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Chemistry 1

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Chemistry 1

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CAMBRIDGE
UNIVERSITY PRESS

PUBLISHED BY THE PRESS SYNDICATE OF THE UNIVERSITY OF CAMBRIDGE
The Pitt Building, Trumpington Street, Cambridge, United Kingdom

CAMBRIDGE UNIVERSITY PRESS
The Edinburgh Building, Cambridge CB2 2RU, UK
40 West 20th Street, New York, NY 10011–4211, USA
477 Williamstown Road, Port Melbourne, VIC 3207, Australia
Ruiz de Alarcón 13, 28014 Madrid, Spain
Dock House, The Waterfront, Cape Town 8001, South Africa

http:// www.cambridge.org

First published 2000
Reprinted 2001 (twice), 2002

Printed in the United Kingdom at the University Press, Cambridge

Typeface Swift *System* QuarkXPress®

A catalogue record for this book is available from the British Library

ISBN 0 521 78778 5 paperback

Produced by Gecko Ltd, Bicester, Oxon

Front cover photograph: Reaction synthesis, Images Colour Library

Introduction

Cambridge Advanced Sciences

The *Cambridge Advanced Sciences* series has been developed to meet the demands of all the new AS and A level science examinations. In particular, it has been endorsed by OCR as providing complete coverage of their specifications. The AS material is presented as a single text for each of biology, chemistry and physics. Material for the A2 year comprises six books in each subject: one of core material and one for each option. Some material has been drawn from the existing *Cambridge Modular Sciences* books; however, the majority is entirely new.

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The presentation of units

You will find that the books in this series use a bracketed convention in the presentation of units within tables and on graph axes. For example, ionisation energies of 1000 kJ mol^{-1} and 2000 kJ mol^{-1} will be represented in this way:

Measurement	Ionisation energy (kJ mol^{-1})
1	1000
2	2000

OCR examination papers use the solidus as a convention, thus:

Measurement	Ionisation energy / kJ mol^{-1}
1	1000
2	2000

Any numbers appearing in brackets with the units, for example (10^{-5} mol dm^{-3} s^{-1}), should be treated in exactly the same way as when preceded by the solidus, /10^{-5} mol dm^{-3} s^{-1}.

Chemistry 1 – the AS chemistry text

Chemistry 1 is all that is needed to cover the whole of the AS chemistry material. It is divided into three parts which correspond to the modules Foundation chemistry, Chains and rings, and How far, how fast? It is designed to be accessible to students with a double-award science GCSE background. This book combines entirely new text and illustrations with revised and updated material from *Foundation Chemistry*, *Chains and Rings*, *How Far? How Fast?*, and *Trends and Patterns*, formerly available in the *Cambridge Modular Sciences* series.

Part 1, Foundation chemistry, is a combination of new material and sections based on chapters from *Foundation Chemistry* and *Trends and Patterns*. Specifically, chapters 1–3 in *Foundation Chemistry* have been adapted to form chapters 1–3 in this book. Chapter 4, 'Periodic patterns', is a combination of *Foundation Chemistry* chapter 5 and a new section on successive ionisation energies. Chapter 5 introduces the elements and compounds of Group II of the Periodic Table, pulling together sections from *Trends and Patterns* chapter 4, *Foundation Chemistry* chapter 5 and new material dealing with the formation of salts and aspects of chalk and lime chemistry. Chapter 6, on the topic of the Group VII elements and their compounds, uses material previously found in *Trends and Patterns* chapter 6 and includes a new section on identifying halides.

Chapter 7 begins Part 2 and introduces the subject of organic chemistry. Much of this chapter has been revised from chapter 1 of *Chains and Rings*. In addition, there is a section on the naming of organic compounds which is based on material from *Foundation Chemistry* chapter 6. Chapters 8, 9 and 10 introduce the chemistry of hydrocarbons in the form of fuels, alkanes and alkenes respectively. Sections of *Chains and Rings* chapter 2 and *Foundation Chemistry* chapter 6 are used in these three chapters, enhanced by new material, including sections on the separation of hydrocarbons for use as fuels and the chemical properties of alkanes. Chapter 11, 'Alcohols', combines *Chains and Rings* chapter 4 with new text on the physical properties of alcohols and structural identification using infrared spectroscopy. Part 2 concludes with 'Halogenoalkanes', which is adapted from chapter 3 of *Chains and Rings*.

Part 3 begins with chapter 13, 'Enthalpy changes'. It makes use of *Foundation Chemistry* chapter 4 and *How Far? How Fast?* chapter 1. In addition there is a large new section on the construction and use of enthalpy cycles. *How Far? How Fast?* chapters 3 and 4 respectively form the basis of the final two chapters. However, there is a substantial amount of entirely new material in chapter 14, on measuring the rates of reactions, collision theory and the concept and application of catalysis; and in chapter 15, 'Equilibria', on Le Chatelier's principle, the Haber process, and finally acids and their reactions.

Acknowledgements

2, 26, 69, 70*l*, 77, 78, 91, 92, 93*t*, 113, 114, 118, 120*r*, 129, 130, 131, 139, 165*b*, 166*t*, 168, 170, 188, Andrew Lambert; 4, 7, University of Cambridge Cavendish Laboratory, Madingley Road, Cambridge; 5, Manchester University/Science and Society Picture Library, Science Museum, London; 18, NSSDC/NASA; 20*t*, 22, 93*b*, courtesy of Aventis Pasteur MSD; 20*b*, 41(*gold*), 44*l*, Dr B. Booth, GeoScience Features Picture Library; 21, Heine Schneebeli/Science Photo Library; 31*b*, Photo Library International; 31*tl*, cr, A. Fisher/GeoScience Features Picture Library; 31*tr*, *cl*, 35, 41*tl*, *tr*, 44*r*, 46, 51*tl*, *cr*, GeoScience Features Picture Library; 32*t*, NASA/Science Photo Library; 32*b*, courtesy of ICI; 34*tl*, The Natural History Museum, London; 34*br*, 71, 124, ©Brian Ratcliff; 36, courtesy of Argonne National Laboratory; 41*c*, *b*, Ancient Art & Architecture Collection; 44*tr*, *br*, 109*r*, 147*tr*, 166*br inset*, Tick Ahearn; 45*l*, 47, 51*bl*, 66*b*, 123*bl*, Peter Gould; 45*r*, Britstock-IFA/TPL; 48*t*, courtesy of Dr Jonathan Goodman, Department of Chemistry, Cambridge University, using the program Eadfrith (©J.M.Goodman, Cambridge University, 1994)/photo by Cambridge University Chemistry Department Photographic Unit; 48*b*, Dr Arthur Lesk, Laboratory of Molecular Biology/Science Photo Library; 51*tc*, Astrid & Hanns–Frieder Michler/Science Photo Library; 51*cb*, Vaughan Fleming/Science Photo Library; 51*br*, Erich Schrempp/Science Photo Library; 51*tr*, *cl*, US Dept of Energy/Science Photo Library; 52, courtesy of the Library & Information Centre of the Royal Society of Chemistry; 53*r*, courtesy of Gordon Woods, Malvern School; 65, 82, 104, 155, Ann Ronan at Image Select; 66*t*, 70*r* Leslie Garland Picture Library; 72, ©Martyn Chillmaid/www.osf.uk.com; 76, Steve Davey/La Belle Aurore; 90, Barry Mayes/Life File; 101, 102, Shell Photo Service; 103, Paolo Koch/Robert Harding Picture Library; 105*t*, Kieran Murray/Ecoscene, 105*b*, 107*r*, Schaffer/Ecoscene; 106t, Martin Bond/Science Photo Library; 106*b*, ©Geoff Miller; 107*l*, US Department of Energy/Science Photo Library; 108*tl*, John Mead/Science Photo Library; 108*tr*, br, Simon Fraser/Science Photo Library; 108*bl*, Wilkinson/Ecoscene; 108*cr*, Helene Rogers/Art Directors & TRIP Photo Library; 109*tl*, reproduced by kind permission of Mercedes-Benz; 109*bl*, FPG Int/Robert Harding Picture Library; 117, Images Colour Library; 120*l*, courtesy of Van den Berg Foods Limited; 123*tr*, E.J. Bent/Ecoscene; 123*br*, James Holmes/Science Photo Library; 128, H. Schmidbauer/Britstock-IFA; 140*b*, 153, 162*b*, 165*t*, 166*b*, Michael Brooke; 140*t*, courtesy of civil Aviation Authority, International Fire Training Centre, UK and Robert Hartness; 141 NASA/Science Photo Library; 146, Astrid & Hanns-Frieder-Michler/Science Photo Library; 147*l*, courtesy of Buxton Lime Industries/photo by SJ Hambleton; 147*br*, Tim Fisher/Life File; 162*a*, Popperfoto; 171, Juliet Coombe/La Belle Aurore; 174, courtesy of Catalytic Systems Division, Johnson Matthey; 176, Budd Titlow/f/stop Pictures; 186, Billingham Ammonia Plant, Terra Nitrogen (UK) Ltd.

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Part 2: Chains and Rings

Part 3: How far, how fast?

Part 1
Foundation Chemistry

CHAPTER 1

Atomic structure

By the end of this chapter you should be able to:

1. recognise and describe *protons*, *neutrons* and *electrons* in terms of their relative charges and relative masses;
2. describe the distribution of *mass* and *charge* within an *atom*;
3. describe the contribution of protons and neutrons to atomic nuclei in terms of *atomic number* and *mass number*;
4. deduce the numbers of protons, neutrons and electrons present in both atoms and *ions* from given atomic and mass numbers;
5. distinguish between *isotopes* on the basis of different numbers of neutrons present;
6. explain the terms *first ionisation energy* and *successive ionisation energies* of an *element* in terms of 1 mole of gaseous atoms or ions;
7. explain that ionisation energies are influenced by *nuclear charge*, *atomic radius* and *electron shielding*;
8. predict the number of electrons in each principal *quantum shell* of an element from its successive ionisation energies;
9. describe the shapes of s and p *orbitals*;
10. describe the numbers and relative energies of the s, p and d orbitals for the principal quantum numbers 1, 2, 3 and also the 4s and 4p orbitals;
11. deduce the *electronic configurations* of atoms up to $Z = 36$ and ions, given the atomic number and charge, limited to s and p *blocks* up to $Z = 36$.

Chemistry is a science of change. Over the centuries people have heated rocks, distilled juices and probed solids, liquids and gases with electricity. From all this activity we have gained a great wealth of new materials – metals, medicines, plastics, dyes, ceramics, fertilisers, fuels and many more (*figure 1.1*). But this creation of new materials is only part of the science and technology of chemistry. Chemists also want to *understand* the changes, to find patterns of behaviour and to discover the innermost nature of the materials.

Our 'explanations' of the chemical behaviour of matter come from reasoning and model-building based on the limited evidence available from

• **Figure 1.1** All of these useful products, and many more, contain chemicals that have been created by applying chemistry to natural materials. Chemists must also find answers to problems caused when people misuse chemicals.

experiments. The work of chemists and physicists has shown us the following:

- All known materials, however complicated and varied they appear, can be broken down into the fundamental substances we call **elements**. These elements cannot be broken down further into simpler substances. So far, about 115 elements are recognised. Most exist in combinations with other elements in **compounds** but some, such as gold, nitrogen, oxygen and sulphur, are also found in an uncombined state. Some elements would not exist on Earth without the artificial use of nuclear reactions. Chemists have given each element a symbol. This symbol is usually the first one or two letters of the name of the element; some are derived from their names in Latin. Some examples are:

Element	*Symbol*
carbon	C
lithium	Li
iron	Fe (from the Latin *ferrum*)
lead	Pb (from the Latin *plumbum*)

- Groups of elements show patterns of behaviour related to their atomic masses. A Russian chemist, Dmitri Mendeleev, summarised these patterns by arranging the elements into a 'Periodic Table'. Modern versions of the Periodic Table are widely used in chemistry. (A Periodic Table is shown in the appendix on page 191 and explained, much more fully, in chapter 4.)
- All matter is composed of extremely small particles (atoms). About 100 years ago, the accepted model for atoms included the assumptions that (i) atoms were tiny particles, which could not be divided further nor destroyed, and (ii) all atoms of the same element were identical. The model had to give way to other models, as science and technology produced new evidence. This evidence could only be interpreted as atoms having other particles inside them – they have an internal structure.

Scientists now believe that there are two basic types of particles – 'quarks' and 'leptons'. These are the building-blocks from which everything is made, from microbes to galaxies. For many explanations or predictions, however, scientists find it helpful to use a model of atomic structure that includes three basic particles in any atom, the **electron**, the **proton** and the **neutron**. Protons and neutrons are made from quarks, and the electron is a member of the family of leptons.

Discovering the electron

Effect of electric current in solutions (electrolysis)

When electricity flows in an aqueous solution of silver nitrate, for example, silver metal appears at the negative electrode (cathode). This is an example of **electrolysis** and the best explanation is that:

- the silver exists in the solution as positively charged particles known as **ions** (Ag^+);
- one silver ion plus one unit of electricity gives one silver atom.

The name 'electron' was given to this unit of electricity by the Irish scientist George Johnstone Stoney in 1891.

Study of cathode rays

At normal pressures gases are usually very poor conductors of electricity, but at low pressures they conduct quite well. Scientists, such as William Crookes, who first studied the effects of passing electricity through gases at low pressures, saw that the glass of the containing vessel opposite the **cathode** (negative electrode) glowed when the applied potential difference was sufficiently high.

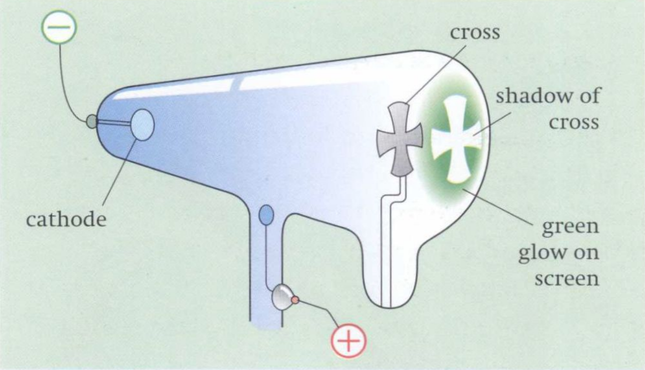
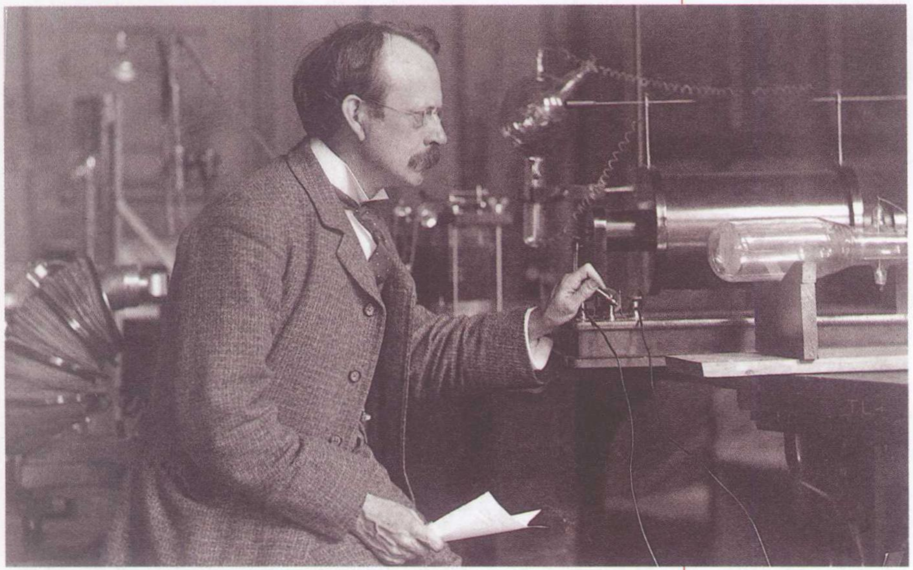


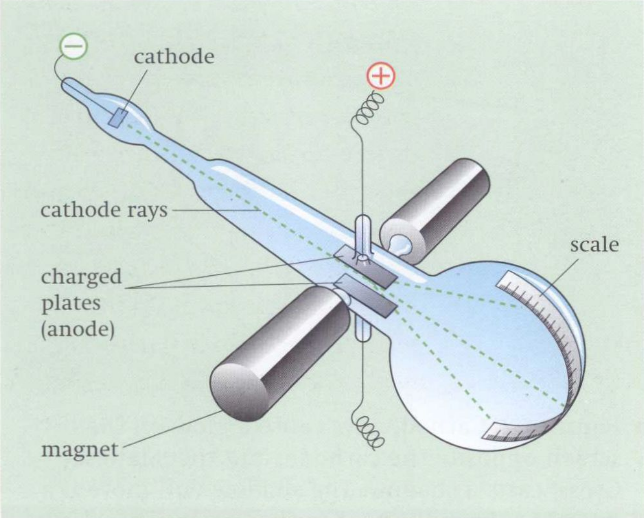
Figure 1.2 Cathode rays cause a glow on the screen opposite the cathode, and the 'Maltese Cross' casts a shadow. The shadow will move if a magnet is brought near to the screen. This shows that the cathode rays are deflected in a magnetic field. The term 'cathode ray' is still familiar today, as in 'cathode-ray oscilloscopes'.

A solid object, placed between the cathode and the glow, cast a shadow (*figure 1.2*). They proposed that the glow was caused by rays coming from the cathode and called these **cathode rays**.

For a while there was some argument about whether cathode rays are waves, similar to visible light rays, or particles. The most important evidence is that they are strongly deflected in a magnetic field. This is best explained by assuming that they are streams of electrically charged particles. The direction of the deflection (towards the positive pole) shows that the particles in cathode rays are negatively charged.



● **Figure 1.3** Joseph (J. J.) Thomson (1856–1940) using his cathode-ray tube.



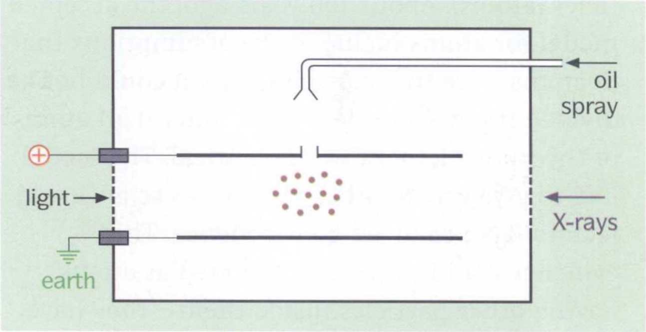
● **Figure 1.4** A drawing of Thomson's apparatus. The electrons move from the hot cathode (negative) through slits in the anode (positive).

J. J. Thomson's *e/m* experiment

The great leap in understanding came in 1897, at the Cavendish Laboratory in Cambridge (*figures 1.3 and 1.4*). J. J. Thomson measured the deflection of a narrow beam of cathode rays in both magnetic and electric fields. His results allowed him to calculate the charge-to-mass ratio (e/m) of the particles. Their charge-to-mass ratio was found to be exactly the same, whatever gas or type of electrodes were used in the experiment. The cathode-ray particles had a tiny mass, only approximately 1/2000 th of the mass of a hydrogen atom. Thomson then decided to call them **electrons** – the name suggested earlier by Stoney for the 'units of electricity'.

Millikan's 'oil-drop' experiment

The electron charge was first measured accurately in 1909 by the American physicist Robert Millikan using his famous 'oil-drop' experiment (*figure 1.5*). He found the charge to be 1.602×10^{-19} C (coulombs). The mass of an electron was calculated to be 9.109×10^{-31} kg, which is 1/1837 th of the mass of a hydrogen atom.



● **Figure 1.5** Robert Millikan's 'oil-drop' experiment. Millikan gave the oil drops negative charge by spraying them into air ionised by X-ray bombardment. He adjusted the charge on the plates so that the upward force of attraction equalled the downward force due to gravity, and a drop could remain stationary. Calculations on the forces allowed him to find the charges on the drops. These were multiples of the charge on an electron.

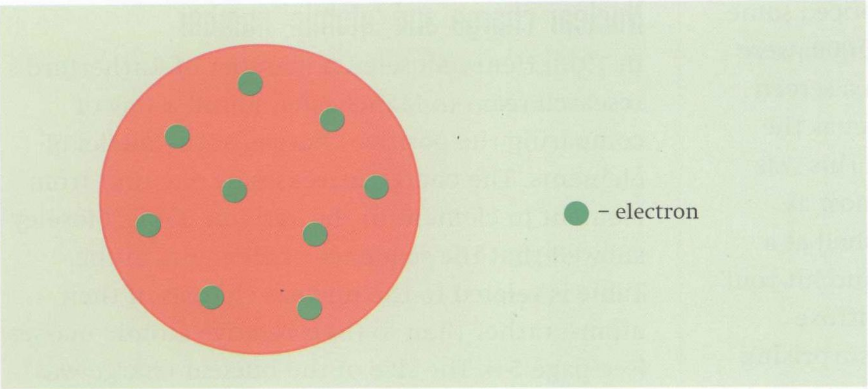


Figure 1.6 J. J. Thomson's 'plum-pudding' model of the atom. The electrons (plums) are embedded in a sphere of uniform positive charge.

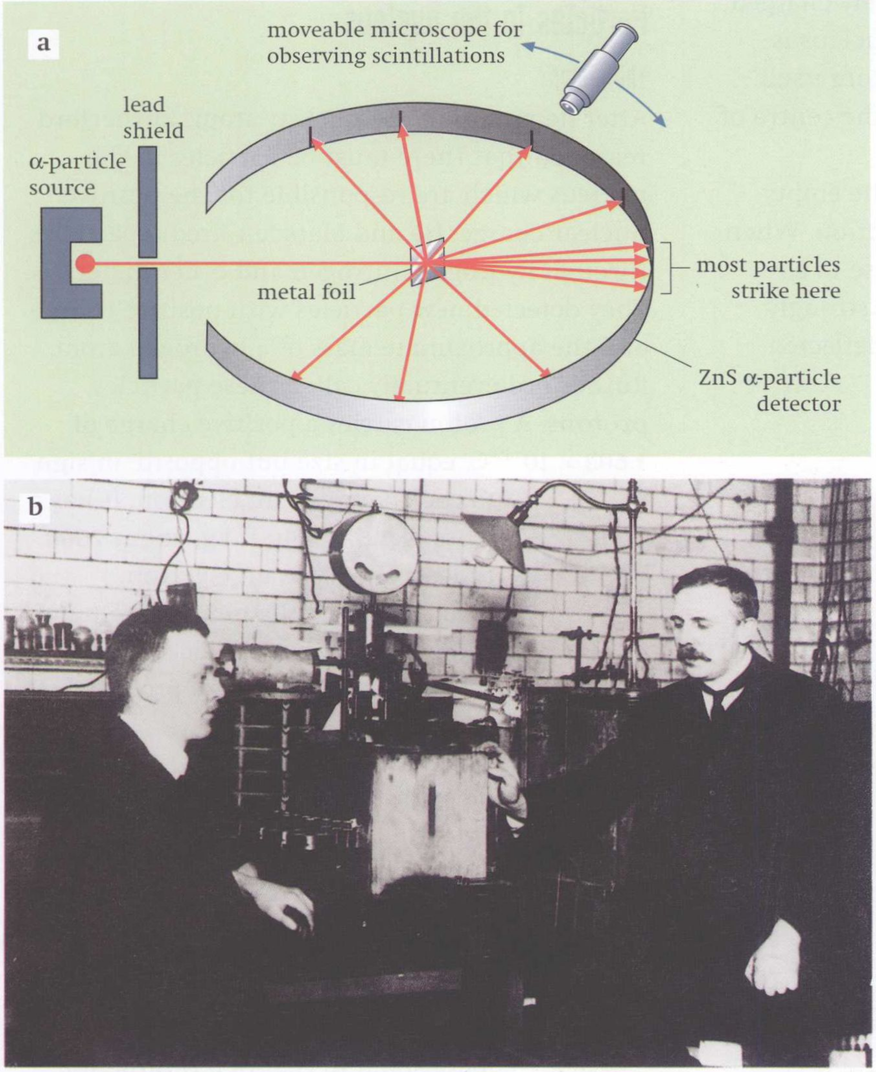


Figure 1.7 Geiger and Marsden's experiment, which investigated how α-particles are deflected by thin metal foils.
a A drawing showing the arrangement of the apparatus.
b Ernest Rutherford (right) and Hans Geiger using their apparatus for detecting α-particle deflections. Interpretation of the results led Rutherford to propose the nuclear model for atoms.

Discovering protons and neutrons

New atomic models: 'plum-pudding' or 'nuclear' atom

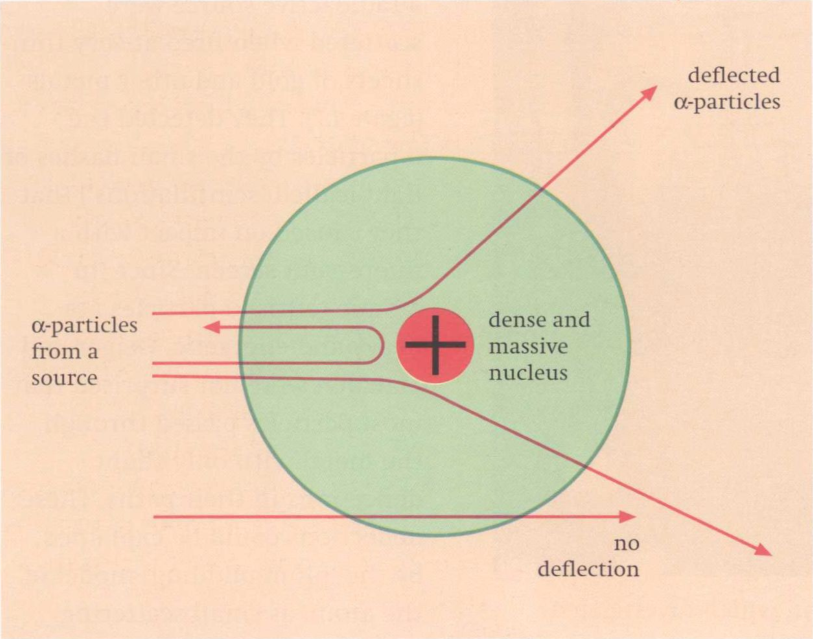
The discoveries about electrons demanded new models for atoms. If there are negatively charged electrons in all electrically neutral atoms, there must also be a positively charged part. For some time the most favoured atomic model was J. J. Thomson's 'plum-pudding', in which electrons (the plums) were embedded in a 'pudding' of positive charge (*figure 1.6*).

Then, in 1909, came one of the experiments that changed everything. Two members of Ernest Rutherford's research team in Manchester University, Hans Geiger and Ernest Marsden, were investigating how α-particles (α is the Greek letter alpha) from a radioactive source were scattered when fired at very thin sheets of gold and other metals (*figure 1.7*). They detected the α-particles by the small flashes of light (called 'scintillations') that they caused on impact with a fluorescent screen. Since (in atomic terms) α-particles are heavy and energetic, Geiger and Marsden were not surprised that most particles passed through the metal with only slight deflections in their paths. These deflections could be explained, by the 'plum-pudding' model of the atom, as small scattering effects caused while the positive α-particles moved through the diffuse mixture of positive charge and electrons.

However, Geiger and Marsden also noticed some large deflections. A few (about one in 20 000) were so large that scintillations were seen on a screen placed on the same side of the gold sheet as the source of positively charged α-particles. This was unexpected. Rutherford said: 'it was almost as incredible as if you had fired a 15-inch shell at a piece of tissue paper and it came back and hit you!'

The plum-pudding model, with its diffuse positive charge, could not explain the surprising Geiger–Marsden observations. However, Rutherford soon proposed his convincing **nuclear model** of the atom. He suggested that atoms consist largely of empty space and that the mass is concentrated into a very small, positively charged, central core called the **nucleus**. The nucleus is about 10 000 times smaller than the atom itself – similar in scale to a marble placed at the centre of an athletics stadium.

Most α-particles will pass through the empty space in an atom with very little deflection. When an α-particle approaches on a path close to a nucleus, however, the positive charges strongly repel each other and the α-particle is deflected through a large angle (*figure 1.8*).



- **Figure 1.8** Ernest Rutherford's interpretation of the Geiger–Marsden observations. The positively charged α-particles are deflected by the tiny, dense, positively charged nucleus. Most of the atom is empty space.

Nuclear charge and 'atomic' number

In 1913, Henry Moseley, a member of Rutherford's research team in Manchester, found a way of comparing the positive charges of the nuclei of elements. The charge increases by one unit from element to element in the Periodic Table. Moseley showed that the sequence of elements in the Table is related to the nuclear charges of their atoms, rather than to their relative atomic masses (see page 54). The size of the nuclear charge was then called the **atomic number** of the element. Atomic number defined the position of the element in the Periodic Table.

Particles in the nucleus

The proton

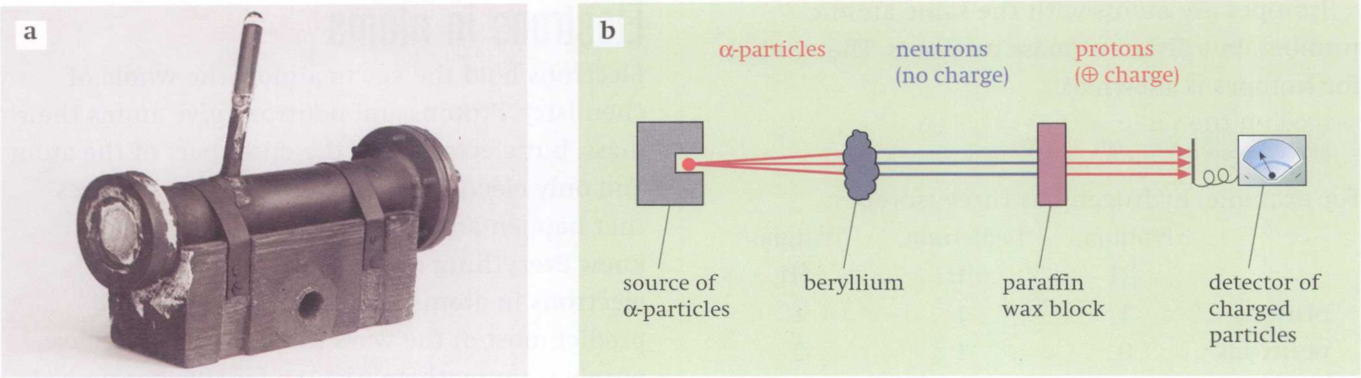
After he proposed the nuclear atom, Rutherford reasoned that there must be particles in the nucleus which are responsible for the positive nuclear charge. He and Marsden fired α-particles through hydrogen, nitrogen and other materials. They detected new particles with positive charge and the approximate mass of a hydrogen atom. Rutherford eventually called these particles **protons**. A proton carries a positive charge of 1.602×10^{-19} C, equal in size but opposite in sign to the charge on an electron. It has a mass of 1.673×10^{-27} kg, about 2000 times as heavy as an electron.

Each electrically neutral atom has the same number of electrons outside the nucleus as there are protons within the nucleus.

The neutron

The mass of an atom, which is concentrated in its nucleus, cannot depend only on protons; usually the protons provide around half of the atomic mass. Rutherford proposed that there is a particle in the nucleus with a mass equal to that of a proton but with zero electrical charge. He thought of this particle as a proton and an electron bound together.

Without any charge to make it 'perform' in electrical fields, detection



● **Figure 1.9**

a Using this apparatus, James Chadwick discovered the neutron.

b Drawing of the inside of the apparatus. Chadwick bombarded a block of beryllium with α-particles ($^{4}_{2}He$). No charged particles were detected on the other side of the block. However, when a block of paraffin wax (a compound containing only carbon and hydrogen) was placed near the beryllium, charged particles were detected and identified as protons (H^+). Alpha-particles had knocked neutrons out of the beryllium, and in turn these had knocked protons out of the wax.

of this particle was very difficult. It was not until 12 years after Rutherford's suggestion that, in 1932, one of his co-workers, James Chadwick, produced sufficient evidence for the existence of a nuclear particle with a mass similar to that of the proton but with no electrical charge (*figure 1.9*). The particle was named the neutron.

Atomic and mass numbers

Atomic number (Z)

The most important difference between atoms of different elements is in the number of protons in the nucleus of each atom. The number of protons in an atom determines the element to which the atom belongs. The atomic number of an element shows:

- the number of protons in the nucleus of an atom of that element;
- the number of electrons in a neutral atom of that element;
- the position of the element in the Periodic Table.

Mass number (A)

It is useful to have a measure for the total number of particles in the nucleus of an atom. This is called the **mass number**. For any atom:

- the mass number is the sum of the number of protons and the number of neutrons.

Summary table

Particle name	*Relative mass*	*Relative charge*
electron	negligible	−1
proton	1	+1
neutron	1	0

Isotopes

In Rutherford's model of the atom, the nucleus consists of protons and neutrons, each with a mass of one atomic unit. The relative atomic masses of elements should then be whole numbers. It was thus a puzzle why chlorine has a relative atomic mass of 35.5.

The answer is that atoms of the same element are not all identical. In 1913, Frederick Soddy proposed that atoms of the same element could have different atomic masses. He named such atoms **isotopes**. The word means 'equal place', i.e. occupying the same place in the Periodic Table and having the same atomic number.

The discovery of protons and neutrons explained the existence of isotopes of an element. In isotopes of one element, the number of protons must be the same, but the number of neutrons may be different.

Remember:

atomic number (Z) = number of protons
mass number (A) = number of protons + number of neutrons

Isotopes are atoms with the same atomic number, but different mass numbers. The symbol for isotopes is shown as

$$^{\text{mass number}}_{\text{atomic number}}\text{X} \quad \text{or} \quad ^{A}_{Z}\text{X}$$

For example, hydrogen has three isotopes:

	Protium, $^{1}_{1}\text{H}$	*Deuterium,* $^{2}_{1}\text{H}$	*Tritium,* $^{3}_{1}\text{H}$
protons	1	1	1
neutrons	0	1	2

It is also common practice to identify isotopes by name or symbol plus mass number only. For example, uranium, the heaviest naturally occurring element ($Z = 92$), has two particularly important isotopes of mass numbers 235 and 238. They are often shown as uranium-235 and uranium-238, as U-235 and U-238 or as ^{235}U and ^{238}U.

Numbers of protons, neutrons and electrons

It is easy to calculate the composition of a particular atom or ion:

number of protons $= Z$

number of neutrons $= A - Z$

number of electrons in neutral atom $= Z$

number of electrons in positive ion $= Z -$ charge on ion

number of electrons in negative ion $= Z +$ charge on ion

For example, magnesium is element 12; it is in Group II, so it tends to form doubly charged (2+) ions. The ionised isotope magnesium-25 thus has the full symbol $^{25}_{12}\text{Mg}^{2+}$, and

number of protons $= 12$

number of neutrons $= 13$

number of electrons $= 10$

SAQ 1.1

a What is the composition (numbers of electrons, protons and neutrons) of neutral atoms of the two main uranium isotopes, U-235 and U-238?

b What is the composition of the ions of potassium-40 (K^{+}) and chlorine-37 (Cl^{-})?

(Use the Periodic Table, page 191, for the atomic numbers.)

Electrons in atoms

Electrons hold the key to almost the whole of chemistry. Protons and neutrons give atoms their mass, but electrons are the outer part of the atom and only electrons are involved in the changes that happen during chemical reactions. If we knew everything about the arrangements of electrons in atoms and molecules, we could predict most of the ways that chemicals behave, purely from mathematics. So far this has proved very difficult, even with the most advanced computers – but it may yet happen.

SAQ 1.2

Suggest why the isotopes of an element have the same chemical properties, though they have different relative atomic masses.

What models are currently accepted about how electrons are arranged around the nucleus? The first simple idea – that they just orbit randomly around the nucleus – was soon rejected. Calculations showed that any moving, electrically charged particles, like electrons, would lose energy and fall into the nucleus.

A model you may have used considers the electrons to be arranged in shells. These ‘shells’ correspond to different energy levels occupied by the electrons.

Arrangements of electrons: energy levels and ‘shells’

There was a great advance in atomic theory when, in 1913, the Danish physicist Niels Bohr proposed his ideas about arrangements of electrons in atoms.

Earlier the German physicist Max Planck had proposed, in his ‘Quantum Theory’ of 1901, that energy, like matter, is ‘atomic’. It can only be transferred in packets of energy he called **quanta**; a single packet of energy is a **quantum**. Bohr applied this idea to the energy of electrons. He suggested that, as electrons could only possess energy in quanta, they would not exist in a stable way, anywhere outside the nucleus, unless they were in fixed or ‘quantised’ energy levels. If an electron gained or lost energy, it could move to higher or lower energy levels, but not somewhere

		Number of electrons in shell		
	Atomic number	$n = 1$	$n = 2$	$n = 3$
H	1	1		
He	2	2		
Li	3	2	1	
Be	4	2	2	
B	5	2	3	
C	6	2	4	
N	7	2	5	
O	8	2	6	
F	9	2	7	
Ne	10	2	8	
Na	11	2	8	1

- **Table 1.1** Electronic configurations of the first 11 elements in the Periodic Table.

in between. It is a bit like climbing a ladder; you can only stay in a stable state on one of the rungs. You will find that, as you read more widely, there are several names given to these energy levels. The most common name is **shells**.

Shells are numbered 1, 2, 3, 4, etc. These numbers are known as **principal quantum numbers** (symbol n). Such numbers correspond to the numbers of rows (or Periods) in the Periodic Table.

We can now write the simple electronic configurations as shown in *table 1.1*. Remember that the atomic number tells us the number of electrons present in an atom of the element. For a given element, electrons are added to the shells as follows:

- up to 2 electrons in shell 1;
- up to 8 electrons in shell 2;
- up to 18 electrons in shell 3.

Some of the best evidence for the existence of electron shells comes from ionisation energies.

Ionisation energy

When an atom loses an electron it becomes a positive ion. We say that it has been **ionised**. Energy is needed to remove electrons and this is generally called **ionisation energy**. More precisely, the **first ionisation energy** of an element is the amount of energy needed to remove one electron from each atom in a mole of atoms of an element in the gaseous state.

The general symbol for ionisation energy is ΔH_i and for a first ionisation energy it is ΔH_{i1}. The process may be shown by the example of calcium as:

$$Ca(g) \rightarrow Ca^{+}(g) + e^{-}; \qquad \Delta H_{i1} = +590\,kJ\,mol^{-1}$$

(If the symbols seem unfamiliar at this stage, see page 148.)

The energy needed to remove a second electron from each ion in a mole of gaseous ions is the **second ionisation energy**. For calcium:

$$Ca^{+}(g) \rightarrow Ca^{2+}(g) + e^{-}; \qquad \Delta H_{i2} = +1150\,kJ\,mol^{-1}$$

Note that the second ionisation energy is much larger than the first. The reasons for this are discussed on page 10.

We can continue removing electrons until only the nucleus of an atom is left. The sequence of first, second, third, fourth, etc. ionisation energies (or **successive ionisation energies**) for the first 11 elements in the Periodic Table are shown in *table 1.2*.

We see that the following hold for any one element:

- The ionisation energies increase. As each electron is removed from an atom, the remaining ion becomes more positively charged. Moving the next electron away from the increased positive charge is more difficult and the next ionisation energy is even larger.
- There are one or more particularly large rises within the set of ionisation energies of each element (except hydrogen and helium).

Ionisation energies of elements are measured mainly by two techniques:

- calculating the energy of the radiation causing particular lines in the emission spectrum of the element;
- using electron bombardment of gaseous elements in discharge tubes.

We now know the ionisation energies of all of the elements.

These data may be interpreted in terms of the atomic numbers of elements and their simple electronic configurations.

Before doing so, we must consider the factors which influence ionisation energies.

		Electrons removed										
		1	**2**	**3**	**4**	**5**	**6**	**7**	**8**	**9**	**10**	**11**
1	H	1310										
2	He	2370	5250									
3	Li	520	7300	11 800								
4	Be	900	1760	14 850	21 000							
5	B	800	2420	3660	25 000	32 800						
6	C	1090	2350	4620	6220	37 800	47 300					
7	N	1400	2860	4580	7480	9450	53 300	64 400				
8	O	1310	3390	5320	7450	11 000	13 300	71 300	84 100			
9	F	1680	3470	6040	8410	11 000	15 200	17 900	92 000	106 000		
10	Ne	2080	3950	6120	9370	12 200	15 200	–	–	–	131 400	
11	Na	510	4560	6940	9540	13 400	16 600	20 100	25 500	28 900	141 000	158 700

● **Table 1.2** Successive ionisation energies for the first 11 elements in the Periodic Table (to nearest 10 kJ mol^{-1}).

Factors influencing ionisation energies

The three strongest influences on ionisation energies of elements are the following:

- *The size of the positive nuclear charge*
  This charge affects all the electrons in an atom. The increase in nuclear charge with atomic number will tend to cause an increase in ionisation energies.
- *The distance of the electron from the nucleus*
  It has been found that, if F is the force of attraction between two objects and d is the distance between them, then

  F is proportional to $1/d^2$
  (the 'inverse square law')

  This **distance effect** means that all forces of attraction decrease rapidly as the distance between the attracted bodies increases. Thus the attractions between a nucleus and electrons decrease as the quantum numbers of the shells increase. The further the shell is from the nucleus, the lower are the ionisation energies for electrons in that shell.
- *The 'shielding' effect by electrons in filled inner shells*
  All electrons are negatively charged and repel each other. Electrons in the filled inner shells repel electrons in the outer shell and reduce the effect of the positive nuclear charge. This is called the **shielding effect**. The greater the shielding effect upon an electron, the lower is the energy required to remove it and thus the lower the ionisation energy.

Consider the example of the successive ionisation energies of lithium. We see a low first ionisation energy, followed by much larger second and third ionisation energies. This confirms that lithium has one electron in its outer shell $n = 2$, which is easier to remove than either of the two electrons in the inner shell $n = 1$. The large increase in ionisation energy indicates where there is a change from shell $n = 2$ to shell $n = 1$.

The pattern is seen even more clearly if we plot a graph of ionisation energies (y axis) against number of electrons removed (x axis). As the ionisation energies are so large, we must use logarithm to base 10 ($\log_{10}$) to make the numbers fit on a reasonable scale. The graph for sodium is shown in *figure 1.10*.

SAQ 1.3

a In *figure 1.10* why are there large increases between the first and second ionisation energies and again between the ninth and tenth ionisation energies?

b How does this graph confirm the suggested simple electronic configuration for sodium of (2,8,1)?

Successive ionisation energies are thus helpful for predicting or confirming the simple electronic configurations of elements. In particular, they