1 Atomic structure

Particles in the atom and atomic radius

- Atoms are mostly empty space but contain three different types of particles – **protons**, **neutrons** and **electrons**. The protons and neutrons form a dense nucleus at the centre of the atom with the electrons in 'shells' some distance away from the nucleus.
- Remember that only the protons and neutrons have significant mass, and that the proton carries a single positive charge while the electron carries a single negative charge.
- Remember also that the protons and neutrons are found in the nucleus of the atom and that the electrons surround the nucleus. These particles behave quite differently if they are passed through an electric field. Protons are attracted to the negative electrode and electrons to the positive electrode, while neutrons are undeviated (Figure 1.1).

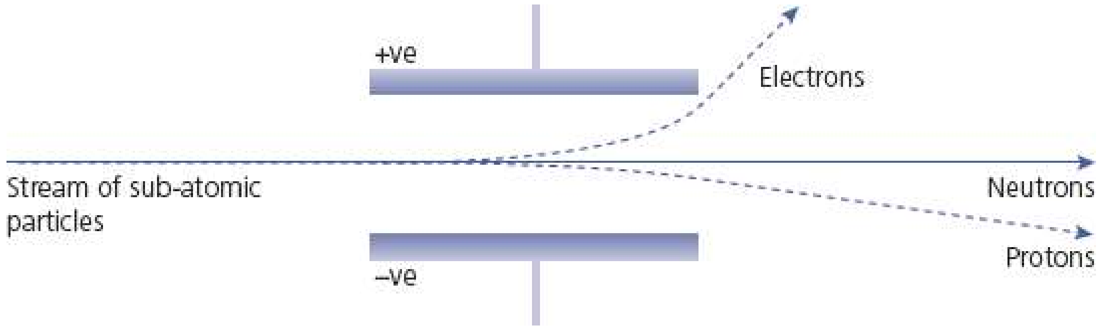

Figure 1.1 Behaviour of protons, neutrons and electrons in an electric field

Charges on particles

Let us look at some different particles that all have the same mass. The numbers of protons, neutrons and electrons in the three particles are shown in Table 1.1. What is the major difference between these three species?

Table 1.1

Particle	Number of protons	Number of neutrons	Number of electrons
A	11	12	10
B	11	12	11
C	11	12	12

- The difference is in the number of electrons each particle possesses, and so the overall charge on the species.
- Because particle A has one more proton than electrons, it has a single positive charge. In B the numbers of protons and electrons are the same, so it is uncharged (neutral). In C there is one more electron than protons, so it has a single negative charge.
- Notice that because all the species have the same number of protons (proton number), they are all forms of the same element, in this case sodium. The two charged species are called **ions**.
- You might be surprised to see sodium (particle A) as an anion, Na^-, but it is theoretically possible (but very unlikely).

KEY TERMS

Protons are particles with a single positive charge found in the nucleus of atoms. They have a relative mass of 1.

Neutrons are uncharged particles, also found in the nucleus of *most* atoms (^1_1H is the exception). They also have a relative mass of 1.

Electrons carry a single negative charge but have no significant mass.

Ions are charged species. A positive ion is called a cation. A negative ion is called an anion.

Isotopes

Proton and neutron numbers

Table 1.2 shows another way in which the numbers of sub-atomic particles can vary.

Table 1.2

Particle	Number of protons	Number of neutrons	Number of electrons
D	12	12	12
E	12	13	12
F	12	14	12

- In this case, it is the number of neutrons that changes, while the element stays the same. These forms of an element are called **isotopes**.
- In Table 1.2, the three species are all isotopes of magnesium.
- The standard way of writing these particles in 'shorthand' form is $^{A}_{Z}X^{n\pm}$. In this form the element symbol is X, A is the nucleon or mass number (the number of protons plus neutrons in the nucleus), Z is the proton or atomic number (the number of protons in the nucleus) and $n\pm$ is the charge (if any) on the particle.

Some elements have atoms with different atomic masses. These atoms have the same number of protons but different numbers of neutrons. They are isotopes and they have identical chemical properties. Most naturally occurring isotopes are stable but some, such as uranium and any artificially produced ones, are unstable and give off radiation.

NOW TEST YOURSELF

1 Write out structures of the six species A–F described in Tables 1.1 and 1.2 using the form $^{A}_{Z}X^{n\pm}$.

Arrangement of electrons in atoms

As the number of protons in the nucleus increases, the masses of atoms increase. After hydrogen, this increase in mass is also due to the neutrons in the nucleus (Table 1.3).

Table 1.3

Element	Protons	Neutrons	Mass number
H	1	0	1
He	2	2	4
Li	3	4	7
Be	4	5	9
B	5	6	11
C	6	6	12

- The addition of electrons to form new atoms is not quite so straightforward because they go into different **orbitals**.
- The electrons also exist in different **energy levels** (sometimes called shells) depending on how close to, or far away from, the nucleus they are.
- The number of protons in a nucleus determines just what the element is. However, it is the *arrangement* of electrons that determines the chemistry of an element and how it forms bonds with other elements.
- So, for example, metal atoms tend to lose electrons, forming positive ions; non-metal atoms tend to accept electrons, forming negative ions.

KEY TERM

Orbitals are regions in space that can hold a certain number of electrons, and which have different shapes.

As more electrons are added, they go successively into orbitals of increasing energy, as shown in Figure 1.2.

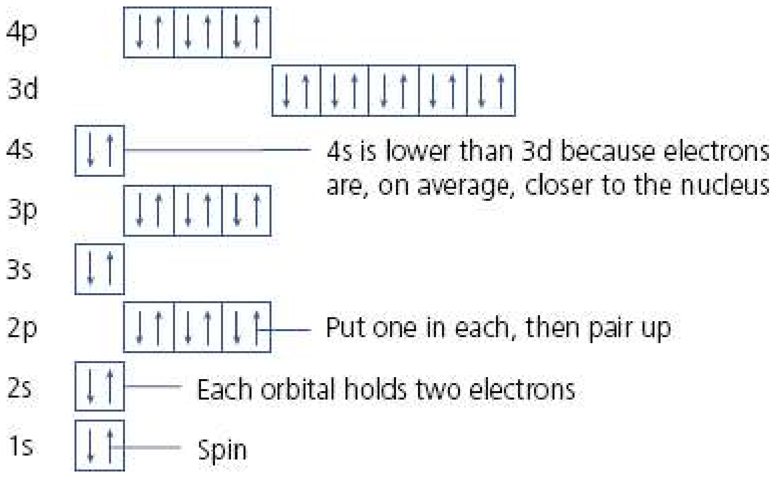

Figure 1.2 Sequence of filling orbitals with electrons

Figure 1.2 illustrates some key points about the arrangement of electrons in atoms.

- The electrons are arranged in energy levels (or shells) from level 1, closest to the nucleus. Moving outwards from the nucleus, the shells gradually increase in energy.
- When filling up the energy levels in an atom, electrons go into the lowest energy level first.
- Most energy levels (except the first) contain sub-levels (or sub-shells) denoted by letters s, p and d.
- Different sub-levels contain different numbers of orbitals, with each orbital holding a maximum of two electrons.
- In sub-levels containing more than one orbital, each of the orbitals is populated singly before any are doubly filled.
- Figure 1.2 has one strange entry – the 4s-orbital has a lower energy than the 3d-orbital.

As the number of protons increases, the electron energy levels fill up in the following sequence: 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p… (Table 1.4). This sequence can be followed in the Periodic Table.

Table 1.4

Element	Electronic configuration	Element	Electronic configuration

Element	Electronic configuration	Element	Electronic configuration
Hydrogen	$1s^1$	Carbon	$1s^2, 2s^2, 2p^2$
Helium	$1s^2$	Nitrogen	$1s^2, 2s^2, 2p^3$
Lithium	$1s^2, 2s^1$	Oxygen	$1s^2, 2s^2, 2p^4$
Beryllium	$1s^2, 2s^2$	Fluorine	$1s^2, 2s^2, 2p^5$
Boron	$1s^2, 2s^2, 2p^1$	Neon	$1s^2, 2s^2, 2p^6$

How big is an atom?

In theory, at least, it should be easy to define the radius of an atom as the distance from the nucleus to the edge of the atom (the limit of the atom's electrons). Unfortunately, the term does not mean the same for every element.

Look at Figure 1.3. You can see three different types of radius. In (a) and (b) the atomic radius is half the internuclear distance, and this applies to metals and covalent molecules respectively. If we have a noble gas such as argon, it does not form molecules, so we use a value called the van der Waals' radius – the radius of an atom which is not chemically bound to another atom.

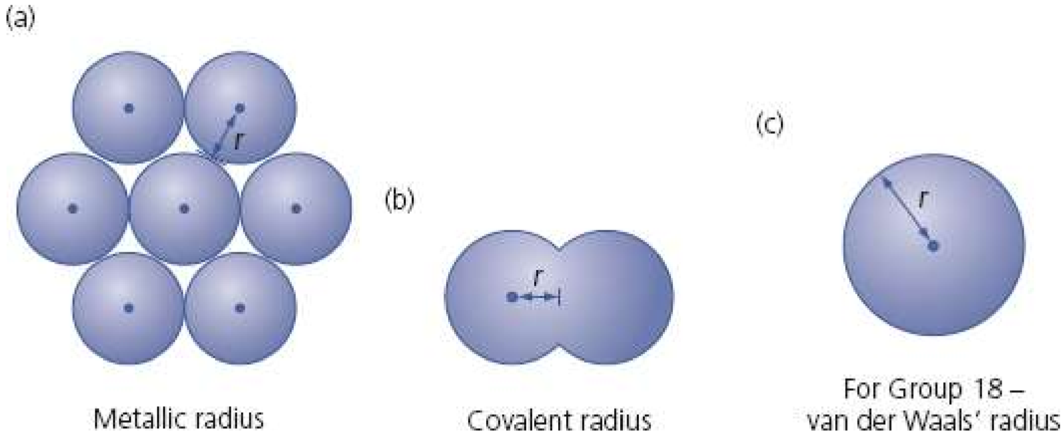

Figure 1.3 Different ways of measuring atomic radius

Figure 1.4 shows the atomic radii of Period 3 elements. As we go across a period the increasing charge on the nucleus means the electrons are more firmly held and the atomic radius decreases. Since argon does not bond to other atoms we have to take a radius based on half the internuclear distance for argon in the solid state.

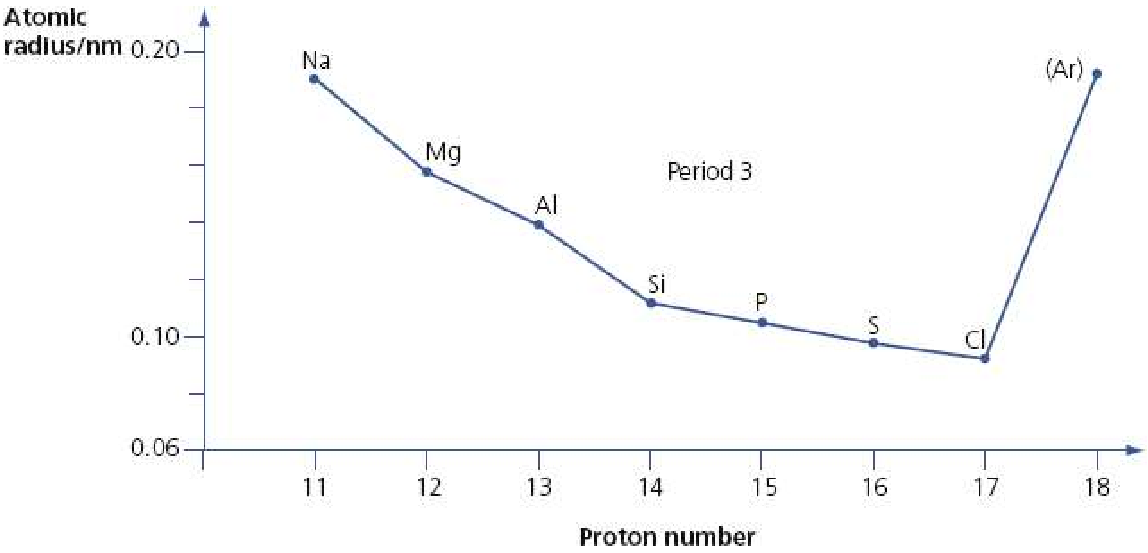

Figure 1.4 Atomic radii across Period 3

In an examination, you may be asked to deduce the electronic configuration of an atom (or ion) given its proton number (and any charge). The next example shows how to do this.

WORKED EXAMPLE

Element X, with proton number 16, forms an ion, X^{2-}. What is the electronic configuration of the ion?

Answer

The ion contains an extra two electrons compared with the atom. This means that it contains a total of (16 + 2) or 18 electrons. Looking at Figure 1.2 you can see that these extra two electrons will fit into the remainder of the 3p-orbital, giving an electronic configuration of $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$ for the ion X^{2-}.

The different orbitals have different shapes. Cross-sections of these orbitals are shown in Figure 1.5.

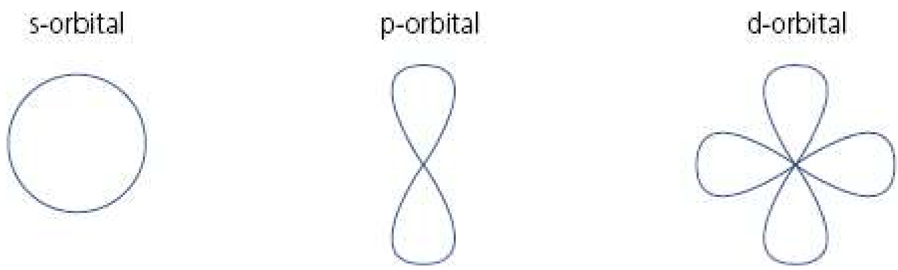

Figure 1.5 Cross-sections of s-, p- and d-orbitals

The location of electrons in the different types of orbital can affect the shapes of molecules.

NOW TEST YOURSELF

2 A particle can be written as $^{25}_{12}Q^{2+}$.

a Write down the numbers of protons, neutrons and electrons in the particle.

b Write the electronic configuration of the particle.

c Find element Q in the Periodic Table. What do you notice is unusual about your answer to part a?

Ionisation energies

The **first ionisation energy** of an atom has a precise definition that you need to remember.

KEY TERM

A **first ionisation energy** is the energy required to convert 1 mole of gaseous atoms of an element into 1 mole of gaseous cations, with each atom losing one electron.

This can be represented as:

$$M(g) \rightarrow M^+(g) + e^-$$

As you might expect, there are changes in the first ionisation energy values as the number of protons in the nucleus increases. This leads to a '2-3-3' pattern for Periods 2 and 3, as shown in Figure 1.6.

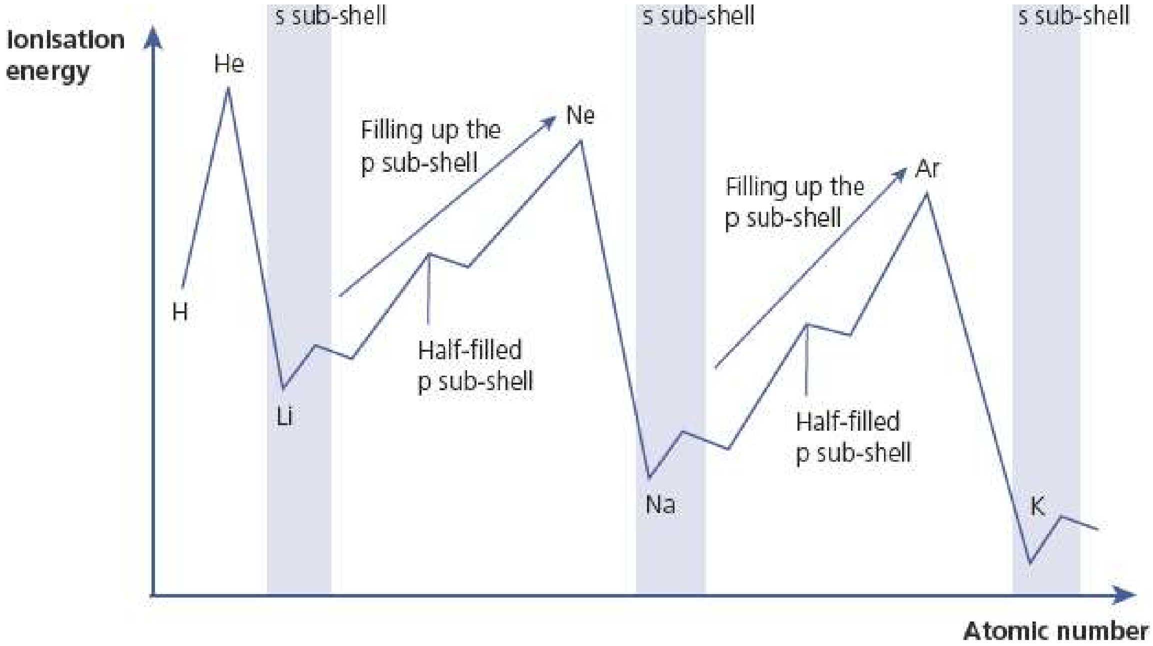

Figure 1.6 Relationship between first ionisation energy and atomic number

The graphs shown in some textbooks look complicated. For the examination you need to know the principles of the change. In an examination you might be asked to explain:

- the general increase in first ionisation energy across a period – proton number/nuclear charge increases across the period; shielding by other electrons is similar; so there is a bigger attraction for the electrons
- the fall between Groups 2 and 3 and/or Groups 15 and 16 – for Groups 2 and 3 the electron is removed from a p-orbital; which is higher in energy than an electron in an s-orbital; so it is easier to remove. For Groups 15 and 16 there are paired electrons in one of the p-orbitals; this causes repulsion; which makes it easier to remove one of these electrons
- the big fall at the end of the period – an electron shell has been completed; which results in more shielding; so there is less attraction for the outer electrons.

Successive ionisation energies

Successive ionisation energies involve the removal of second and subsequent electrons, for example:

- second ionisation energy:

$$M^{+}(g) \rightarrow M^{2+}(g) + e^{-}$$

- third ionisation energy:

$$M^{2+}(g) \rightarrow M^{3+}(g) + e^{-}$$

Knowledge of successive ionisation energies for an unknown element enables us to deduce which group the element is in. You know that successive ionisation energies increase as outer electrons are removed, and that a big increase occurs when an electron is removed from a new inner orbital closer to the nucleus.

NOW TEST YOURSELF

3 The graph in Figure 1.7 shows successive ionisation energies for an element Z. In which group of the Periodic Table is Z? Explain your answer.

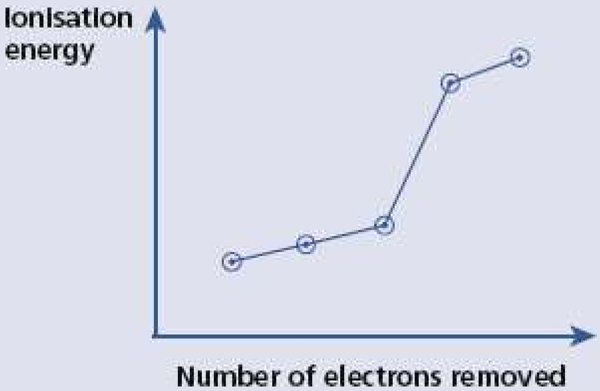

Figure 1.7 Graph of successive ionisation energies for element Z

REVISION ACTIVITY

a A particle can be described by the symbol $^{40}_{19}X^{+}$. How many protons, neutrons and electrons are in the particle?

b Compare the composition of particles of element X to those of the element with atomic number 19 in the Periodic Table. What is unusual about them?

c What do we call the energy change for the process shown here?

$$^{40}_{19}X(g) \rightarrow ^{40}_{19}X^{+}(g) + e^{-}$$

d Would you expect the energy change for the process described below to be smaller, the same or larger than that in part c? Explain your answer.

$$^{40}_{19}X^{+}(g) \rightarrow ^{40}_{19}X^{2+}(g) + e^{-}$$

END OF CHAPTER CHECK

By now you should be able to:

- describe the structure and size of atoms
- identify isotopes of elements and discuss their properties
- describe the number of electrons in atoms, their energy levels and atomic orbitals
- identify and explain trends in ionisation energies of atoms and their relationship with the Periodic Table