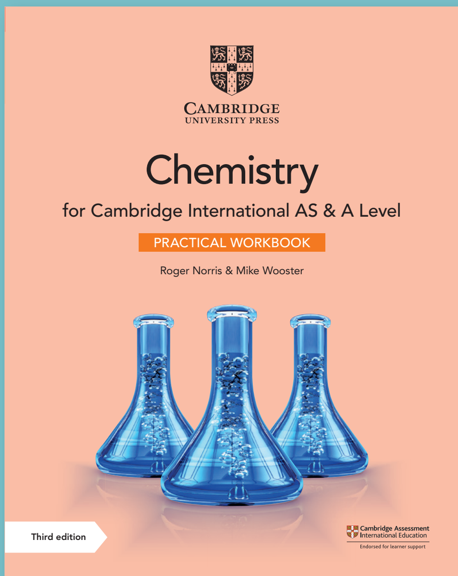


The coursebook provides comprehensive support for the full Cambridge International AS & A Level Chemistry syllabus (9701). It clearly explains facts, concepts and practical techniques, and uses real world examples of scientific principles. Two chapters provide full guidance to help learners develop investigative skills. Questions within each chapter help them to develop their understanding, while exam-style questions provide essential practice.

The workbook contains over 100 exercises and exam-style questions, carefully constructed to help learners develop the skills that they need as they progress through their Chemistry course.

The exercises also help learners develop understanding of the meaning of various command words used in questions, and provide practice in responding appropriately to these.

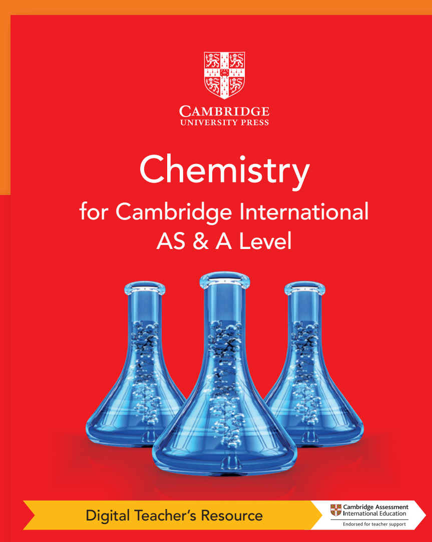


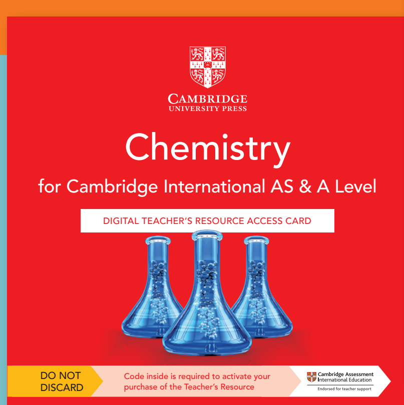


This write-in book provides learners with a wealth of hands-on practical work, giving them full guidance and support that will help them to develop all of the essential investigative skills. These skills include planning investigations, selecting and handling apparatus, creating hypotheses, recording and displaying results, and analysing and evaluating data.

The teacher's resource supports and enhances the questions and practical activities in the coursebook. This resource includes detailed lesson ideas, as well as answers and exemplar data for all questions and activities in the coursebook and workbook.

The practical teacher's guide, included with this resource, provides support for the practical activities and experiments in the practical workbook. Teaching notes for each topic area include a suggested teaching plan, ideas for active learning and formative assessment, links to resources, ideas for lesson starters and plenaries, differentiation, lists of common misconceptions and suggestions for homework activities. Answers are included for every question and exercise in the coursebook, workbook and practical workbook.
Detailed support is provided for preparing and carrying out for all the investigations in the practical workbook, including tips for getting things to work well, and a set of sample results that can be used if learners cannot do the experiment, or fail to collect results.





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Chemistry

for Cambridge International AS & A Level

COURSEBOOK

Lawrie Ryan & Roger Norris

Third edition

Digital Access

Cambridge Assessment
International Education

Endorsed for full syllabus coverage

Contents

How to use this book

Throughout this book, you will notice lots of different features that will help your learning. These are explained below.

LEARNING INTENTIONS

These set the scene for each chapter, help with navigation through the coursebook and indicate the important concepts in each topic.

BEFORE YOU START

This contains questions and activities on subject knowledge you will need before starting this chapter.

SCIENCE IN CONTEXT

This feature presents real-world examples and applications of the content in a chapter, encouraging you to look further into topics. There are discussion questions at the end which look at some of the benefits and problems of these applications.

PRACTICAL ACTIVITIES

This book does not contain detailed instructions for doing particular experiments, but you will find background information about the practical work you need to do in these boxes. There are also two chapters, P1 and P2, which provide detailed information about the practical skills you need to develop during the course.

Questions

Appearing throughout the text, questions give you a chance to check that you have understood the topic you have just read about. You can find the answers to these questions in the Digital version of the coursebook.

KEY DEFINITION

Key definitions for important scientific principles, laws and theories are given in the margin and highlighted in the text when it is first introduced. You will also find these definitions in the Glossary at the back of this book.

KEY WORDS

Key vocabulary is highlighted in the text when it is first introduced. Definitions are then given in the margin, which explain the meanings of these words and phrases.

You will also find definitions of these words in the Glossary at the back of this book.

COMMAND WORDS

Command words that appear in the syllabus and might be used in exams are highlighted in the exam-style questions when they are first introduced. In the margin, you will find the Cambridge International definition*.

You will also find the same definitions in the Glossary at the back of this book.

*The information in this section is taken from the Cambridge International syllabus for examination from 2022. You should always refer to the appropriate syllabus document for the year of your examination to confirm the details and for more information. The syllabus document is available on the Cambridge International website at www. cambridgeinternational.org.

WORKED EXAMPLES

Wherever you need to know how to use a formula to carry out a calculation, there are worked examples boxes to show you how to do this.

REFLECTION

These activities ask you to look back on the topics covered in the chapter and test how well you understand these topics and encourage you to reflect on your learning.

IMPORTANT

Important equations, facts and tips are given in these boxes.

SUMMARY CHECKLISTS

There is a summary of key points at the end of each chapter.

EXAM-STYLE QUESTIONS

These questions provide more demanding exam-style questions, some of which may require use of knowledge from previous chapters. Answers to these questions can be found in the Elevate edition of the coursebook.

SELF-EVALUATION CHECKLIST

There are 'I can' statements at the end of each chapter which match the learning intentions at the beginning of the chapter. You might find it helpful to rate how confident you are for each of these statements when you are revising. You should revisit any sections that you rated "Needs more work" or "Almost there".

I can	See section...	Needs more work	Almost there	Ready to move on

Chapter 1
Atomic structure

LEARNING INTENTIONS

In this chapter you will learn how to:

- describe the structure of the atom as mostly empty space surrounding a very small nucleus that consists of protons and neutrons and state that electrons are found in shells in the space around the nucleus
- describe the position of the electrons in shells in the space around the nucleus
- identify and describe protons, neutrons and electrons in terms of their relative charges and relative masses
- use and understand the terms atomic (proton) number and mass (nucleon) number
- describe the distribution of mass and charges within an atom
- deduce the behaviour of beams of protons, neutrons and electrons moving at the same velocity in an electric field
- understand that ions are formed from atoms or molecules by gain or loss of electrons
- deduce the numbers of protons, neutrons and electrons present in both atoms and ions given atomic (proton) number, mass (nucleon) number and charge
- define the term *isotope* in terms of numbers of protons and neutrons

CONTINUED

- use the notation $^{x}_{y}A$ for isotopes, where ^{x} is the mass (nucleon) number and $_{y}$ is the atomic (proton) number
- explain why isotopes of the same element have the same chemical properties
- explain why isotopes of the same element have different physical properties (limited to mass and density).

BEFORE YOU START

1 Without looking at the Periodic Table, make a list of the names and symbols for the elements in Periods 1, 2 and 3. Compare your list with another learner then check to see if the symbols are correct.

2 How can you deduce the formula of a simple ion (e.g. a chloride ion or an aluminium ion) by reference to the Periodic Table?

3 Take turns in challenging another learner to write down the formula of a simple ion. Check your answers afterwards using a textbook.

4 Make a list of the subatomic particles in an atom giving their relative mass and relative charges as well as their position in the atom, structure of the atom and isotopes. Compare your answers with those of another learner. Were you in agreement?

5 Write down a definition of the term *isotope*. Put a circle around the three most important words in your definition. Compare your definition to the one in a textbook.

6 What do the terms *mass number* and *proton number* mean? Write down your definitions and compare yours with another learner.

7 Ask another learner to use a data book or the internet to select an isotope. Use this data to deduce the number of protons, neutrons and electrons in an atom or ion of this isotope, e.g. Cr atom or Cr^{3+} ion. If you are unsure, check your answer with someone else in the class or with a teacher.

8 Take a photocopy of the modern Periodic Table and cross out or cut out the group numbers and period numbers. Get another learner to select an element. You then have to state in which period and group that element belongs. Take turns in doing this until you are sure that you can easily identify the group and period of an element.

9 Ask another learner to select an element. You then have to state if the element is a metal, non-metal or metalloid (metalloids have some characteristics of both metals and non-metals). If you are both uncertain, consult a textbook or the internet. Take turns in doing this until you are sure that you can easily identify the position of metals, non-metals and metalloids.

10 Explain to another learner in terms of numbers of electrons and protons why a sodium ion has a single positive charge but an oxide ion has a 2– charge.

11 Explain to another learner what you know about attraction or repulsion of positive and negative charges.

DEVELOPING AN IDEA: NANOMACHINES

Progress in science depends not only on original thinking but also on developing the ideas of other people. The idea of an atom goes back over 2000 years to the Greek philosopher Demokritos. About 350 years ago, Robert Boyle looked again at the idea of small particles but there was no proof. John Dalton moved a step closer to proving that atoms exist: he developed the idea that atoms of the same kind had the same weight, thinking this could explain the results of experiments on combining different substances in terms of rearrangement of the atoms.

At the beginning of the 20th century, J.J. Thomson (see Figure 1.6) suggested three models of the atom. His preferred model was to imagine an atom as a spherical cloud of positive charge in which electrons were placed. A few years later, scientists working under the direction of Ernest Rutherford (see Figure 1.4) fired alpha particles (which we now know are positively charged helium nuclei) at very high speeds at strips of metal only 0.0005 mm thick. Most of the alpha particles went through the strip. This would fit with the idea of atoms being a cloud of charge with very little mass to deflect (change the direction of) the alpha particles. But one alpha particle in every 20 000 was deflected at an angle of more than 90°. From this, Rutherford deduced that there must be something very small and positively charged in the atom. The atomic nucleus had been discovered!

Figure 1.1: Richard Feynman.

In 1960 Richard Feynman (Figure 1.1) suggested that tiny machines could be made from a few hundred atoms grouped together in clusters. At the time, these ideas seemed like 'science fiction'. But several scientists took up the challenge and the science of nanotechnology was born.

In nanotechnology, scientists design and make objects that may have a thickness of only a few thousand atoms or less. Groups of atoms can be moved around on special surfaces (Figure 1.2). In this way, scientists have started to develop tiny machines that will help deliver medical drugs to exactly where they are needed in the body.

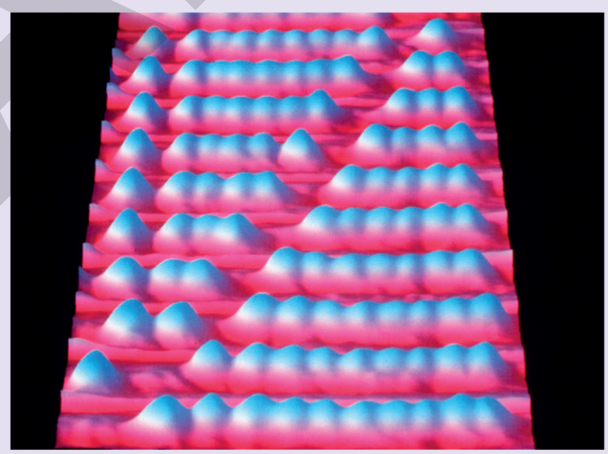

Figure 1.2: Each of the blue peaks in this image is an individual molecule. The molecules can be moved over a copper surface, making this a molecular abacus or counting device.

Questions for discussion

Discuss with another learner or group of learners:

- Why do you think that tiny clusters of atoms are useful for catalysts?
- How do you think that you could make tiny clusters of metal atoms on a cold surface?

CONTINUED

HINT: Think about breathing onto a cold surface.

- What other uses could be made of tiny groups / clusters of atoms?
- What advantages and disadvantages could there be in using tiny clusters of atoms to help deliver medical drugs and in cancer treatment?
- What else do you think nanomachines could be used for?

1.1 Elements and atoms

Every substance in our world is made up from chemical **elements**. These chemical elements cannot be broken down further into simpler substances by chemical means. A few elements, such as nitrogen and gold, are found on their own in nature, not combined with other elements. Most elements, however, are found in combination with other elements as compounds.

Every element has its own chemical symbol. The symbols are often derived from Latin or Greek words. Some examples are shown in Table 1.1.

Element	Symbol
carbon	C
lithium	Li (from Greek 'lithos')
iron	Fe (from Latin 'ferrum')
potassium	K (from Arabic 'al-qualyah' or from the Latin 'kalium')

Table 1.1: Some examples of chemical symbols.

Chemical elements contain only one type of **atom**. An atom is the smallest part of an element that can take part in a chemical change. Atoms are very small. The diameter of a hydrogen atom is approximately 10^{-10} m, so the mass of an atom is also very small. A single hydrogen atom weighs only 1.67×10^{-27} kg.

1.2 Inside the atom

The structure of an atom

Every atom has nearly all of its mass concentrated in a tiny region in the centre of the atom called the nucleus. The nucleus is made up of particles called nucleons. There are two types of nucleon: **protons** and **neutrons**. Atoms of different elements have different numbers of protons.

Outside the nucleus, particles called **electrons** move around in regions of space called orbitals (see page xx). Chemists often find it convenient to use a simpler model of the atom in which electrons move around the nucleus in electron shells. Each shell is a certain distance from the nucleus at its own particular **energy level** (see page xx). In a neutral atom, the number of electrons is equal to the number of protons. A simple model of a carbon atom is shown in Figure 1.3.

KEY WORDS

element: a substance containing only one type of atom. All the atoms in an element have the same proton number.

atom: the smallest part of an element that can take part in a chemical change. Every atom contains protons in its nucleus and electrons outside the nucleus. Most atoms have neutrons in the nucleus. The exception is the isotope of hydrogen ^1_1H.

protons: positively charged particles in the nucleus of the atom.

neutrons: uncharged particles in the nucleus of the atom.

electrons: negatively charged particles surrounding the nucleus.

energy levels: the specific distances from the nucleus corresponding to the energy of the electrons. Electrons in energy levels further from the nucleus have more energy than those closer to the nucleus. Energy levels are split up into sub-levels which are given the names s, p, d, etc.

IMPORTANT

When we use a simple model of the atom we talk about shells (n = 1, n = 2, etc) and sub-shells 2s, 2p, etc. In this model, the electrons are at a fixed distance from the nucleus. This model is useful when we discuss ionisation energies (Chapter 2).

When we discuss where the electrons really are in space, we use the orbital model. In this model, there is a probability of finding a particular electron within certain area of space outside the nucleus. We use this model for discussing bonding and referring to electrons in the sub-shells.

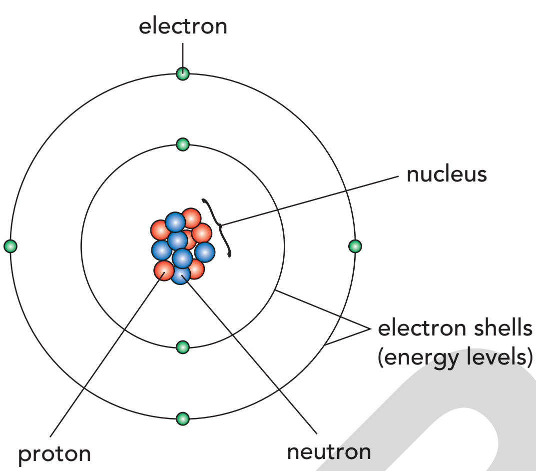

Figure 1.3: A model of a carbon atom. This model is not very accurate but it is useful for understanding what happens to the electrons during chemical reactions.

Atoms are tiny, but the nucleus of an atom is much smaller. If the diameter of an atom were the size of a football stadium, the nucleus would only be the size of a pea. This means that most of the atom is empty space! Electrons are even smaller than protons and neutrons.

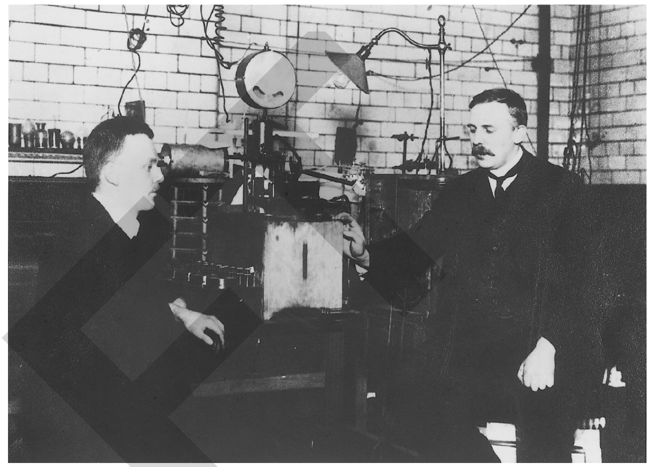

Figure 1.4: Ernest Rutherford (left) and Hans Geiger (right) using their alpha particle apparatus. Interpretation of the results led to Rutherford proposing the nuclear model for atoms.

PRACTICAL ACTIVITY 1.1

Experiments with subatomic particles

We can deduce the electric charge of subatomic particles by showing how beams of electrons, protons and neutrons behave in electric fields. If we fire a beam of electrons past electrically charged plates, the electrons are deflected (change direction) away from the negative plate and towards the positive plate (Figure 1.5a). This shows us that the electrons are negatively charged because opposite charges attract each other and like charges repel each other.

A cathode-ray tube (Figure 1.5b) can be used to produce beams of electrons. At one end of the tube is a metal wire (cathode), which is heated to a high temperature when a low voltage is applied to it. At the other end of the tube is a fluorescent screen, which glows when electrons hit it.

The electrons are given off from the heated wire and are attracted towards two metal plates, which are positively charged. As they pass through the metal plates, the electrons form a beam. When the electron beam hits the screen a spot of light is produced. When an electric field is applied across this beam the electrons are deflected. The fact

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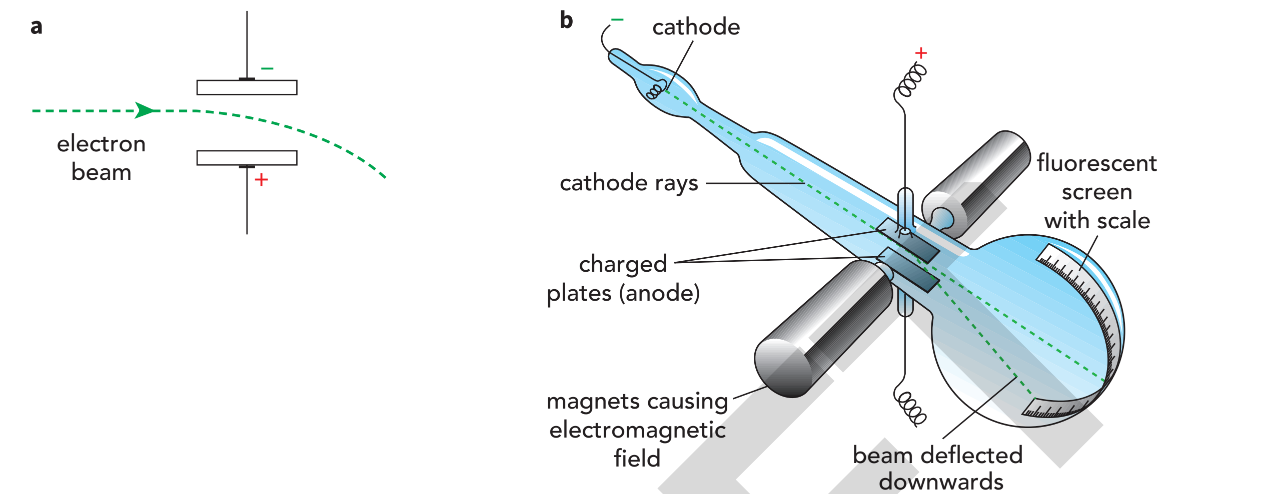


Figure 1.5: a The beam of electrons is deflected away from a negatively charged plate and towards a positively charged plate. **b** The electron beam in a cathode-ray tube is deflected by an electromagnetic field. The direction of the deflection shows us that the electron is negatively charged.

that the electrons are so easily attracted to the positively charged **anode** and that they are easily deflected by an electric field shows us that:

- electrons have a negative charge
- electrons have a very small mass.

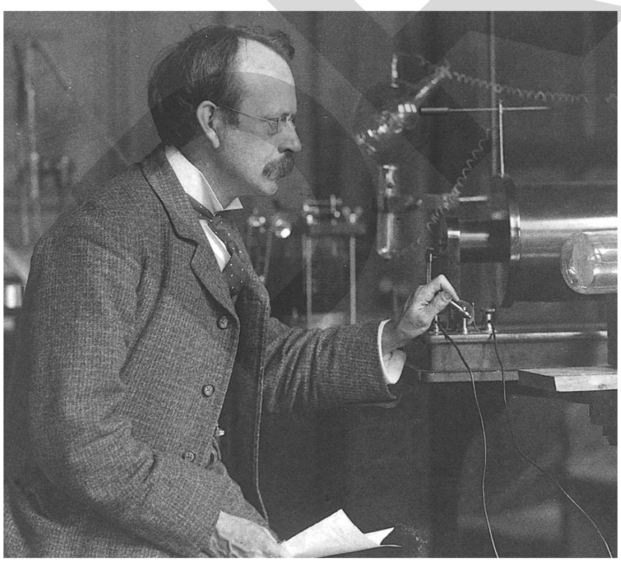


Figure 1.6: J. J. Thomson calculated the charge to mass ratio of electrons. He used results from experiments with electrons in cathode-ray tubes.

KEY WORD

anode: the positive electrode.

In recent years, experiments have been carried out with beams of electrons, protons and neutrons moving at the same velocity in an electric field.

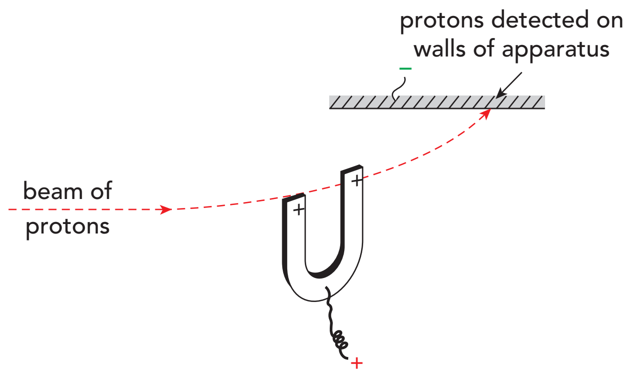


Figure 1.7: A beam of protons is deflected away from a positively charged area. This shows us that protons have a positive charge.

CONTINUED

The results of these experiments show that:

- a proton beam is deflected away from a positively charged plate; as like charges repel, the protons must have a positive charge (Figure 1.7)
- an electron beam is deflected towards a positively charged plate; as opposite charges attract, the electrons must have a negative charge
- a beam of neutrons is not deflected; this shows that they are uncharged.

In these experiments, huge voltages have to be used to deflect the proton beam. This contrasts with the very low voltages needed to deflect an electron beam. These experiments show us that protons are much heavier than electrons. If we used the same voltage to deflect electrons and protons, the beam of electrons would have a far greater deflection than the beam of protons. This is because a proton is about 2000 times heavier than an electron.

IMPORTANT

Remember that like charges repel each other and unlike charges attract each other.

Question

1 A beam of electrons is passing close to a highly negatively charged plate. When the electrons pass close to the plate, they are deflected away from the plate.

  a What deflection would you expect, if any, when the experiment is repeated with beams of **i** protons and **ii** neutrons? Explain your answers.

  b Which subatomic particle (electron, proton or neutron) would deviate the most? Explain your answer.

Subatomic particle	Symbol	Relative mass	Relative charge
electron	e	$\frac{1}{1836}$	−1
neutron	n	1	0
proton	p	1	+1

Table 1.2: Comparing electrons, neutrons and protons.

Masses and charges: a summary

Electrons, protons and neutrons have characteristic charges and masses. The values of these are too small to be very useful when discussing general chemical properties. For example, the charge on a single electron is -1.602×10^{-19} coulombs. We therefore compare their masses and charges by using their relative charges and masses. These are not the actual charges and masses. They are the charges and masses compared with each other in a simple ratio. These are shown in Table 1.2.

1.3 Numbers of nucleons

Atomic (proton) number and mass (nucleon) number

The number of protons in the nucleus of an atom is called the atomic number (proton number) (Z). Every atom of the same element has the same number of protons in its nucleus. It is the atomic number that makes an atom what it is. For example, an atom with an **atomic number** of 11 must be an atom of the element sodium. No other element can have 11 protons in its nucleus. The Periodic Table of elements is arranged in order of the atomic numbers of the individual elements (see Appendix 1, page xxx).

The mass number (nucleon number) (A) is the number of protons plus neutrons in the nucleus of an atom.

KEY DEFINITION

atomic number: the number of protons in the nucleus of an atom. Also called the proton number. Remember that in writing isotopes, this is the figure which is subscript.

How many neutrons?

We can use the mass number and atomic number to find the number of neutrons in an atom. As:

mass number = number of protons + number of neutrons

Then:

$$\text{number of neutrons} = \text{mass number} - \text{atomic number} = A - Z$$

For example, an atom of aluminium has a mass number of 27 and an atomic number of 13. So an aluminium atom has 27 – 13 = 14 neutrons.

Question

2 Use the information in Table 1.3 to deduce the number of electrons and neutrons in a neutral atom of:

a vanadium

b strontium

c phosphorus

Atom	Mass number	Proton number
vanadium	51	23
strontium	84	38
phosphorus	31	15

Table 1.3: Information table for Question 2

Isotopes

All atoms of the same element have the same number of protons. However, they may have different numbers of neutrons. Atoms of the same element that have different numbers of neutrons are called **isotopes**.

KEY DEFINITION

isotope: atoms of the same element with different mass numbers. Note that the word 'atom' is essential in this definition.

Isotopes are atoms of the same element with different **mass numbers**.

KEY WORD

mass number: the number of protons + neutrons in an atom. Also called the nucleon number.

Isotopes of a particular element have the same chemical properties because they have the same number of electrons. They have slightly different physical properties, such as small differences in density or small differences in mass, because they have different numbers of neutrons.

We can write symbols for isotopes. We write the nucleon number at the top left of the chemical symbol and the proton number at the bottom left.

The symbol for the isotope of boron with 5 protons and 11 nucleons is written $^{11}_{5}B$:

nucleon number ⟶ 11
proton number ⟶ 5
$^{11}_{5}\mathrm{B}$

Hydrogen has three isotopes. The atomic structure and isotopic symbols for the three isotopes of hydrogen are shown in Figure 1.8.

When writing generally about isotopes, chemists also name them by leaving out the proton number and placing the mass number after the name. For example, the isotopes of hydrogen can be called hydrogen-1, hydrogen-2 and hydrogen-3.

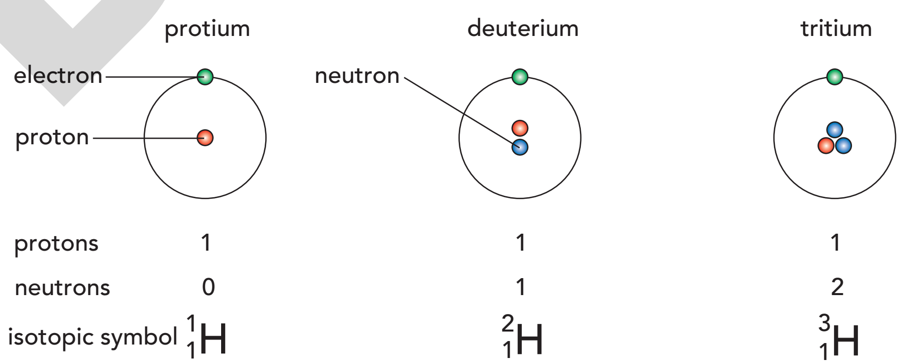


Figure 1.8: The atomic structure and isotopic symbols for the three isotopes of hydrogen.

Remember that in writing isotopes, mass number is the figure which is superscript.

Isotopes can be radioactive or non-radioactive. Specific radioisotopes (radioactive isotopes) can be used to check for leaks in oil or gas pipelines and to check the thickness of paper. They are also used in medicine to treat some types of cancer and to check the activity of the thyroid gland in the throat.

Question

3 Use the Periodic Table on page XXX to help you.
Write isotopic symbols for the following neutral atoms:

a bromine-81

b calcium-44

c iron-58

d palladium-110

How many protons, neutrons and electrons?

In a neutral atom the number of positively charged protons in the nucleus equals the number of negatively charged electrons outside the nucleus. When an atom gains or loses electrons, ions are formed, which are electrically charged. For example:

Cl	+	e^-	→	Cl^-
chlorine atom		1 electron gained		chloride ion
17 protons				17 protons
17 electrons				18 electrons

The chloride ion has a single negative charge because there are 17 protons (+) and 18 electrons (–).

Mg	→	Mg^{2+}	+	$2e^-$
magnesium atom		magnesium ion		2 electrons removed
12 protons		12 protons		
12 electrons		10 electrons		

The magnesium ion has a charge of 2+ because it has 12 protons (+) but only 10 electrons (–).

The isotopic symbol for an ion derived from sulfur-33 is $^{33}_{16}S^{2-}$. This sulfide ion has 16 protons, 17 neutrons (because 33 – 16 = 17) and 18 electrons (because 16 + 2 = 18).

IMPORTANT

Ions: charged particles formed by the loss or gain of electrons from an atom or group of covalently bonded atoms. Remember that positive ions are formed when one or more electrons are *lost* by an atom and that negative ions are formed when one or more electrons are *gained* by an atom.

WORKED EXAMPLE

1 Deduce the number of electrons in the ion $^{52}_{24}Cr^{2+}$.

Solution

Step 1: Work out the number of protons. This is the subscripted number 24.

Step 2: Number of protons = number of electrons in the neutral atom. So number of electrons in the atom is 24.

Step 3: For a positive ion subtract the number of charges (because electrons have been lost from the atom). For a negative ion add the number of charges (because electrons have been gained).

So for Cr^{2+}, 24 – 2 = 22 electrons.

Questions

4 Deduce the number of electrons in each of these ions:

a $^{40}_{19}K^+$

b $^{15}_{7}N^{3-}$

c $^{18}_{8}O^{2-}$

d $^{71}_{31}Ga^{3+}$

5 In which one of the following ways are isotopes of the same element exactly the same?

A The sum of the number of electrons and the number of neutrons in each atom.

B The mass of the nucleus in each atom.

C The number of electrons in each atom.

D The sum of the number of protons and the number of neutrons in each atom.

6 Deduce the number of electrons, protons and neutrons in each of these ions:

a $^{81}_{35}Br^{-}$

b $^{136}_{38}Ce^{3+}$

REFLECTION

Read the second paragraph on page XX again about Rutherford's work in discovering the nucleus. Discuss these questions with another learner:

1. Why, in Rutherford's experiments, did most of the alpha particles go straight through the metal foil and so few bounced back?
2. Suggest what happened to the alpha particles that went a little way from the nucleus. Use ideas of attractive or repulsive forces.
3. Use your knowledge of what you have learned in this chapter to think about any other experiments that could have been used.

How much did you contribute to the discussion? Could you have contributed more?

SUMMARY

Beams of protons and electrons are deflected by electric fields but neutrons are not.
The atom consists of positively charged protons and neutral neutrons in the nucleus, surrounded by negatively charged electrons arranged in energy levels (shells).
Isotopes are atoms with the same atomic number but different mass numbers. They only differ in the number of neutrons they contain.

EXAM-STYLE QUESTIONS

1 Boron is an element in Group 13 of the Periodic Table.

Boron has two isotopes.

a **Deduce** the number of **i** protons, **ii** neutrons and **iii** electrons in one neutral atom of the isotope $^{11}_{5}B$. **[3]**

b What do you understand by the term *isotope*? **[1]**

c **State** the relative masses and charges of:

i an electron **[2]**

ii a neutron **[2]**

iii a proton **[2]**

[Total: 10]

COMMAND WORDS

Deduce: conclude from available information.

State: express in clear terms.

CONTINUED

2 Zirconium, Zr, and hafnium, Hf, are metals.

An isotope of zirconium has 40 protons and 91 nucleons.

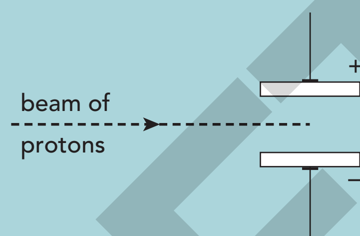
a **i** Write the isotopic symbol for this isotope of zirconium. **[1]**

ii State the number of neutrons present in one atom of this isotope. **[1]**

b The symbol for a particular ion of hafnium ion is $^{180}_{72}Hf^{2+}$.

Deduce the number of electrons that are present in one of these hafnium ions. **[1]**

c The subatomic particles present in zirconium and hafnium are electrons, neutrons and protons. A beam of protons is fired into an electric field produced by two charged plates, as shown in the diagram.



i Describe how the beam of protons behaves when it passes through the gap between the charged plates. **[2]**

ii Explain your answer. **[2]**

d Describe and explain what happens when a beam of neutrons passes through the gap between the charged plates. **[2]**

[Total: 9]

3 **a** Describe the structure of an atom, giving details of the subatomic particles present. **[6]**

b Explain the terms *atomic number* and *nucleon number*. **[2]**

c Copy and complete the table: **[2]**

Neutral atom	Atomic number	Nucleon number	Numbers of each subatomic particle present
Mg	12	24	
Al	13	27	

d Explain why atoms are neutral. **[1]**

e An oxygen atom has 8 protons in its nucleus. Explain why it cannot have 9 protons. **[1]**

f When deducing the relative mass of an atom, the electrons are not used in the calculation. Explain why not. **[1]**

[Total: 13]

4 The symbols below describe two isotopes of the element uranium.

$$^{235}_{92}U \quad ^{238}_{92}U$$

a State the meaning of the term *isotope*. **[1]**

b State two ways in which these two isotopes of uranium are identical. **[2]**

c State how these isotopes differ. **[2]**

d State the number of electrons present in one U^{2+} ion. **[1]**

[Total: 6]

COMMAND WORDS

Describe: state the points of a topic / give characteristics and main features.

Explain: set out the reasons why something happens or make the relationships between things clear. Make sure that you write down the evidence logically.

CONTINUED

5 The table below shows the two naturally occurring isotopes of chlorine.

a Copy and complete the table.

	${}^{35}_{17}Cr$	${}^{37}_{17}Cr$
number of protons		
number of electrons		
number of neutrons		

[3]

b The relative atomic mass of chlorine is 35.5. What does this tell you about the relative abundance of the two naturally occurring isotopes of chlorine? **[2]**

c Magnesium chloride contains magnesium ions, Mg^{2+}, and chloride ions, Cl^{-}.

i Explain why a magnesium ion is positively charged. **[1]**

ii Explain why a chloride ion has a single negative charge. **[2]**

[Total: 8]

SELF-EVALUATION CHECKLIST

After studying this chapter, complete a table like this:

I can	See section...	Needs more work	Almost there	Ready to move on
understand that every atom has an internal structure with a nucleus in the centre and the negatively charged electrons arranged in ‘shells’ outside the nucleus	1.2			
understand that most of the mass of the atom is in the nucleus, which contains protons (positively charged) and neutrons (uncharged)	1.2			
understand that beams of protons and electrons are deflected by electric fields but neutrons are not	1.2			
understand that atoms of the same element have the same number of protons; this is called the atomic (proton) number	1.3			
understand that the mass (nucleon) number (A), is the total number of protons and neutrons in an atom	1.3			

CONTINUED

I can	See section...	Needs more work	Almost there	Ready to move on
deduce the number of neutrons in an atom by subtracting the atomic number from the mass number ($A-Z$)	1.3			
understand that in a neutral atom, the number of electrons equals the number of protons: when there are more protons than electrons, the atom becomes a positive ion; when there are more electrons than protons, a negatively charged ion is formed	1.3			
understand that isotopes are atoms with the same atomic number but different mass numbers; they only differ in the number of neutrons they contain.	1.3			

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Chemistry

for Cambridge International AS & A Level

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Contents

1 Atomic structure

Syllabus overview

This chapter covers all the topics covered in Chapter 1 of the Coursebook and Workbook.

- This section of the syllabus provides learners with the following knowledge about the structure of the atom and the properties and arrangement of the subatomic particles within the atom.
- It also deals with how the numbers of these subatomic particles change for ions and isotopes.
- The number of electrons within the atom will link to Chapter 2 (Electrons in atoms).
- This section is theoretical and there are no realistic opportunities to assess practical techniques.
- The assessment objectives AO1 and AO2 can be covered in this topic and some of the mathematical skills listed in Section 6 of the syllabus can be used.

Topic teaching plan

Syllabus topic	Number of lessons	Outline of lesson content	Resources
1.1.1–1.1.5	1	Run through the history of how the structure of the atom was worked out and use this to illustrate how scientific theories develop.	**Coursebook:** Science in context 'Developing an idea: Nanomachines' 1.2 'Inside the atom' Exam-style question 3 **Workbook:** Exercise 1.4
		Describe the relative masses and electrical charges of protons, neutrons and electrons.	**Coursebook:** 1.2 'Masses and charges: a summary' Exam-style question 3 **Workbook:** Exercises 1.1 and 1.2
		The behaviour of beams of subatomic particles in an electric field.	**Coursebook:** 'Practical activity 1.1: Experiments with subatomic particles', Question 1 **Workbook:** Exam-style question 2d
		Investigate how the atomic number and mass (nucleon) number can be used to calculate the numbers of the three subatomic particles: proton, electron and neutron.	**Coursebook:** 1.3 'Numbers of nucleons', Question 2 **Workbook:** Exercise 1.4 Exam-style question 1

(Continued)

Syllabus topic	Number of lessons	Outline of lesson content	Resources
1.1.6 and 1.2.1–1.2.4	1	Explore the differences between atoms of isotopes of the same element.	**Coursebook:** 1.3 'Isotopes', Question 3 **Workbook:** Exercise 1.3 and Exam-style questions 1 and 2
		Explain how ions are formed and use the number of charges, atomic number and mass (nucleon) number to calculate the numbers of subatomic particles in an ion.	**Coursebook:** 1.3 'How many protons, neutrons and electrons?', Question 4 Exam-style questions 1 and 2 **Workbook:** Exercise 1.3 Exam-style question 2
		Explain how isotopes differ in terms of their physical and chemical properties.	**Coursebook:** 1.3 'Isotopes'

Topics 1.1.1 to 1.1.5

Learners will:

- run through the history of how the structure of the atom was worked out and use this to illustrate how scientific theories develop
- describe the relative masses and electrical charges of protons, neutrons and electrons
- learn about the behaviour of beams of subatomic particles in an electric field
- investigate how the atomic number and mass (nucleon) number can be used to calculate the numbers of the three subatomic particles: proton, electron and neutron.

Suggested teaching time:

1 hour

Links to other components in this series

Component	Resource	Description
Coursebook	Science in Context 'Developing an idea: Nanomachines' 1.2 'Inside the atom' Exam-style question 3 1.2 'Masses and charges: a summary', Question 3 Practical activity 1.1: 'Experiments with subatomic particles', Question 1 1.3 'Numbers of nucleons', Question 2	• Discuss how the model of the atom evolved and construct a timeline on how the concept of the atom evolved • Draw and label a model of the atom • List the relative charges and masses of subatomic particles • Describe the relationship between the atomic and mass (nucleon) numbers and the numbers of electrons, protons and neutrons • Investigate the experiments on particles in an electric field.

Component	Resource	Description
Workbook	Exercise 1.1 Exercise 1.2 Exercise 1.3 Exercise 1.4 Exam-style question 1 Exam-style question 2d	• Summarise our knowledge of atomic structure • Review Rutherford's experiment.

Common misconceptions

- Because they sound similar, the words nucleon and neutron sometimes get confused.
- When electrons and protons are deviated by an electric field, the deviations are *not* equal. Some learners will not consider the effect of mass on the deviation.

Lesson starters

This lesson could well be the first lesson of the course and the learners will almost certainly have covered most of the subject matter at IGCSE/O Level. Therefore, depending on the learners, some of the knowledge could be assumed and the major part of the subject matter can be treated as a recall exercise.

Two suggestions are given here. The choice of which activity is used will depend on what resources are available, the time available and how the learners are progressing with this topic.

1 Idea A (15–20 minutes)

Use some of the facts about the atom's tiny size. There are plenty of short comparisons you could use to illustrate the small sizes of atoms and molecules. For example, one drop of water is about 0.1 cm^3 or 0.1 g. When you work out how many water molecules there are in that small amount of water it comes to a large number but if you calculate how long it would take to count out the water molecules at 1 a second it comes to about 106 million million years. The question then arises, 'How do we know these particles/atoms exist and how did we discover what makes them up?'

The story of the atom's structure is not just history; it also shows how scientific theories evolve according to the evidence available.

In this activity you can give learners the names of the scientists involved and the theories with which they are associated. This information can be written on a sheet but jumbled up so learners can reconstruct the information.

To reinforce the learners' understanding of what happened in Rutherford's experiment, they can go on the internet to view video clips and animations that explain what he and his co-workers did and found out: search online for 'Discovery of the atomic nucleus' or 'Animations of Rutherford's experiment'.

› **Assessment ideas:** Ask learners, working in small groups, to construct a timeline showing the order in which the theories were put forward and by whom. They can also write a short criticism of what is wrong about each theory. Workbook Exercise 1.4 on the discovery of the nucleus reinforces how the model of the atom evolved.

2 Idea B (15 minutes)

Use 'Before you Start' in the Coursebook to set a quick recall test (see Teaching techniques 'Diagnostic exercises'). This can show you how much information learners can recall.

› **Assessment ideas:** Learners can self-assess and record what they find difficult on paper. You can collect these in or ask learners to record their results in their books for future reference. To make sure the exercise does not take too long, only allow 1–2 minutes per set of questions and 1 minute each for the answers.

Main activities

Below are several teaching activities which you can pick and choose from in order to tailor the lesson to your class's needs.

1 Numbers of subatomic particles (20–25 minutes)

The information required is in Coursebook Section 1.2 'Masses and charges: a summary' and Section 1.3 'Numbers of nucleons'. Learners can use the relationships between atomic and mass (nucleon) number and the numbers of electrons, protons and neutrons in a *neutral* atom. It is worthwhile at this stage to discuss why the atomic number is defined as the number of protons and *not* the number of electrons or neutrons (see Coursebook Section 1.3 'Numbers of nucleons').

Element	Symbol	Atomic number	Mass (nucleon) number	Number of protons	Number of electrons	Number of neutrons

Table 1.1

Give at least two relevant numbers for each of, for example, 10 elements and ask your learners to complete Table 1.1. You could also test their knowledge of symbols by leaving one of the first two columns empty for some elements.

A more difficult approach would be to leave both the first two columns empty and ask them to use the Periodic Table in the Coursebook to identify each element.

› **Assessment ideas: Hinge-point Question:** Which of the sets of numbers in Table 1.2 *cannot* be used to calculate the numbers of subatomic particles in a neutral atom? The correct answer is B. See below for explanations.

	First number	Second number
A	atomic number	mass (nucleon) number
B	atomic number	number of protons
C	mass (nucleon) number	number of neutrons
D	number of electrons	mass (nucleon) number

Table 1.2

This is a quick and good diagnostic tool to assess how well learners understand the relationships between the numbers.

Explanations:

A Incorrect. If they respond A they do not realise that the atomic number gives the number of protons and electrons, and that the number of neutrons can be found by subtracting the atomic number (Z) from the mass number (A).

B Correct: the atomic number will only give the number of protons and electrons, which is the same as the second number. They cannot calculate the number of neutrons from these numbers.

C Incorrect. They haven't rearranged the equation: number of neutrons = $A - Z$. They are given A and they are given the number of neutrons. Therefore they can find the number of protons and electrons from A – number of neutrons = number of protons (number of electrons).

D Incorrect. They haven't realised that the number of electrons in a neutral atom is the same as the atomic number (the number of protons) and if this is subtracted from the mass number, they will obtain the number of neutrons.

2 A concept map for atomic structure (20 minutes)

The learners can construct a *concept map* (see Teaching techniques 'Concept mapping'), using what they know about atomic structure and the Coursebook. The words they can use are:

atom, electron, proton, neutron, nucleus, nucleon

plus any more words that you feel are relevant (e.g. atomic number, mass (nucleon) number, negative plate, negative charge).

This can be done in more than one way and so this allows your learners some choice. Tell them that there are many variations of a creative, correct answer.

Assessment ideas: For assessment purposes, the concept maps can either be handed in for marking, photographed using their phones or peer assessed (between groups).

The quality of a concept map depends very much on the phrases/words placed on the arrows, but it is quite difficult for learners to get it wrong.

Differentiation

Stretch and challenge

- Ask learners to write one paragraph summarising the deflection of subatomic particles in electrical fields. Their descriptions should contain explanations of the results obtained (see Coursebook Practical activity 1.1: Experiments with subatomic particles).
- Give more words to incorporate into the concept map.

Support

- When completing the numbers of subatomic particles in Table 1.1, give learners support by completing the first two or three sections with the learners. Ask learners, 'How did you calculate the number of neutrons?' to encourage them to reflect on how they approached the process.
- Suggest some words to use to help learners devise their concept maps, but delay telling them where they should go.

WRAP UP AND REFLECTION IDEAS

If there is sufficient time, split up the learners into groups to do some winding-up exercises:

- Play 'Guess the word' (see Teaching techniques 'Group work'). Choose words related to the chapter.
- Look at the completed concept maps and ask each group to do a one-minute presentation describing what they have done and why they have used the links they have. What did they find easy/hard about this exercise? Many learners find this sort of exercise difficult at first, but once they get some practice, there is a strong sense of ownership associated with this type of learning.

CROSS-CURRICULAR LINKS

Literacy

- The concept map is mainly a literacy exercise and will test their ability to form clear sentences/ clauses from simple words and scientific vocabulary.

Numeracy

- Simple mathematical skills such as addition, subtraction and use of a formula ($A - Z$ = number of neutrons).

Topics 1.1.6 and 1.2.1 to 1.2.4

Learners will:

- explore the differences between atoms of isotopes of the same element
- explain how ions are formed and use the number of charges, atomic number and mass (nucleon) number to calculate the numbers of subatomic particles in an ion
- explain how isotopes differ in terms of their physical and chemical properties.

Suggested teaching time:

1 hour

Links to other components in this series

Component	Resource	Description
Coursebook	1.3 'Isotopes', Question 3 1.3 'How many protons, neutrons and electrons?', Question 4 1.3 'Isotopes' Exam-style questions 1 and 2	• Learn the definition of an isotope and why isotopes have the same chemical properties but slightly different physical properties • Investigate how ions are formed and use this knowledge to calculate the numbers of electrons in positive and negative ions • Either interpret the symbols ${}^{A}_{Z}X$ or write the symbols using the numbers of subatomic particles present.
Workbook	Exercise 1.3 Exam-style questions 1 and 2	• Assess understanding of isotopes and the nomenclature used to describe them.

Common misconceptions

Many learners believe that *all* isotopes are radioactive.

Lesson starters

Two suggestions are given here. The choice of which activity is used will depend on what resources are available, the time available and how the learners are progressing with this topic.

1 Idea A (10 minutes)

Review the previous lesson.

> **Assessment ideas:** Give learners a quick-fire test (see Teaching techniques 'Diagnostic exercises') on the concepts learned from the previous lesson.

2 Idea B (10 minutes)

The word isotope means same (iso = same) place (tope = place). The *place* refers to the place in the Periodic Table. Ask learners why these (neutral) atoms are placed in the same place in the Periodic Table? What must they have that is identical and what could be different? Hopefully, this discussion will lead to understanding the definition of an isotope and that the proton or atomic number is unique to each element.

Main activities

Below are several teaching activities which you can pick and choose from in order to tailor the lesson to your class's needs.

1 Isotopes (20 minutes)

Use the Coursebook Section 1.3 'How many protons, neutrons and electrons?'

If it hasn't been introduced already, define the term *isotope* and the symbol ${}^{x}_{y}A$. Then explain to the learners that they will be expected to use this notation to find the number of protons (y), electrons (= y) and neutrons ($x - y$), or use the numbers of subatomic particles to find the values of x and y. Encourage learners to use the internet to reinforce their understanding: search online for 'Videos on isotopes'.

Assessment ideas: Some simple exercises on isotopes: using Questions 2 and 3 in the Coursebook or questions such as:

What are the similarities and differences for these two isotopes?

A ${}^{35}_{17}Cl$

B ${}^{37}_{17}Cl$

Note: they need to give the numbers of subatomic particles present.

You can use hinge-point Questions 4 and 5 (in Section 1.3 'How many protons, electrons and neutrons?').

2 Ions, isotopes and numbers of electrons, protons and neutrons (15–20 minutes)

Ask learners to complete a few equations which illustrate the formation of ions. Depending on the ability of the learners, these can be a mixture of positive and negative ions or separated into two categories.

To build confidence the ions formed can be familiar (e.g. $Na \rightarrow Na^+$ and $Cl \rightarrow Cl^-$) and gradually increased in difficulty.

At the end of the exercise they should be able to write definitive statements about the formation of ions and whether the number of electrons remaining increases or decreases. They should also be able to give the change in the electron number from the charge on the ion.

Assessment ideas: Ask learners to fill in a table similar to Table 1.1 above. You can use the headings in Table 1.3.

Particle	e^- lost or gained in forming ion	Number of electrons	Number of protons	Number of neutrons
${}^{37}_{17}Cl^-$	1 gained	18	17	20

Table 1.3

Alternatively, use Coursebook Section 1.3, Question 4.

At this stage, the whole class or separate groups could discuss why the number of protons is given as the defining number for an element. You could then ask learners to write a paragraph with the appropriate explanation.

Assessment ideas: Isotopes have the same number of protons and electrons. Give learners two properties, such as chemical reactivity and density. Ask them to explain, giving reasons, which of these two properties is different for isotopes and which is unchanged.

Differentiation

Stretch and challenge

Ask learners to explain why the formation of an ion is through either oxidation or reduction.

Support

The equations for the formation of ions are partially completed at first and then made slightly harder.

e.g. …. – $2e^- \rightarrow Fe^{2+}$;   Cl …. $e^- \rightarrow Cl^-$

WRAP UP AND REFLECTION IDEAS

Hinge-point question

Which of the statements A to D in Table 1.4 is a correct statement about the two particles Q and R below?

$^{52}_{24}Cr^{3+}$ $^{54}_{24}Cr^{2+}$

Q   R

A	Q has one more electron than R	Incorrect. Q has more positive charges, therefore it has one less electron than R.
B	Q and R both have 28 neutrons	Incorrect. R has 30 neutrons (54 – 24). The learner has calculated the number for Q and assumed they were the same, showing they have misunderstood what isotopes are.
C	Q has 24 protons	Correct
D	R has 24 electrons and 24 protons	Incorrect. Q and R both have 24 protons but Q has $21e^-$ and R has $22e^-$. They have not accounted for the charge on the ions.

Table 1.4

- If they were incorrect, what did they misunderstand?
- Can they explain what is wrong about the incorrect alternatives?
- To assess what concepts are *understood/partially understood/not understood*, the learners can complete a quick 'traffic lights' exercise (see Teaching techniques 'Diagnostic exercises'). A sample statement/question is: 'The number of neutrons in the atom of an element can vary.'

CROSS-CURRICULAR LINKS

Numeracy

- Learners will need to remember the formula for calculating the number of neutrons from the atomic number and the mass number.
- They will need to translate the numbers in the symbols $^{A}_{Z}X$ for both neutral atoms and ions.

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Chemistry

for Cambridge International AS & A Level

WORKBOOK

Roger Norris & Mike Wooster

Third edition

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Contents

Chapter 1

Atomic structure

CHAPTER OUTLINE

In this chapter you will learn how to:

- describe the structure of the atom and the relative charges and masses of protons, neutrons and electrons
- describe how protons, neutrons and electrons behave in electric fields
- deduce the number of protons, neutrons and electrons in atoms and ions
- define proton (atomic) number, mass (nucleon) number and isotopes
- explain why isotopes have the same chemical properties but some of their physical properties are different
- use the symbolism $^{x}_{y}A$ for isotopes.

Exercise 1.1 Atomic structure

This exercise will familiarise you with the properties of the three types of subatomic particle.

TIP

Remember to read the stem of the question carefully. Here it states that some words can be used more than once.

Copy and complete these sentences using words from this list. Some words may be used more than once.

electron **negative** **neutrons** **positively** **protons** **relative** **shells**

An atom contains a dense **nucleus** surrounded by ____________ of **electrons**. The nucleus contains the nucleons (____________ and ____________). **Protons** are charged, electrons have a ____________ charge and ____________ are uncharged. The ____________ and **neutrons** have the same ____________ mass. The mass of an ____________ is negligible (hardly anything).

KEY WORDS

nucleus: the dense core at the centre of an atom containing neutrons (except the ^{1}H isotope) and protons.

electron: negatively charged particle found in orbitals outside the nucleus of an atom. It has negligible mass compared with a proton.

proton: positively charged particle in the nucleus of an atom.

neutron: uncharged particle in the nucleus of an atom, with the same relative mass as a proton.

Exercise 1.2 Terms used in atomic structure

This exercise will familiarise you with some terms related to atomic structure.

Match the boxes 1 to 4 on the left with the descriptions A to D on the right.

1	**Atomic number**	A	The tiny central core of the atom
2	**Mass number**	B	The number of protons plus neutrons in the nucleus
3	**Neutrons**	C	The number of protons in the nucleus of an atom
4	**Nucleus**	D	Uncharged particles in the nucleus

TIP

It is important that you learn the exact meanings of scientific words such as *mass number* and key definitions such as *isotopes*.

When defining terms, you must be precise.

Exercise 1.3 Isotopes

This exercise will familiarise you with the concept of **isotopes** and help you deduce the number of particular subatomic particles in an atom.

TIP

Number of neutrons = mass number – atomic number

KEY WORD

isotope: atoms of an element with the same number of protons but a different number of neutrons.

a Deduce the number of protons and electrons or neutrons represented by the letters A to **F**.

Isotope	Number of protons	Number of electrons	Number of neutrons
$^{86}_{36}Kr$	36	**A**	50
$^{115}_{49}In$	**B**	49	**C**
$^{50}_{24}Cr$	**D**	**E**	**F**

Table 1.1: Isotopes.

TIP

The top number in an isotopic formula is the number of protons + neutrons and the bottom number is the proton number.

b Here is a 'cell' of the Periodic Table:

38
Sr
87.6

i Explain why the relative atomic mass is not a whole number.

ii An isotope of strontium has a nucleon number of 90. How many neutrons are there in this isotope?

iii Explain in terms of the charge on the subatomic particles why the strontium ion has a 2+ charge.

c How many protons, neutrons and electrons do the following species have?

i $^{27}_{13}Al$

ii $^{133}_{55}Cs^{+}$

iii $^{17}_{8}O^{2-}$

Exercise 1.4 The discovery of the nucleus

This exercise explores how the nucleus was discovered and will familiarise you with the behaviour of charged particles.

TIP

When answering questions about unfamiliar material, always:

- Read the information carefully, noting down the key points.
- Take note of the command words such as explain and suggest. The definitions are given in the glossary if you're not sure what these mean.

In 1910, researchers in Manchester, UK, fired alpha-particles (α-particles) at thin sheets of gold foil. Some of the α-particles passed straight through the foil (course A in Figure 1.1). Others were deflected slightly (course B). About 1 in every 20 000 was deflected backwards (course C).

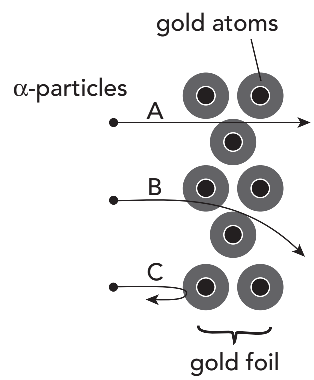

Figure 1.1: Alpha-particles are fired at gold foil.

a Alpha-particles are helium nuclei. Helium atoms have 2 protons and 2 neutrons. Write the isotopic symbol for a helium nucleus.

b Suggest, in terms of the structure of the atoms, why most α-particles passed straight through the foil.

c Explain why some α-particles were deflected slightly.

d Suggest, in terms of the structure of the atoms, why so few α-particles were deflected backwards.

e Suggest what would happen in this experiment if a beam of neutrons were fired at the gold foil. Explain your answer.

f Explain why two different isotopes of helium have different densities.

EXAM-STYLE QUESTIONS

1 This question is about isotopes and subatomic particles. The diagram in Figure 1.2 shows the structure of an isotope of lithium.

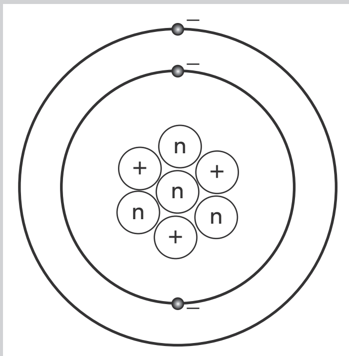

Figure 1.2

a **Describe** the number, charge and relative mass of each subatomic particle present. **[5]**

b Explain why two different isotopes of lithium have the same chemical properties. **[1]**

c Write the isotopic symbol for the lithium atom shown. **[2]**

d Explain why a lithium ion is positively charged. **[1]**

[Total: 9]

2 Cobalt and nickel are next to each other in the Periodic Table.

27 Co 58.9	28 Ni 58.7

a Which one of these elements has the higher atomic number? Explain your answer. **[1]**

b **Suggest** why nickel has a lower relative atomic mass than cobalt. **[3]**

c The isotopic symbols of two isotopes are:

$^{59}_{27}\mathrm{Co}$ $\quad$ $^{58}_{28}\mathrm{Ni}$

i Which one of these isotopes has a greater number of neutrons? **Explain** your answer. **[1]**

ii Which one of these isotopes has fewer electrons? Explain your answer. **[1]**

iii An ion of cobalt has 27 protons and 24 electrons. Give the symbol for this ion. **[1]**

TIPS

In part **a**, don't forget that there are three types of particle as well as three things to describe. For simple questions you may have to write two points to get one mark.

In part **c**, don't forget the charge on the lithium ion.

COMMAND WORD

Describe: state the points of a topic / give characteristics and main features.

TIP

Note, the number of marks available. In parts **2 b** and **2 d** you need to give at least three separate points in order to gain full marks.

COMMAND WORD

Suggest: apply knowledge and understanding to situations where there are a range of valid responses in order to make proposals / put forward considerations.

CONTINUED

d A beam of electrons is fired through an electric field between two charged plates.

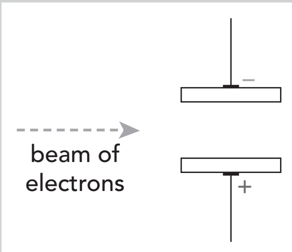

Figure 1.3

Describe how the electron beam behaves when it passes through the plates. Explain your answer. **[3]**

[Total: 10]

COMMAND WORD

Explain: set out purposes or reasons / make the relationships between things evident / provide why and/or how and support with relevant evidence.

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Chemistry

for Cambridge International AS & A Level

PRACTICAL WORKBOOK

Roger Norris & Mike Wooster

Third edition

Contents

Chapter 1

Masses, moles and atoms

CHAPTER OUTLINE

This relates to Chapter 1: Atomic structure; Chapter 2: Electrons in atoms; Chapter 3: Atoms, molecules and stoichiometry in the coursebook.

In this chapter you will complete investigations on:

- 1.1 Empirical formula of hydrated copper(II) sulfate crystals
- 1.2 Relative atomic mass of magnesium using molar volumes
- 1.3 Percentage composition of a mixture of sodium hydrogencarbonate and sodium chloride
- 1.4 Relative atomic mass of calcium by two different methods – molar volume and titration

Practical investigation 1.1: Empirical formula of hydrated copper(II) sulfate crystals

In this investigation you will determine the empirical formula (refer to Chapter 3 of the coursebook if required) of **hydrated** copper(II) sulfate by finding the value of x in $CuSO_4.xH_2O$. You will weigh out some hydrated copper(II) sulfate in an evaporating basin, heat it to constant mass, determine the mass of water present in your sample and then find the molar ratio $CuSO_4 : H_2O$.

KEY WORD

hydrated (crystals): Crystals that have a specific number of moles of water associated with their crystal structure.

YOU WILL NEED

Equipment:

• pipe-clay triangle • evaporating basin • Bunsen burner and tripod • tongs
• glass stirring rod • two heat-resistant pads • spatula

Access to:

• supply of gas • top-pan balance that reads to at least two decimal places

Safety considerations

- Make sure you have read the advice in the Safety section at the beginning of this book and listen to any advice from your teacher before carrying out this investigation.
- You must wear eye protection at all times and tie your hair back if it is long.
- Copper(II) sulfate is an irritant and is harmful if swallowed.
- Carry the evaporating basin to the top-pan balance on a heat-resistant pad. Do not use the tongs to carry it.

Method

1 Weigh an empty evaporating basin, and then weigh the mass of crystals that you have been given. Record your measurements here:

Mass of basin + $CuSO_4{\cdot}xH_2O$ crystals .... g

Mass of basin .... g

Mass of $CuSO_4{\cdot}xH_2O$ crystals .... g

2 Put the pipe-clay triangle and the evaporating basin containing your crystals on the tripod as shown in Figure 1.1.

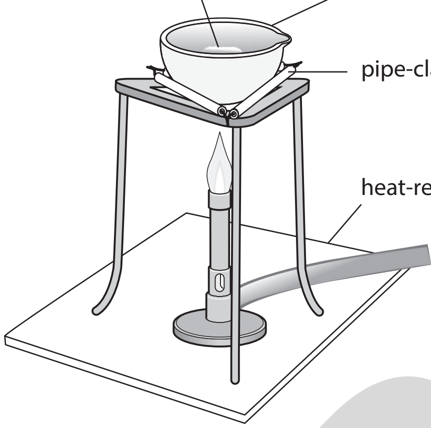

Figure 1.1: Heating a solid

3 Copper(II) sulfate **decomposes** if heated too strongly. Heat the crystals *very gently*. A low, just-blue Bunsen flame should be used for this.

4 While you are heating the crystals, stir them using the glass stirring rod. At the same time grip the evaporating basin using the tongs to prevent it toppling over and spilling the contents.

5 At first, the copper(II) sulfate will be 'sticky' but after a short time it should not cling to the glass rod and will become powdery.

6 The colour of the copper(II) sulfate will change from blue to light blue, and then to a very light grey – almost white.

7 When it gets to this stage, weigh the evaporating basin and anhydrous copper(II) sulfate.

Mass of basin + copper(II) sulfate = ... g

8 Reheat the powder for a short while and then reweigh it. If constant mass is obtained, then all the water of crystallisation will have been driven out of the crystals.

Mass of basin + copper(II) sulfate = ... g

KEY WORD

decomposition: The breakdown of a substance into two or more other substances.

9 If the mass has decreased, then keep on reheating and reweighing until a constant mass is obtained.

Repeat (1) mass of basin + copper(II) sulfate = …………….. g

Repeat (2) mass of basin + copper(II) sulfate = …………….. g

Repeat (3) mass of basin + copper(II) sulfate = …………….. g

TIP

Note that anhydrous $CuSO_4$ absorbs water from the air when it is cool.

Analysis, conclusion and evaluation

a Calculate the mass of the anhydrous copper(II) sulfate remaining, and then the mass of water that has been lost from the crystals on heating. This is the water of crystallisation.

Mass of anhydrous $CuSO_4$ = …………….. g

Mass of water of crystallisation = …………….. g

b Using the grid supplied, draw a set of axes with the mass of anhydrous copper(II) sulfate on the horizontal (*x*) axis and the mass of the water of crystallisation on the vertical (*y*) axis. Use suitable scales and label the axes.

- Plot the points on the graph.
- Reject any **anomalous points** (that are obviously wrong).
- Draw a best-fit line through the remaining points.

c Use your line to calculate the mass of water that combines with 1.60 g of anhydrous copper(II) sulfate ($CuSO_4$).

Mass of water = ………………… g

d From your result, calculate the number of moles of water that combine with 1 mole of anhydrous $CuSO_4$.

……………………………………………………………………………………………

……………………………………………………………………………………………

……………………………………………………………………………………………

e Calculate the value of x in the formula $CuSO_4{\cdot}xH_2O$.

x = ………………..………………..

f Which point on your graph should you be most confident about?

Explain your answer.

……………………………………………………………………………………………

……………………………………………………………………………………………

……………………………………………………………………………………………

g Explain any points:

i that lie *above* your best-fit line

……………………………………………………………………………………………

……………………………………………………………………………………………

……………………………………………………………………………………………

KEY WORD

anomalous points: Points that do not fit in with the pattern of the rest of the results.

TIP

Remember that your line must go through the origin (0, 0).

TIP

If the point lies above the line, the ratio of water to anhydrous copper(II) sulfate is higher than it should be.

ii that lie *below* your best-fit line.

....

....

....

TIP

If the point lies below the line, the ratio of water to anhydrous copper(II) sulfate is lower than it should be.

h Copper(II) sulfate crystals lose their water of crystallisation between 100 °C and 350 °C. They start to decompose at approximately 600 °C.

Briefly describe a better way of heating the copper(II) sulfate crystals in this experiment and explain why it is an improvement on the method you used.

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Practical investigation 1.2: Relative atomic mass of magnesium using molar volumes

The objective of this investigation is to find the relative atomic mass of magnesium using its reaction with dilute hydrochloric acid to give hydrogen gas.

Refer to Chapter 3: Atoms, molecules and stoichiometry in the coursebook for more details of the theory. The equation for the reaction between magnesium and hydrochloric acid is:

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

1 mol of any gas occupies 24 000 cm^3 (at room temperature and pressure).

This reaction can be used to find the relative atomic mass of magnesium. By determining the number of moles of hydrogen produced by a known mass of magnesium (m), the number of moles (n) of magnesium can be determined.

The relative atomic mass of magnesium can be found using $A_r = \frac{m}{n}$

Because the masses of short lengths of magnesium ribbon are very small and difficult to measure on a top-pan balance, you will measure out a 10 cm length and weigh it. You will then estimate the masses of different shorter lengths and use these for your experiments.

YOU WILL NEED

Equipment:

• apparatus for the collection and measurement of a gas • small piece of steel wool • one 10.0 cm length of magnesium ribbon • 30 cm ruler • plastic gloves (see 'Safety considerations') • scissors

Access to:

• a top-pan balance reading to *at least* two decimal places • 2 mol dm^{-3} hydrochloric acid

Safety considerations

- Make sure you have read the advice in the Safety section at the beginning of this book and listen to any advice from your teacher before carrying out this investigation.
- You must wear eye protection at all times.
- Magnesium is highly flammable.
- Hydrogen is a flammable gas.
- 2 mol dm^{-3} hydrochloric acid is an irritant.
- Steel wool sometimes splinters, so use gloves if you have sensitive skin.
- If you are using a glass measuring cylinder for collecting the gas or a gas syringe, then take care when clamping it because over-tightening could shatter the glass.

Method

1. Get a 10.0 cm length of magnesium ribbon and *gently* clean it using the steel wool.
2. Weigh the cleaned magnesium ribbon and record its mass.

   Mass of ribbon .... g
3. Cut the ribbon into 2 × 0.5 cm, 2 × 1.0 cm, 2 × 1.5 cm and 2 × 2.0 cm lengths.
4. From your mass for 10.0 cm of ribbon, estimate the masses of the 1.0 cm, 1.5 cm and 2.0 cm lengths.

   Estimated mass of 1.0 cm lengths .... g

   Estimated mass of 1.5 cm lengths .... g

   Estimated mass of 2.0 cm lengths .... g

5 Depending on which gas-collecting system you are going to use, set up your apparatus as shown in Figure 1.2.

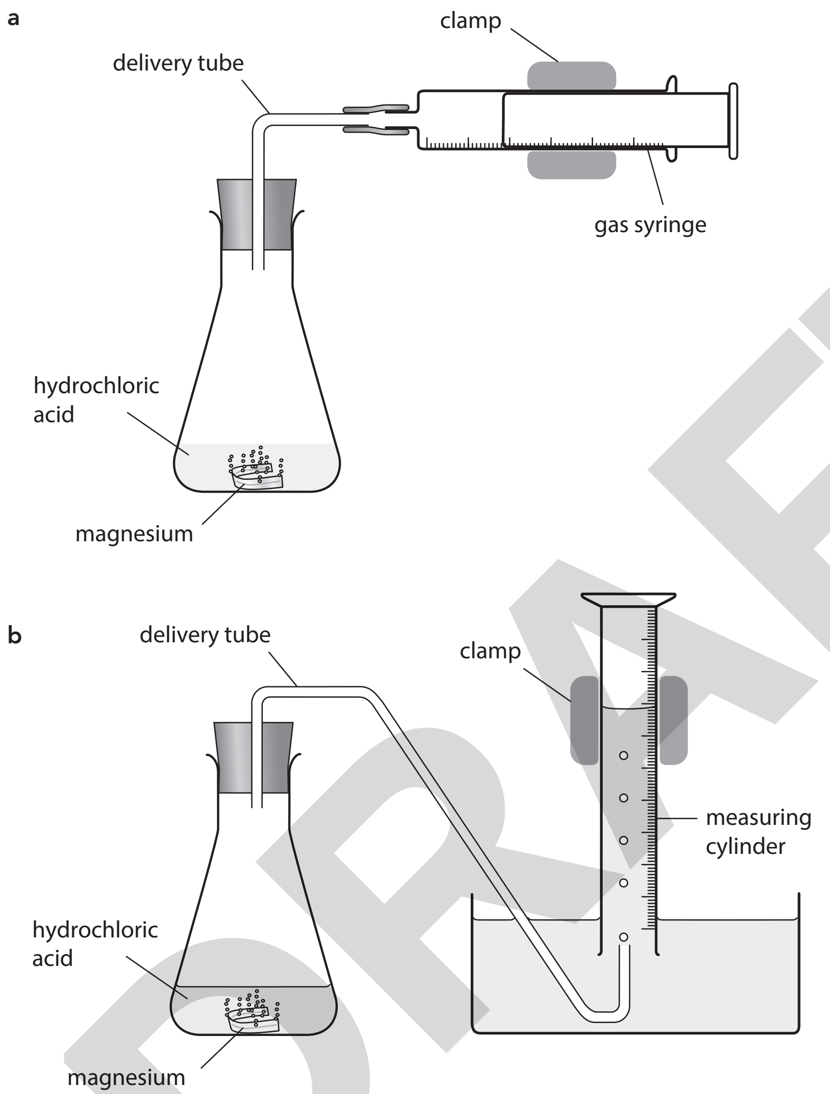

Figure 1.2: Different ways of collecting gases

6 Measure out 25.0 cm^3 of hydrochloric acid into the conical flask.

- **a** Set up the apparatus ready for the measurement of a gas.
- **b** Add one of the 1 cm lengths of magnesium ribbon to the acid, quickly replace the bung, and start collecting the gas.
- **c** Continually swirl the flask because the magnesium will stick to its sides.
- **d** When the reaction is finished, record the volume of gas produced.

  Volume of gas given by a 1.0 cm length of ribbon = ... cm^3

7 Repeat step 6 with all the other known lengths of magnesium ribbon.

Volume of gas (from 1.0 cm of ribbon) .... cm^3

Volume of gas (from 1.5 cm of ribbon) .... cm^3

Volume of gas (from 1.5 cm of ribbon) .... cm^3

Volume of gas (from 2.0 cm of ribbon) .... cm^3

Volume of gas (from 2.0 cm of ribbon) .... cm^3

Results

Use Table 1.1 to record the masses of the ribbon used and the corresponding volumes of hydrogen produced.

Length of Mg ribbon/cm	Mass of Mg/g	Volume of gas produced/cm^3		
		Experiment 1	Experiment 2	Average
0.5 cm				
1.0 cm				
1.5 cm				
2.0 cm				

Table 1.1: Results table

Analysis, conclusion and evaluation

a Plot a graph of mass of magnesium along the horizontal axis (x-axis) against the volume of gas up the vertical axis (y-axis). You should use at least three-quarters of the space available on the graph.

- Discard any results that are obviously wrong.
- Draw a best-fit line through your points.

b Using your graph, calculate the mass of magnesium that gives 24.0 cm^3 of hydrogen gas.

....

....

....

c From this value, calculate the number of moles of magnesium that give this volume of gas and use $A_r = \frac{\text{mass of magnesium}}{\text{number of moles}}$ to find the relative atomic mass of magnesium. Assume that under the conditions of the experiment, 1 mol of gas occupies 24 dm^3 (or 24 000 cm^3).

....

....

....

d Compare your value for A_r with the value given in your Periodic Table.
Using the following formula, calculate the percentage error in your result.

$$\text{Percentage error} = \frac{\text{actual value} - \text{experimental value}}{\text{actual value}} \times 100$$

....

....

....

e What was the maximum error for the top-pan balance that you used?

$$\text{The percentage error for your weighing} = \frac{\text{maximum error}}{\text{mass weighed out}} \times 100\%$$

....

....

f Calculate the percentage error from your measurements of lengths of magnesium ribbon.

....

....

TIP

Look back at the Practical skills chapter to see how to calculate the percentage error from your readings.

TIP

The ruler measures to 1 mm and the maximum error is ± 0.5 mm or 0.05 cm. Therefore, a 2 cm length is really 2.0 ± 0.05 mm and the percentage error $= \frac{0.05}{2.0} \times 100\% = 2.5\%$

g Using this information, calculate the total error from your length measurements. Remember, you made just one weighing but several volume and length measurements and these should be added up.

i Calculated error from length measurements:

....

....

ii Possible errors from volume measurements:

....

....

....

....

iii Total possible percentage error from apparatus readings:

....

....

....

h What other factors could limit the **accuracy** of your results and contribute to the error?

....

....

....

KEY WORD

accuracy (of measurement): How close a measurement is to its true value.

Practical investigation 1.3: Percentage composition of a mixture of sodium hydrogencarbonate and sodium chloride

In this practical, you will investigate the percentage composition of a mixture of sodium hydrogencarbonate and sodium chloride using an acid–base titration.

YOU WILL NEED

Equipment:

• 150 cm^3 conical flask • 250 cm^3 volumetric flask • wash bottle of distilled water • burette stand • 25 cm^3 pipette • white tile • 250 cm^3 beaker • 100 cm^3 beaker • stirring rod • small dropper • small filter funnel for burette and larger one for volumetric flask • 50 cm^3 burette • weighing boat

Access to:

• top-pan balance reading to two, or ideally, three decimal places • mixture of sodium hydrogencarbonate and sodium chloride • 0.100 $mol\,dm^{-3}$ hydrochloric acid • methyl orange indicator and dropper • distilled water

Safety considerations

- Make sure you have read the advice in the Safety section at the beginning of this book and listen to any advice from your teacher before carrying out this investigation.
- Eye protection must be worn at all times.
- Hydrochloric acid is an irritant.
- Methyl orange is poisonous. If you get it on your skin, wash it off immediately.

Part 1: Making up the solution of the mixture

Method

1 Weigh out 1.90–2.10 g of the mixture of sodium hydrogencarbonate and sodium chloride.

Weight of this mixture .... g

2 Dissolve this solid sample in distilled water and make up to a total volume of 250 cm^3 in your volumetric flask as described in the Practical skills chapter.

Part 2: The titrations

Method

1 Titrate 25 cm^3 samples of this solution against the standard 0.100 $mol\,dm^{-3}$ hydrochloric acid.

Use methyl orange as the indicator.

2 You should look back at the Practical skills chapter to remind yourself how to do this.

Results

Complete Table 1.2.

	Rough titration/cm^3	First accurate titration/cm^3	Second accurate titration/cm^3	Third accurate titration/cm^3
Final burette reading/cm^3				
Starting burette reading/cm^3				
Titre/cm^3				

Table 1.2: Results table

Analysis, conclusion and evaluation

a Identify the concordant titres and write the average of these values.

Concordant titres = ... cm^3 and .... cm^3

Average of concordant titres = ... cm^3

Using the data collected, you can calculate the number of moles of sodium hydrogencarbonate present in your sample. You can then calculate the mass of this compound and, from that, the composition of the mixture.

The equation for the reaction between hydrochloric acid and sodium hydrogencarbonate is:

$$NaHCO_3(aq) + HCl(aq) \rightarrow NaCl(aq) + CO_2(g) + H_2O(l)$$

b Calculate the following:

i The volume of 0.100 $mol\,dm^{-3}$ hydrochloric acid needed to react completely with the sodium hydrogencarbonate present in 25 cm^3 of the mixture

= ………………… cm^3

ii The number of moles of hydrochloric acid reacting = …………….. ×

………….. mol = number of moles of sodium hydrogencarbonate present in

25.00 cm^3 of solution = ………………… mol

iii Mass of sodium hydrogencarbonate present (Remember $m = n \times M_r$)

= ……………………….. g

iv Total mass of mixture = …………………………. g

v Therefore, mass of sodium chloride present in mixture = ………………… g

vi Percentage of sodium hydrogencarbonate present in mixture

= ……………………

vii What is the actual percentage composition of the mixture? (Ask your teacher/supervisor.) Answer = …………………….. %

c You should also calculate the percentage error in your results, as you did in Investigation 1.2.

$$\text{Percentage error} = \frac{\text{actual value} - \text{experimental value}}{\text{actual value}} \times 100$$

d Identify and calculate the **systematic errors** in your experiment from the following apparatus:

i The top-pan balance

……………………………………………………………………………………

……………………………………………………………………………………

ii The pipette

……………………………………………………………………………………

……………………………………………………………………………………

iii The burette readings

……………………………………………………………………………………

……………………………………………………………………………………

e Identify the **random errors** in your experiment.

……………………………………………………………………………………

……………………………………………………………………………………

f What was the main contribution (if any) to your percentage error?

……………………………………………………………………………………

……………………………………………………………………………………

KEY WORDS

systematic errors: Errors due to data being inaccurate in a consistent way, e.g. all the results in an experiment measuring gas volumes are 1 cm^3 higher than they should be. Systematic errors are often caused by errors in the experimental procedure or equipment.

random errors: Errors that are due to chance changes in the experiment or by the experimenter, e.g. the electric current in a water-bath heater shuts down for a short time, human error in a one-time misreading.

g How could this be overcome?

....

....

TIP

Remember that in each titration you make two readings, each with a possible error of $\pm0.05\ cm^3$. So, for example, a titre of $20.00\ cm^3$ has a maximum possible error of $\pm0.10\ cm^3$.

Practical investigation 1.4: Relative atomic mass of calcium by two different methods – molar volume and titration

The equation for the reaction between calcium and water is:

$$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$$

This reaction can be used to find the relative atomic mass of calcium by measuring the number of moles of hydrogen produced by a known mass of calcium. The number of moles of calcium (n) can then be calculated using $A_r = m/n$

The calcium hydroxide formed in the reaction can then be titrated against standard hydrochloric acid.

YOU WILL NEED

Equipment:

• apparatus for measuring gas volumes (as used in Investigation 1.2) • small filter funnel for burette • $50.00\ cm^3$ burette • weighing boat • $150\ cm^3$ conical flask • wash bottle of distilled water • burette stand • $25.00\ cm^3$ pipette • white tile • $250\ cm^3$ beaker • $25.0\ cm^3$ measuring cylinder • methyl orange indicator in dropper bottle

Access to:

• top-pan balance reading to *at least* two decimal places • $0.200\ mol\,dm^{-3}$ hydrochloric acid • *fresh* calcium granules • distilled water

Safety considerations

- Make sure you have read the advice in the Safety section at the beginning of this book and listen to any advice from your teacher before carrying out this investigation.
- You must wear eye protection at all times.
- Calcium reacts vigorously with water. Do not handle it with bare hands.
- Hydrogen is a flammable gas.
- $0.2\ mol\,dm^{-3}$ hydrochloric acid is an irritant.
- If you are using a glass measuring cylinder for collecting the gas or a gas syringe, then take care when clamping it. Over-tightening the clamp could shatter the glass.
- Calcium hydroxide is an alkali and should be regarded as being corrosive. If you get any on your skin then wash it off immediately.

Part 1: Determination by molar volume

Method

1 Setup your apparatus for reacting calcium with water and collecting the gas formed during the reaction. Use either of the two arrangements shown in Figure 1.2.

2 Measure 25 cm^3 of distilled water and pour it into the conical flask.

3 Weigh out between 0.040 g and 0.080 g of calcium.

4 Make sure that your gas-collection apparatus is ready.

5 Add the calcium granules to the conical flask and quickly replace the stopper. Swirl the flask vigorously to make sure that all the calcium has reacted.

6 When the reaction is finished note the volume of gas evolved and record it in Table 1.3.

Results

Mass of calcium/g	Volume of hydrogen/ cm^3	Burette reading for hydrochloric acid/cm^3	
		2nd	
		1st	
		Titre	
		2nd	
		1st	
		Titre	
		2nd	
		1st	
		Titre	

Table 1.3: Results table

Analysis, conclusion and evaluation

a Assume that 1 mol of a gas occupies 24 000 cm^3 at room temperature and pressure.

i Calculate the number of moles of hydrogen formed in your first experiment.

……………………………………………………………………………………..

ii From this, calculate the number of moles of calcium.

……………………………………………………………………………………..

iii Calculate the relative atomic mass of calcium.

……………………………………………………………………………………..

……………………………………………………………………………………..

....

....

....

b Using the data shown in your Periodic Table, calculate the percentage error in your result.

$$\text{Percentage error} = \frac{\text{actual value} - \text{experimental value}}{\text{actual value}} \times 100\%$$

....

....

c Systematic errors: calculate the percentage errors in your apparatus.

i The weighing out of the calcium.

....

....

ii The measurement of gas volume.

....

....

iii Identify any random errors in the method.

....

....

iv Are there any improvements you would make to this method?

....

....

> **TIP**
>
> Look back at the Practical skills chapter for full details on carrying out titrations.

Part 2: Determination by titration

Method

1 Remove the flask from the gas collection apparatus and wash any liquid *and* white solid on the sides into the solution.

2 **a** Fill your burette to near the zero mark with 0.200 $mol\,dm^{-3}$ hydrochloric acid.

b Put a white tile under the burette tap.

c Add a few drops of methyl orange indicator to the calcium hydroxide in the conical flask. *There are no opportunities for a rough titration.*

3 a Add the acid to the flask and after each addition swirl the flask vigorously.

b When the indicator shows signs of a colour change to orange red, add the acid more slowly – *one drop at a time* until an orange colour is obtained.

c Note the final burette reading.

4 a Wash the flask thoroughly using tap water and then rinse it with distilled water.

b Repeat these steps twice using a new mass of calcium each time.

Results

Complete Table 1.4.

Mass of calcium/g	Volume of hydrogen/cm^3	Burette reading for hydrochloric acid/cm^3	
		2nd	
		1st	
		Titre	
		2nd	
		1st	
		Titre	
		2nd	
		1st	
		Titre	

Table 1.4: Results table

Analysis, conclusion and evaluation

a Calculate the number of moles of hydrochloric acid reacting with the calcium hydroxide.

i From this value, calculate the number of moles of calcium hydroxide, and therefore the number of moles of calcium.

ii Calculate the relative atomic mass of calcium.

Repeat these calculations if you have more than one set of results.

………………………………………………………………………………………………

………………………………………………………………………………………………

………………………………………………………………………………………………

………………………………………………………………………………………………

………………………………………………………………………………………………

b Using the value shown in your Periodic Table, calculate the percentage error in your results for the following:

i Weighing out the calcium.

....

....

ii The titrations.

....

....

....

....

iii Systematic errors: calculate the total percentage errors in your measurements.

....

....

iv Random errors: identify these in this method.

....

....

c Are there any improvements you would make to this method?

....

....