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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

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- 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.

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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.

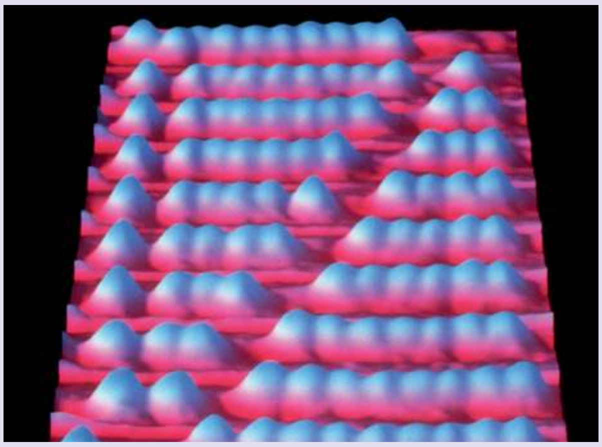

Figure 1.2: Each of the blue peaks in this image is an individual molecule. The molecules can be moved over a copper surface, making this a molecular abacus or counting device.

Questions for discussion

Discuss with another learner or group of learners:

- Why do you think that tiny clusters of atoms are useful for catalysts?

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- How do you think that you could make tiny clusters of metal atoms on a cold surface?

  Tip: 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 Section 2.3). 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 Section 2.3). 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 $^{1}_{1}H$.

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

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

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

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.

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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.

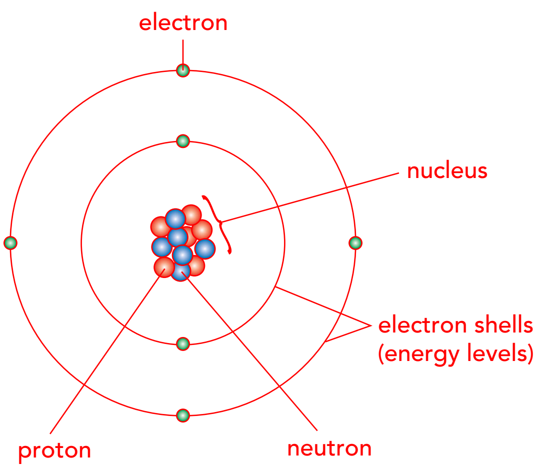

Figure 1.3: A model of a carbon atom. This model is not very accurate but it is useful for understanding what happens to the electrons during chemical reactions.

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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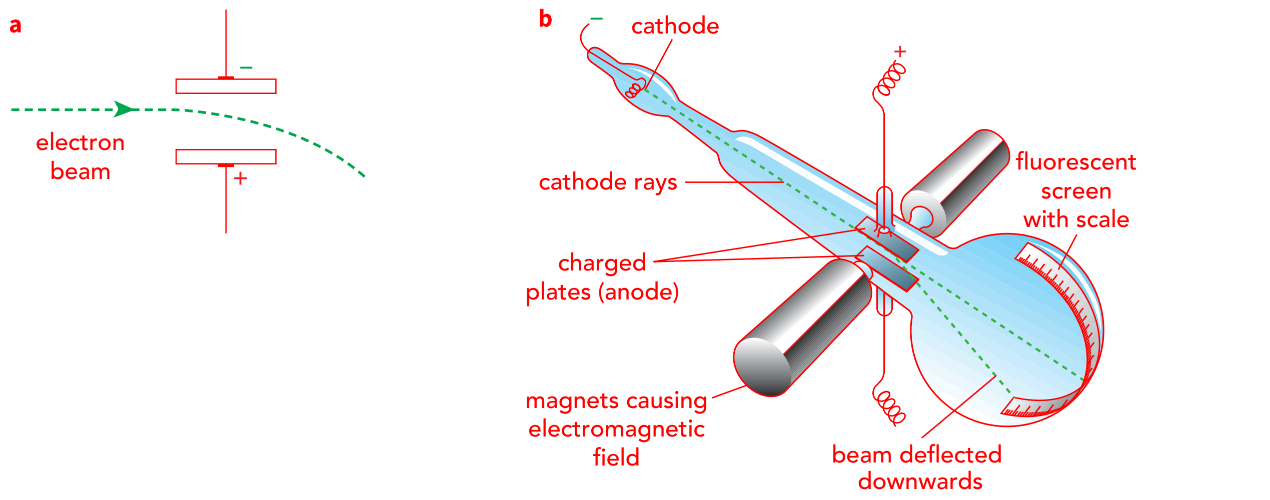

Figure 1.5: a The beam of electrons is deflected away from a negatively charged plate and towards a positively charged plate. **b** The electron beam in a cathode-ray tube is deflected by an electromagnetic field. The direction of the deflection shows us that the electron is negatively charged.

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.

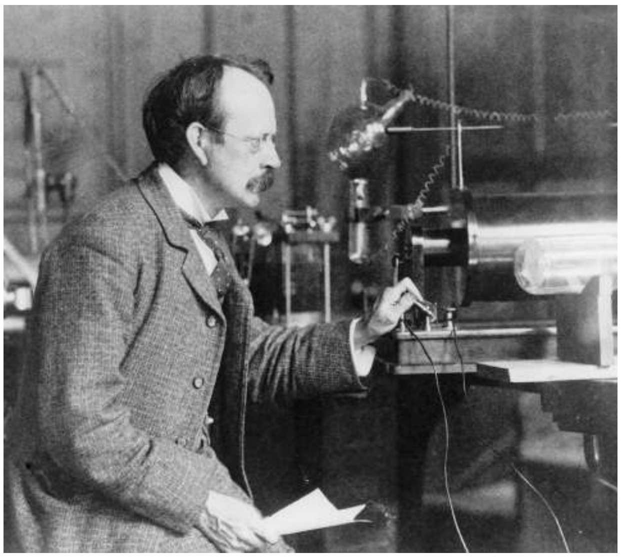

Figure 1.6: J. J. Thomson calculated the charge to mass ratio of electrons. He used results from experiments with electrons in cathode-ray tubes.

KEY WORD

anode: the positive electrode (where oxidation reactions occur).

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

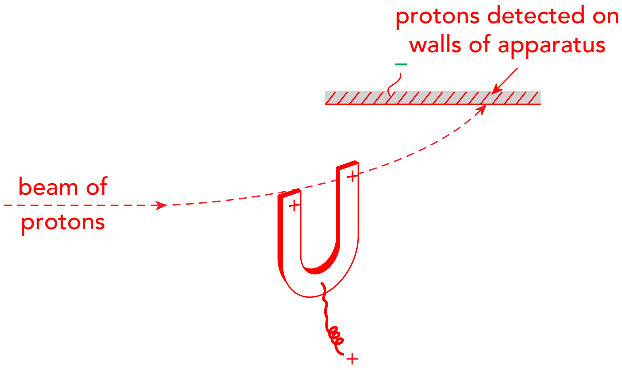

Figure 1.7: A beam of protons is deflected away from a positively charged area. This shows us that protons have a positive charge.

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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.

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.

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.

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).

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 isotopic symbols, ${}^{x}_{y}A$, this is the figure which is subscript.

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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:

number of neutrons = mass number − 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	88	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 $^{11}_{5}B$:

$$\begin{array}{l} \text{nucleon number} \longrightarrow \\ \text{proton number} \longrightarrow \end{array} {}^{11}_{\ 5}\mathrm{B}$$

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.

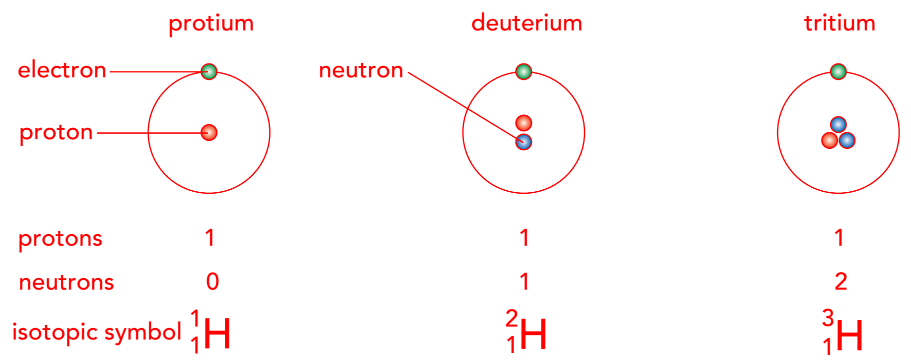

Figure 1.8: The atomic structure and isotopic symbols for the three isotopes of hydrogen.

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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 in Appendix 1 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^-	$\rightarrow$	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	$\rightarrow$	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.

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

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

b $^{58}_{38}Ce^{3+}$

REFLECTION

Read the paragraph in 'Developing an idea: Nanomachines' at the beginning of this chapter 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.

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