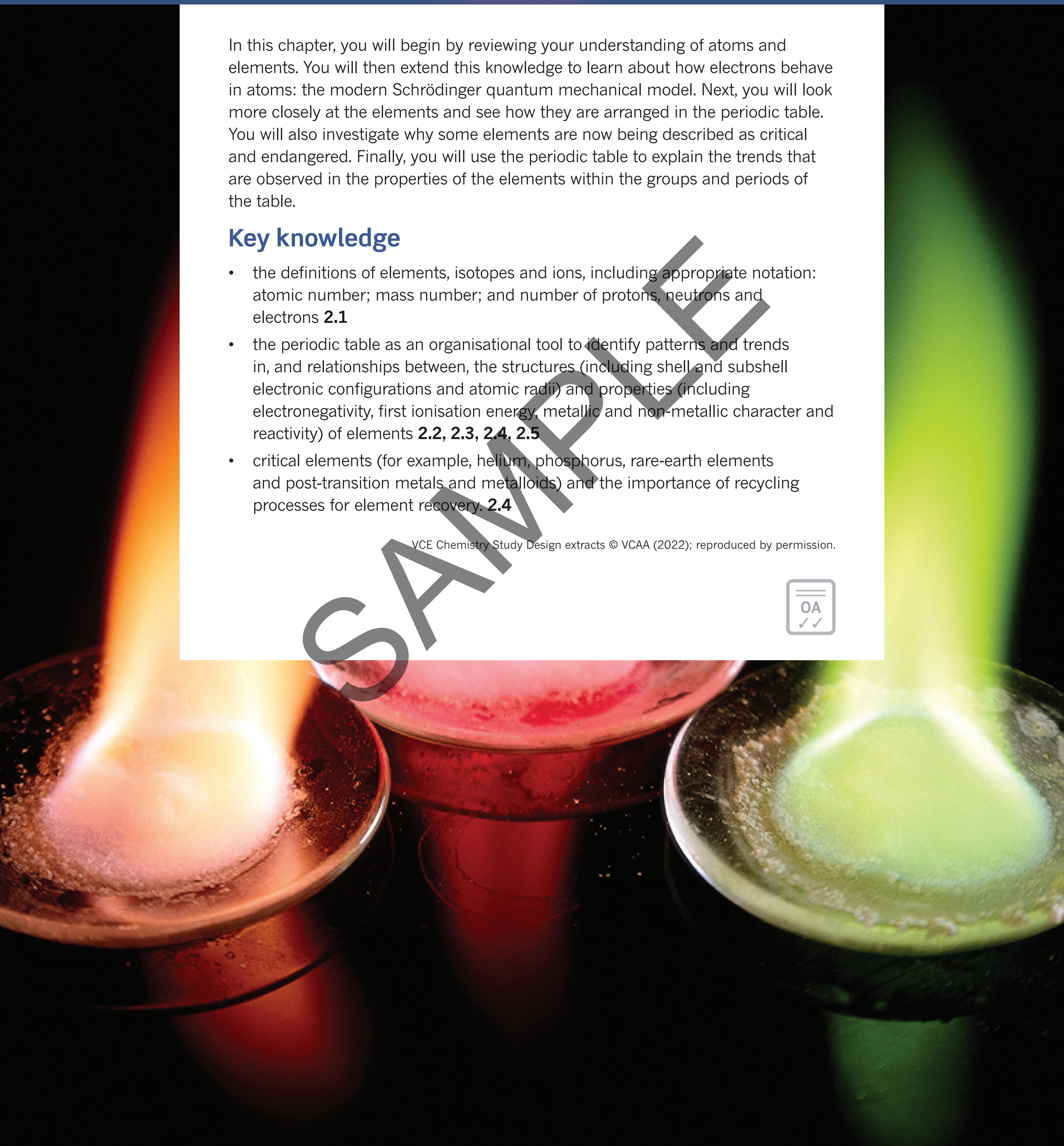
CHAPTER

02 Elements and the periodic table

In this chapter, you will begin by reviewing your understanding of atoms and elements. You will then extend this knowledge to learn about how electrons behave in atoms: the modern Schrödinger quantum mechanical model. Next, you will look more closely at the elements and see how they are arranged in the periodic table. You will also investigate why some elements are now being described as critical and endangered. Finally, you will use the periodic table to explain the trends that are observed in the properties of the elements within the groups and periods of the table.

Key knowledge

- the definitions of elements, isotopes and ions, including appropriate notation: atomic number; mass number; and number of protons, neutrons and electrons **2.1**
- the periodic table as an organisational tool to identify patterns and trends in, and relationships between, the structures (including shell and subshell electronic configurations and atomic radii) and properties (including electronegativity, first ionisation energy, metallic and non-metallic character and reactivity) of elements **2.2, 2.3, 2.4, 2.5**
- critical elements (for example, helium, phosphorus, rare-earth elements and post-transition metals and metalloids) and the importance of recycling processes for element recovery. **2.4**



2.1 The atomic world

Over time, scientists have gained a deep understanding of the structure of **atoms**, which are the basic building blocks of **matter**. As atoms are too small to be seen with even the most powerful optical microscope, much of what scientists know about atoms has come from theoretical models and indirect observations.

A scientific **model** is a tool used by scientists to understand something they cannot see directly. Using their observations, they are able to construct a theoretical picture of what they are trying to describe. As new data becomes available, the model can develop and become more accurate.

A scientific model is a tool that may be used by scientists to explain something they cannot see directly.

ATOMIC THEORY

In 1802, an English scientist called John Dalton (Figure 2.1.1) presented the first **atomic theory of matter**. Dalton proposed that all matter is made up of tiny spherical particles, which are indivisible and indestructible.

Dalton also accurately described **elements** as materials containing just one type of atom and **compounds** as materials containing different types of atoms in fixed ratios. Subsequent experiments showed that Dalton's atomic theory of matter was mostly correct. However, scientists now know that atoms are not indivisible or indestructible. Atoms are made up of even smaller **subatomic particles**.

Elements are materials that contain just one type of atom. Compounds are materials that contain different types of atoms, in fixed ratios.

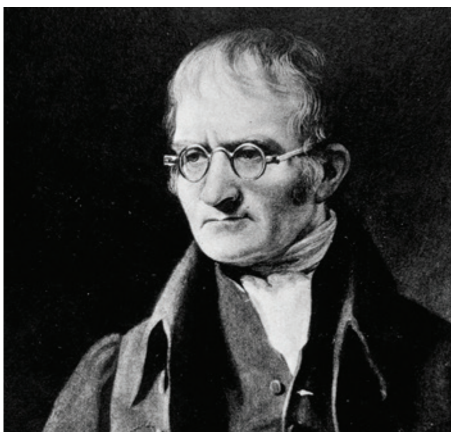


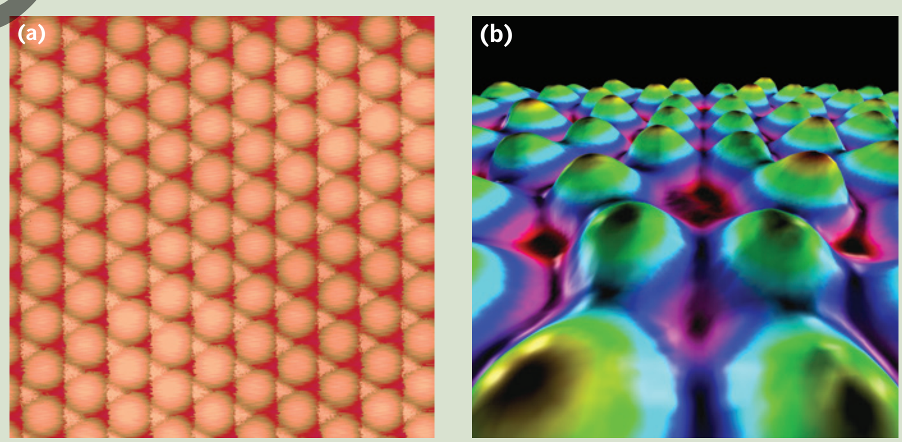
FIGURE 2.1.1 John Dalton (1766–1844) proposed that matter was composed of atoms.

CHEMFILE

Viewing atoms

Dalton's atomic theory of matter assumed that atoms are spherical. However, atoms cannot be seen with conventional microscopes. Therefore, there was no way to confirm the shape of atoms until 1981 when a microscope capable of viewing atoms was developed by IBM researchers Gerd Binnig and Heinrich Rohrer. This type of microscope is known as a scanning tunnelling microscope (STM). Using STMs, scientists confirmed that atoms are indeed spherical.

STMs use an extremely sharp metal tip to detect atoms. The tip is scanned, line-by-line, across the surface of a crystal. As it moves, the tip measures minute height differences in the crystal's surface due to the individual atoms. This is similar to the way a person with a vision impairment uses their finger to sense braille on a page. The data from the tip is then sent to a computer that constructs an image of the atoms. STM images of a lattice of copper atoms and silicon atoms on the surface of a silicon chip are shown in the figures below.



(a) Coloured image of copper atoms produced by a scanning tunnelling microscope (STM).
(b) Silicon atoms on a silicon chip imaged with a scanning tunnelling microscope

STRUCTURE OF ATOMS

Atoms are made up of a small, positively charged nucleus surrounded by a much larger cloud of negatively charged electrons as shown in Figure 2.1.2. The nucleus is in turn made up of two types of subatomic particle—protons and neutrons. The **protons** are positively charged and the **neutrons** have no charge.

Electrons

Electrons are negatively charged particles. You can imagine them forming a cloud of negative charge around the nucleus. This cloud gives the atom its size and volume.

An electron is approximately 1800 times smaller than a proton or neutron. Therefore, electrons contribute very little to the total mass of an atom. However, the space occupied by the cloud of electrons is 10 000–100 000 times larger than the nucleus.

Negative particles attract positive particles. This is called **electrostatic attraction**. Electrons are bound to the nucleus by the electrostatic attraction to the protons within the nucleus. The charge on an electron is equal but opposite to the charge on a proton. Electrons are said to have a charge of –1 whereas protons have a charge of +1.

In some circumstances, electrons can be easily removed from an atom. For example, when you rub a rubber balloon on a woollen jumper or dry hair, electrons are transferred to the balloon and the balloon develops a negative charge. The negative charge is observed as an electrostatic force that can attract hair or even stick the balloon to a wall (Figure 2.1.3). You will look at a different way of removing electrons from atoms, via a chemical reaction, when looking at redox reactions in Chapter 12.

The electricity that powers lights and appliances is the result of electrons moving in a current through wires. Sparks and lightning are also caused by electrons moving through air.

The nucleus

The **nucleus** of an atom is approximately 10 000–100 000 times smaller than the size of the atom. To put this in perspective, if an atom were the size of the Melbourne Cricket Ground (Figure 2.1.4), then the nucleus would be about the size of a pea in the centre. Nonetheless, the nucleus contributes around 99.97% of the atom's mass. This means that atomic nuclei are extremely dense.

The subatomic particles in the nucleus, the protons and neutrons, are referred to collectively as **nucleons**. Protons are positively charged particles with a mass of approximately 1.673×10^{-27} kg. Neutrons are almost identical in mass to protons.

FIGURE 2.1.4 If an atom were the size of the Melbourne Cricket Ground, then the nucleus would be the size of a pea.

Atoms are made up of protons (positive charge), neutrons (neutral) and electrons (negative charge).

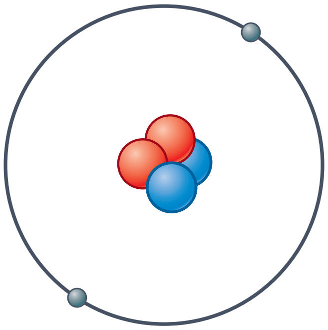


FIGURE 2.1.2 A simplified model of the atom. This shows a helium atom with two protons and two neutrons (in the central nucleus) and two electrons. The ring symbolises that the electrons are held in an orbit; however, evidence suggests they move throughout an area that is more like a cloud.

An electron is approximately 1800 times smaller than a proton or a neutron. However, electrons occupy most of the space in an atom.

The charge on an electron is equal but opposite to the charge on a proton.

FIGURE 2.1.3 It is the build-up of electrons, as static charge on the balloon, that attracts this girl's hair to the surface of the balloon.

Most of the mass of an atom is concentrated in the extremely dense nucleus.

Protons and neutrons are almost identical in mass.

Table 2.1.1 summaries the properties of protons, neutrons and electrons.

TABLE 2.1.1 Properties of the subatomic particles

Particle	Symbol	Charge	Mass relative to a proton	Mass (kg)
proton	p	+1	1	1.673×10^{-27}
neutron	n	0	1	1.675×10^{-27}
electron	e	–1	$\frac{1}{1800}$	9.109×10^{-31}

ELEMENT SYMBOLS

Atoms can be identified by how many protons they have. An element is made up of atoms that all contain the same number of protons in their nucleus. Scientists have discovered 118 different elements and about 98 of these occur in nature. The other elements have only been observed in the laboratory.

Each element has a name and a unique **chemical symbol**. Table 2.1.2 lists the chemical symbols of some well-known elements.

TABLE 2.1.2 Chemical symbols and names of some well-known elements

Element	Symbol	Element	Symbol
aluminium	Al	mercury	Hg
argon	Ar	nitrogen	N
carbon	C	oxygen	O
chlorine	Cl	potassium	K
copper	Cu	silver	Ag
hydrogen	H	sodium	Na
iron	Fe	uranium	U

The chemical symbol is made up of one or two letters. The first letter is always capitalised and subsequent letters are always lower case.

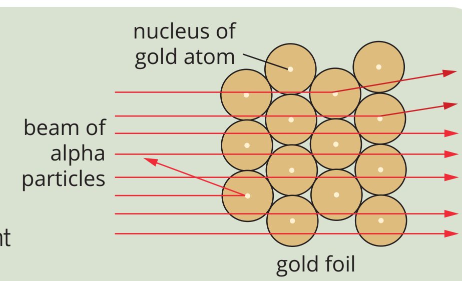
In many cases, the chemical symbol corresponds to the name of the element. For example, nitrogen has the chemical symbol N, chlorine has the chemical symbol Cl and uranium has the chemical symbol U.

However, some chemical symbols do not seem to correspond to the name of the element. For example, sodium has the chemical symbol Na, potassium has the chemical symbol K and iron has the chemical symbol Fe. This is because the chemical symbols have been derived from the Latin or Greek names of the elements. In Latin, sodium is known as *natrium*, potassium is known as *kalium* and iron is known as *ferrum*.

CHEMFILE

Atoms are mostly empty space

Between 1899 and 1911, New Zealand–born physicist Ernest Rutherford conducted experiments in which he fired a beam of alpha particles (helium nuclei) at a piece of extremely thin gold foil. You can imagine this experiment as a little like throwing pebbles at an object in the dark to deduce its shape. By listening to whether the pebbles hit something or pass straight through, you can build up a picture of the object in front of you. Rutherford noted that most of the alpha particles passed straight through the gold foil (see figure below). Even more surprisingly, Rutherford found that a small number of particles bounced back, some almost directly back at the source. From this observation, he deduced that the gold atoms were made up almost entirely of empty space, with a small, extremely dense nucleus.



Rutherford's gold foil experiment: only those alpha particles that closely approach the nuclei in the gold foil are deflected significantly. Most particles pass directly through the foil.

Representing atoms

The number of protons in an atom's nucleus is known as the **atomic number** and is represented by the symbol Z.

All atoms that belong to the same element must have the same number of protons and therefore have the same atomic number, Z. For example, all hydrogen atoms have Z = 1, all carbon atoms have Z = 6 and all gold atoms have Z = 79.

The total number of protons and neutrons in the nucleus is known as the **mass number** and is represented by the symbol A. The mass number represents the total mass of the nucleus.

As all atoms are electrically neutral, the number of electrons in an atom is equal to the number of protons in an atom. The atomic number therefore tells you both the number of protons and the number of electrons. For example, carbon atoms, with Z = 6, have six protons and six electrons.

The number of protons, neutrons and electrons defines the basic structure of an atom. A standard way of representing an atom is to show its atomic and mass numbers as shown in Figure 2.1.5. This is known as **nuclide notation**.

For an aluminium atom, this would be written like this: $^{27}_{13}\mathrm{Al}$

From this representation, you can determine that:

- the number of protons is 13 because the number of protons is equal to the atomic number (Z)
- the number of neutrons is 14 because the number of neutrons plus the number of protons is equal to the mass number; therefore, you can subtract the atomic number from the mass number to determine the number of neutrons (A – Z)
- the number of electrons is 13 because atoms have no overall charge, therefore the number of electrons must equal the number of protons.

Worked example 2.1.1

CALCULATING THE NUMBER OF SUBATOMIC PARTICLES

Calculate the number of protons, neutrons and electrons for the atom with this nuclide symbol: $^{40}_{18}\mathrm{Ar}$

Thinking	Working
The atomic number is equal to the number of protons.	The number of protons = Z = 18
Find the number of neutrons. Number of neutrons = mass number – atomic number	The number of neutrons = A – Z = 40 – 18 = 22
Find the number of electrons. The number of electrons is equal to the atomic number because the total negative charge is equal to the total positive charge.	Number of electrons = Z = 18

Worked example: Try yourself 2.1.1

CALCULATING THE NUMBER OF SUBATOMIC PARTICLES

Calculate the number of protons, neutrons and electrons for the atom with this nuclide symbol: $^{235}_{92}\mathrm{U}$

The atomic number is equal to the number of protons in the nucleus, and is represented by the symbol Z.

The mass number is the sum of all the particles in the nucleus, i.e. protons plus neutrons. This represents the total mass of the atom.

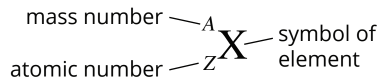


FIGURE 2.1.5 The standard way of representing an atom, showing its atomic number and mass number

CHEMFILE

Using isotopes to study climate change

One of the ways scientists study climate change is to study frozen bubbles of air, deep within ice that has remained frozen for thousands of years. The first ice-core studies were done using samples from Greenland and Antarctica, but more recently scientists have been searching for suitable ice deposits in the world's warmer regions like Africa and South America. One such expedition took scientists to the summit of Mount Kilimanjaro in Tanzania.

The oldest ice-cores collected in the expedition were 11 700 years old (see figure below). By determining the ratio of water in the ice containing the oxygen-18 isotope versus water containing the oxygen-16 isotope, it is possible to determine the temperature of the air when the water originally fell as rain. If there is a larger amount of oxygen-18, the temperature was higher; if there is a larger amount of oxygen-16, the temperature was lower.

A scientist begins their analysis of an ice-core sample.

Isotopes

Isotopes are atoms with the same atomic number, but with different mass numbers. They contain the same number of protons, but the number of neutrons in the nucleus is different.

All atoms that belong to the same element have the same number of protons in the nucleus and therefore the same atomic number, Z. However, not all atoms that belong to the same element have the same mass number, A. For example, hydrogen atoms can have a mass number of 1, 2 or 3. In other words, hydrogen atoms may contain just a single proton, a proton and a neutron, or a proton and two neutrons as shown in Figure 2.1.6. Atoms that have the same number of protons (atomic number) but different numbers of neutrons (and therefore different mass numbers) are known as **isotopes**.

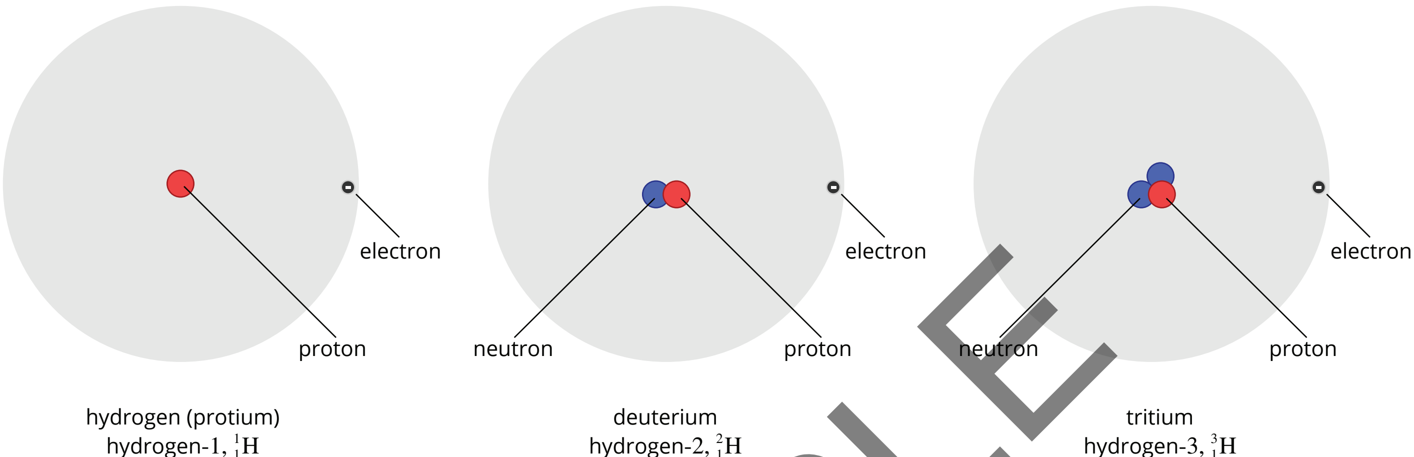


FIGURE 2.1.6 The three isotopes of hydrogen are given special names. A hydrogen atom with just one proton in its nucleus is known as hydrogen or protium. A hydrogen atom with one proton and one neutron is known as deuterium. A hydrogen atom with one proton and two neutrons is known as tritium.

Carbon also has three naturally occurring isotopes. These three isotopes are known as carbon-12, carbon-13 and carbon-14. Carbon-12 atoms have a mass number of 12, carbon-13 atoms have a mass number of 13 and carbon-14 atoms have a mass number of 14. In the 1950s and 1960s, nuclear weapons testing caused a spike in carbon-14 in the atmosphere. This has been declining in the last 50 years. These three carbon isotopes can be represented as shown in Figure 2.1.7.

$^{12}_{6}C$ carbon–12 $\quad$ $^{13}_{6}C$ carbon–13 $\quad$ $^{14}_{6}C$ carbon–14

FIGURE 2.1.7 Ways of representing the three isotopes of carbon

Isotopes have identical chemical properties but different physical properties such as mass and density. Some isotopes are **radioactive**. This means their nucleus is not stable and will break down spontaneously into a more stable form by emitting particles as radiation.

Ions

Nuclide symbols can also be used to represent ions. **Ions** are atoms that have lost or gained one or more electrons. An atom that loses electrons becomes positively charged overall (as the positive charges in the nucleus now outnumber the negatively charged electrons.) Similarly, an atom that gains electrons becomes negatively charged. You will learn more about ions in Chapter 5.

When an atom gains or loses one or more electrons, it becomes an ion. An ion is a charged particle. Atoms that lose electrons become positive ions. Atoms that gain electrons become negative ions.

Worked example 2.1.2

CALCULATING THE NUMBER OF SUBATOMIC PARTICLES IN AN ION

Calculate the number of protons, neutrons and electrons for the ion with this nuclide symbol:

$${}^{25}_{12}Mg^{2+}$$

Thinking	Working
The atomic number is equal to the number of protons.	The number of protons = Z = 12
Find the number of neutrons. Number of neutrons = mass number – atomic number	The number of neutrons = A – Z = 25 – 12 = 13
Find the number of electrons in an uncharged atom. The number of electrons is equal to the atomic number.	Number of electrons in an uncharged atom = Z = 12
Find the number of electrons in the ion. The number of electrons is equal to the atomic number minus two, because the total negative charge is two less than the total positive charge.	Number of electrons in ion = Z – (charge) = Z – 2 = 12 – 2 = 10

Worked example: Try yourself 2.1.2

CALCULATING THE NUMBER OF SUBATOMIC PARTICLES IN AN ION

Calculate the number of protons, neutrons and electrons for the ion with this nuclide symbol:

$${}^{14}_{7}N^{3-}$$

2.1 Review

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SUMMARY

- All matter is made of atoms, which are composed of a small, positively charged nucleus surrounded by a negatively charged cloud of electrons.
- The mass of an atom is mostly determined by the mass of the nucleus, while the diameter of an atom is determined by the cloud of electrons.
- The nucleus is made up of two subatomic particles—protons and neutrons. These particles are referred to as nucleons.
- Protons have a positive charge, electrons have a negative charge and neutrons have no charge.
- Protons and neutrons are similar in mass while electrons are approximately 1800 times smaller.
- The charges on protons and electrons are equal but opposite.
- An element contains atoms of the same type and has a chemical symbol that is made up of one or two letters. The first letter is capitalised and the second letter is lower case.
- You can determine the number of subatomic particles in an atom from an element's atomic number and mass number:

$$\text{mass number} \longrightarrow A \qquad {}^{A}_{Z}\mathrm{X} \longleftarrow \text{symbol of element} \qquad \text{atomic number} \longrightarrow Z$$

  - Z = number of protons = number of electrons
  - A – Z = number of neutrons
- Isotopes are atoms with the same atomic number but different mass numbers, i.e. they have the same number of protons but different numbers of neutrons.
- Isotopes have the same chemical properties but different physical properties such as mass, density and radioactivity.
- Ions are atoms that have lost or gained electrons to become charged particles.

KEY QUESTIONS

Knowledge and understanding

1 How many times larger is the atom compared to its nucleus?

2 What subatomic particles make up most of the mass of an atom and where are they found?

3 How are electrons held within the cloud surrounding the nucleus?

4 What term is given to the number of protons and neutrons in the nucleus of an atom?

Analysis

5 How many electrons would an atom of ${}^{65}_{30}\mathrm{Zn}^{2+}$ contain?

6 Yttrium-90 (atomic number 39) is used for treatment of cancer, particularly non-Hodgkin's lymphoma and liver cancer, and it is being used more widely, including for arthritis treatment.

a Write the nuclide symbol for yttrium-90.

b How many neutrons does each atom of this isotope contain?

7 Which of the following nuclide symbols represent isotopes of the same element?

a ${}^{14}_{6}\mathrm{X}$

b ${}^{14}_{7}\mathrm{Y}$

c ${}^{12}_{6}\mathrm{Z}$

8 Suggest why it might be easier to separate ${}^{46}_{20}\mathrm{Ca}$ and ${}^{46}_{22}\mathrm{Ti}$ than ${}^{46}_{20}\mathrm{Ca}$ and ${}^{40}_{20}\mathrm{Ca}$.

2.2 Emission spectra and the Bohr model

When fireworks explode, they create a spectacular show of coloured light (Figure 2.2.1). The light is produced by metal atoms that have been heated by the explosion. This coloured light posed a significant problem for early scientists. The models the scientists were using could not explain the source of the light. However, the light was a clue that ultimately led to a better understanding of the arrangement of electrons in atoms.

FIGURE 2.2.1 The spectacular colours in this New Year's Eve fireworks display are emitted by metal atoms that have been heated to very high temperatures.

EMISSION SPECTRA

When atoms are heated, they can produce coloured light. You may have observed this phenomenon in a flame test. **Flame tests** make use of the fact that some metallic elements can be identified by the characteristic colour produced when a sample is passed through a flame. Figure 2.2.2 shows the characteristic colours produced by some metals.

When atoms are heated, they can produce coloured light. Observations of this phenomenon led to a greater understanding of the behaviour of electrons.

FIGURE 2.2.2 The colour of the flame is determined by the different metal compounds present and can be used to identify these metals. The flame colours shown here are for (from left to right): barium (yellow–green), lithium (crimson), strontium (scarlet), sodium (yellow), copper (green) and potassium (lilac).

If the light from the flame test is passed through a prism, it produces a spectrum with a black background and a number of coloured lines. Figure 2.2.3 shows the apparatus used to produce these spectra.

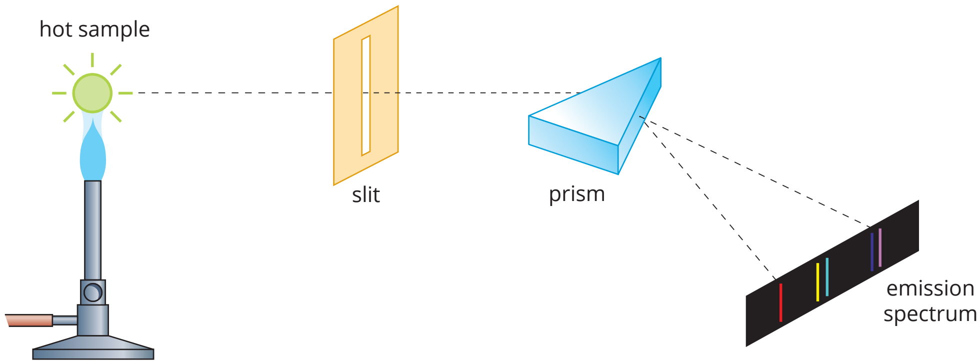

FIGURE 2.2.3 The apparatus used to analyse the coloured light given out when an element is heated. The coloured lines are called an emission spectrum.

These spectra are known as line spectra or **emission spectra**. Each emission spectrum is unique for a particular element and can be used to identify different elements.

The line spectrum produced by helium is shown in Figure 2.2.4.

FIGURE 2.2.4 The emission spectrum of helium is made up of lines ranging from violet to red in colour.

Each line in the spectrum corresponds to light of a different energy. Violet lines correspond to light with high energies. As the colour of the light changes to blue, green, yellow and orange, the energy of the light decreases. Red light is the lowest energy light visible to the human eye. Just as some metals have a characteristic flame colour, elements have a characteristic emission spectrum.

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THE BOHR MODEL

In 1913, Niels Bohr developed a new model of the hydrogen atom that explained its emission spectrum. The **Bohr model** proposed the following:

- Electrons revolve around the nucleus in fixed, circular orbits.
- These orbits correspond to specific energy levels in the atom.
- Electrons can only occupy fixed energy levels and cannot exist between two energy levels.
- Orbits of larger radii correspond to energy levels of higher energy.

In the Bohr model, it is possible for electrons to move between the energy levels by absorbing or emitting energy. Bohr's model (Figure 2.2.5) gave close agreement between the calculated energies for lines in the hydrogen spectrum and the observed values in the spectrum.

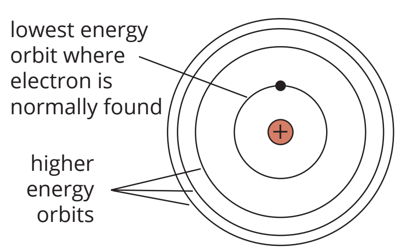


FIGURE 2.2.5 The Bohr model of a hydrogen atom. Bohr suggested that electrons moved in orbits of particular energies.

ELECTRON SHELLS

Scientists quickly extended Bohr's model of the hydrogen atom to other atoms. Scientists proposed that electrons were grouped in different energy levels, called **electron shells**. The electron shells are labelled with the number $n = 1, 2, 3 \ldots$, as shown in Figure 2.2.6.

The orbit in which an electron moved depended on the energy of the electron; electrons with lower energy are in orbits close to the nucleus while high-energy electrons are in outer orbits.

The Bohr model of the atom proposes that electrons are grouped into different energy levels, called shells.

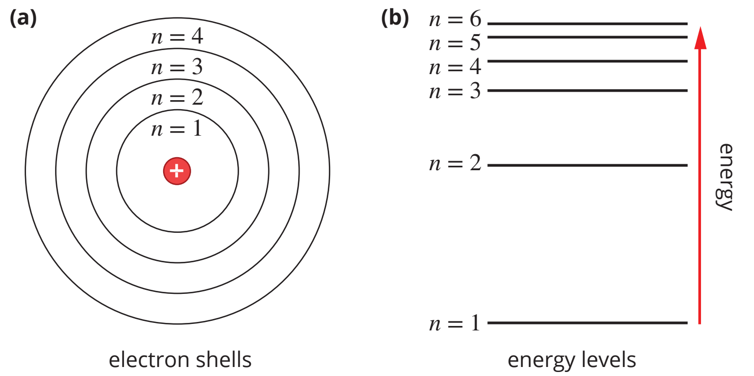


FIGURE 2.2.6 (a) The electron shells of an atom (n) are labelled using integers. The first shell is the shell closest to the nucleus and the radius of each shell increases as the shell number increases. (b) Each shell corresponds to an energy level that electrons can occupy. The first shell has the lowest energy and the energy increases as the shell number increases.

The lowest energy shell is the shell closest to the nucleus and is labelled $n = 1$. Shells with higher values of n correspond to higher energy levels. As the values of n increase, the energy levels get closer together.

LINKING EMISSION SPECTRA TO THE SHELL MODEL

Heating an element can cause an electron to absorb energy and jump to a higher **energy state**. Shortly afterwards, the electron returns to the lower energy state, releasing a fixed amount of energy as light. The electron can return in a number of different ways, some of which are shown in Figure 2.2.7. Each one of the different pathways produces light of a particular colour in the emission spectrum.

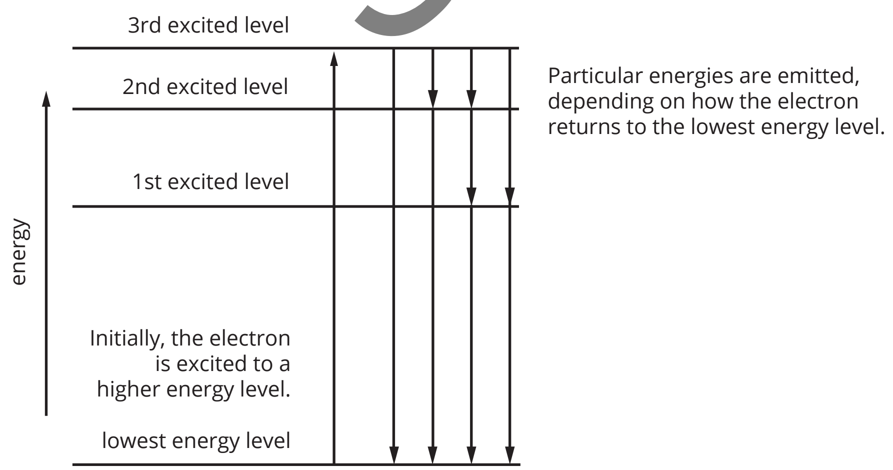


FIGURE 2.2.7 Emission of energy as electrons move from a higher to a lower energy state

Consider a hydrogen atom with one proton and one electron. Usually, the electron exists in the $n = 1$ shell. This is the lowest energy state of the atom and is called the **ground state**.

When the hydrogen atom is heated, the electron absorbs energy and jumps to a higher energy level. This is known as an **excited state**.

Shortly afterwards, the electron returns to the ground state. The electron may return directly to the ground state or may move to other energy levels before returning to the ground state. For example, an electron in the $n = 4$ shell may move to the $n = 2$ shell before returning to the $n = 1$ ground state.

As an electron falls to a lower energy shell, it emits energy, some of which is in the form of coloured light.

As the electron falls to a lower energy shell, it emits energy, some of which is in the form of coloured light. This energy is exactly equal to the energy difference between the two energy levels. Each coloured line on the spectrum corresponds to a specific amount of energy (Figure 2.2.8).

FIGURE 2.2.8 The emission spectrum of the hydrogen atom has four lines in the visible range—violet, blue, green and red.

ELECTRONIC CONFIGURATION IN SHELLS

The electronic configuration of an atom is a representation of the arrangement of electrons in energy levels.

The different energy levels or shells can hold different numbers of electrons. The arrangement of these electrons around the nucleus is called the **electronic configuration**. Electron shells can hold a maximum number of electrons, as shown in Table 2.2.1. You can calculate the maximum number of electrons a shell can hold using the formula $2n^2$, where n = shell number. Even though from the third shell on, shells can hold more than eight electrons, a **valence shell** (outer shell) can only hold eight electrons. Once a valence shell reaches eight electrons, the next shell starts filling.

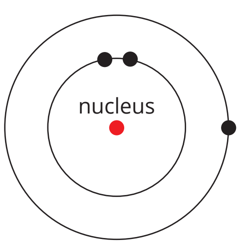
TABLE 2.2.1 The maximum number of electrons held by each shell of an atom. Shell 1 is the shell closest to the nucleus.

Electron shell (n)	Maximum number of electrons
1	2
2	8
3	18
4	32
n	$2n^2$

Electron shells fill in a particular order. The first two electrons go into the first shell. The next eight electrons go into the second shell. The third shell can hold 18 electrons, but once it contains eight electrons, the next two electrons go into the fourth shell. Only then does the third shell fill up. It took the development of Schrödinger's quantum mechanical model to explain these complexities of electronic configurations, as you will read in the next section.

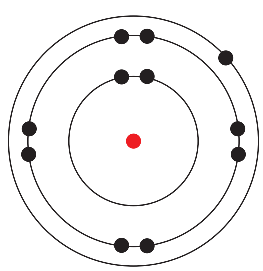
A Bohr diagram is a simple diagram that shows the arrangement of electrons around the nucleus. In such diagrams, only the shells that are occupied are drawn. The Bohr diagram and electronic configuration for lithium are shown in Figure 2.2.9. The electronic configuration can be shown as a series of numbers. For example, the electronic configuration of a magnesium atom with two electrons in the first shell, eight in the second and two in the third can be written as 2,8,2.

Figure 2.2.10 shows a Bohr diagram and electronic configuration for an atom with three electron shells, sodium. The Bohr diagram and electronic configuration for an atom with four shells is shown in Figure 2.2.11.



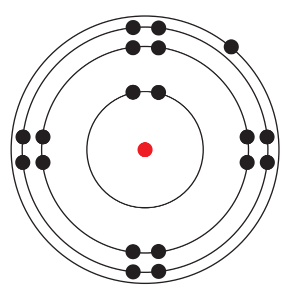
Li 2,1

FIGURE 2.2.9 The Bohr diagram and electronic configuration for the lithium atom with three electrons



Na 2,8,1

FIGURE 2.2.10 The electronic configuration of a sodium atom



K 2,8,8,1

FIGURE 2.2.11 The electronic configuration of potassium

Worked example 2.2.1

ELECTRONIC CONFIGURATION FOR UP TO 36 ELECTRONS

Apply the order of filling of the shell model to determine the electronic configuration of an atom with 28 electrons.

Thinking	Working
Recall the maximum number of electrons that each shell can hold.	Shell (*n*) Maximum number of electrons 1 2 2 8 3 18 4 32
Place the first 18 electrons in the shells from the lowest energy to the highest energy. Do not exceed the maximum number of electrons allowed.	Shell (*n*) Electrons in atom 1 2 2 8 3 8 4
Place the next two electrons in the fourth shell.	Shell (*n*) Electrons in atom 1 2 2 8 3 8 4 2
Continue filling the third shell until it holds up to 18 electrons. Put any remaining electrons in the fourth shell.	Shell (*n*) Electrons in atom 1 2 2 8 3 16 4 2 The 8 remaining electrons from the previous step have gone into the third shell.
Write the electronic configuration by listing the number of electrons in each shell separated by commas.	The electronic configuration is: 2,8,16,2

Worked example: Try yourself 2.2.1

ELECTRONIC CONFIGURATIONS FOR UP TO 36 ELECTRONS

Apply the order of filling of the shell model to determine the electronic configuration of an atom with 34 electrons.

2.2 Review

OA ✓✓

SUMMARY

- When atoms are heated, they can emit electromagnetic radiation in the form of coloured light. When the light is passed through a prism, it produces a spectrum made up of lines of different colours. These spectra are known as emission or line spectra.
- The Bohr model of the atom was the first atomic model to explain the origin of emission spectra and assumes that electrons can only exist in fixed, circular orbits of specific energies. These orbits later came to be known as energy levels or shells.
- When an electron absorbs energy it can jump to a higher energy shell and when an electron falls back to its original state it can emit energy in the form of coloured light. Each line in the emission spectrum corresponds to a specific electron transition between shells.
- Each shell can hold a maximum number of electrons. This can be calculated using the rule $2n^2$, where n = shell number.
- The valence shell cannot contain more than eight electrons.

KEY QUESTIONS

Knowledge and understanding

1 Explain how emission spectra provide evidence for electron shells in the Bohr model.

2 What form of energy is emitted when an electron moves from a higher energy shell to a lower energy shell?

3 Draw a Bohr diagram and write the electronic configuration for an atom of Argon. (Refer to the periodic table at the back of the book to find the atomic number of argon.)

Analysis

4 Determine the maximum number of electrons that the fifth shell can hold.

5 Compare an atom's electronic shell configuration to its position in the periodic table. Describe any patterns you observe.

6 An atom has an electronic configuration 2,6,8. Identify the atom and state why this electronic arrangement is unexpected. Propose a reason for this arrangement.

2.3 The Schrödinger model of the atom

The shell model of the atom that developed from the work of Niels Bohr mathematically explained the lines in the emission spectrum of hydrogen atoms. However, there were some things that the model could not explain. The shell model:

- cannot accurately predict the emission spectra of atoms with more than one electron
- is unable to explain why electron shells can only hold $2n^2$ electrons
- does not explain why the fourth shell accepts two electrons before the third shell is completely filled.

These limitations of the shell model indicate that the model is incomplete. Obtaining a better model of the atom required scientists to think about electrons in an entirely different way.

A QUANTUM MECHANICAL VIEW OF ATOMS

The Bohr model of the atom was revolutionary when it was proposed. Before Bohr, scientists believed that electrons could orbit the nucleus at any distance. This picture of the electrons was based on how scientists observed the world around them. For example, planets can revolve at any distance around the Sun. However, Bohr's theory stated that electrons only occupy specific, circular orbits. This was the first suggestion that the physics inside atoms might be very different to the physics we experience in our daily lives.

Quantum mechanics

The word 'quantum' simply means a specific amount. In the Bohr model, the electrons can only have specific amounts of energy depending on which shell they are in. The energy of the electrons is said to be **quantised**.

In 1926, the German scientist Erwin Schrödinger proposed that electrons behave as waves around the nucleus. Using a mathematical approach and this wave theory, Schrödinger developed a model of the atom called **quantum mechanics** (see Figure 2.3.1). The Schrödinger model of the atom is the model that scientists use today.

Quantum mechanics describes the behaviour of extremely small particles like electrons. You rarely experience quantum mechanics in your everyday life. As a result, the predictions of quantum mechanics are often difficult to visualise. Nonetheless, quantum mechanics accurately predicts the behaviour of electrons in atoms.

THE SCHRÖDINGER MODEL

The fundamental difference between the Bohr model and the Schrödinger model of the atom is the way they view the electrons. Bohr viewed electrons as tiny, solid particles that revolve around the nucleus in circular orbits. Schrödinger viewed electrons as having wave-like properties, just like light. In this model, the electrons occupy a three-dimensional space around the nucleus known as an **orbital**. Figure 2.3.2 on the following page compares the Bohr model and the Schrödinger model.

By assuming that electrons have wave-like properties, Schrödinger found the following.

- There are major energy levels in an atom that, for historical reasons, are called shells.
- These shells contain separate energy levels of similar energy, called **subshells**, which he labelled *s*, *p*, *d* and *f*. Each subshell can only hold a certain number of electrons.
- The first shell ($n = 1$) contains only an *s*-subshell. The second shell contains *s*- and *p*-subshells. The third shell contains *s*-, *p*- and *d*-subshells and so on. The subshells for the first four shells are summarised in Table 2.3.1 on the following page.

The model of the atom that is used today is called the quantum mechanical or Schrödinger model.

FIGURE 2.3.1 Edwin Schrödinger (1887–1961), the Austrian physicist whose research into subatomic particles is the basis of quantum mechanics.

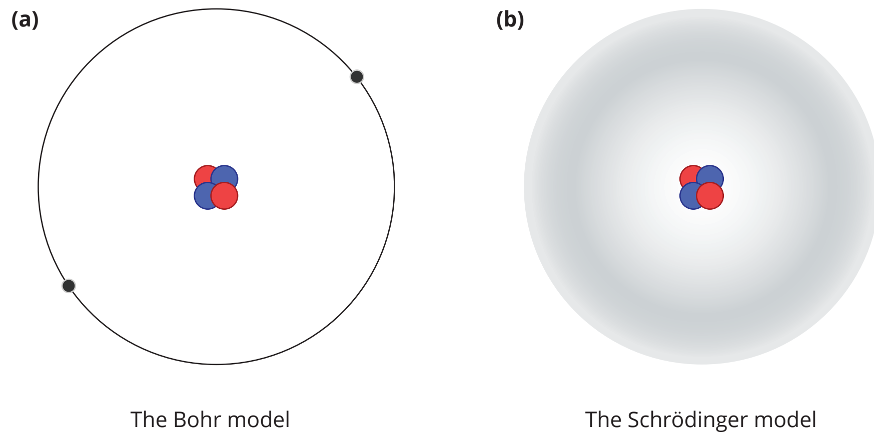


FIGURE 2.3.2 (a) Bohr regarded electrons as particles that travel along a defined path in circular orbits. (b) In Schrödinger's quantum mechanical approach, the electrons behave as waves and occupy a three-dimensional space around the nucleus. The region occupied by the electrons is known as an orbital.

The quantum mechanical model predicts subshells and orbitals within each of the shell levels.

- Each subshell is made up of smaller components known as orbitals. Orbitals can be described as regions of space surrounding the nucleus of an atom in which electrons may be found. Any orbital can hold a maximum of 2 electrons. An *s*-subshell has just one orbital. A *p*-subshell has three orbitals. A *d*-subshell has five orbitals and an *f*-subshell has seven. The number of orbitals in each subshell for the first four shells is summarised in Table 2.3.1.
- The total number of orbitals in a shell is given by n^2. Each orbital can contain a maximum of two electrons. Therefore, the total number of electrons per shell is given by $2n^2$. For example, the second shell contains *s*- and *p*-subshells and so contains a total of 1 + 3 = 4 orbitals. Each orbital can contain two electrons so the second shell contains up to 2 × 4 = 8 electrons. The number of electrons in each subshell and shell for the first four shells is summarised in Table 2.3.1.

TABLE 2.3.1 Energy levels within an atom

Shell number (n)	Subshells	Number of orbitals in subshell	Maximum number of electrons per subshell	Maximum number of electrons per shell
1	1*s*	1	2	2
2	2*s*	1	2	
	2*p*	3	6	8
3	3*s*	1	2	
	3*p*	3	6	
	3*d*	5	10	18
4	4*s*	1	2	
	4*p*	3	6	
	4*d*	5	10	
	4*f*	7	14	32

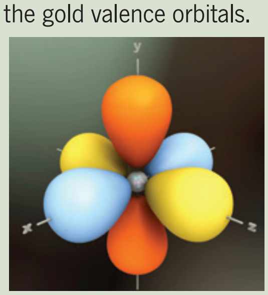
Note that the Schrödinger model accurately predicts the maximum number of electrons that each shell can hold. This is something that the Bohr model could not explain.

CHEMFILE

A model built on mathematical probability

In the Schrödinger model, electrons can be regarded as behaving like clouds of charge, where the cloud density at a particular location in the atom indicates the probability of finding an electron at that position. So, this model is based on probability, rather than certainty.

These mathematical regions of probability are called orbitals and are determined using mathematical wave theory. The shape of some of these orbitals may be surprising. The following model represents these orbitals as they would occur in a gold atom. You can see a model of the gold atom showing the Bohr energy levels, as well as models of the shapes of different types of orbitals and of the gold valence orbitals.



A computer-generated representation of *p*-orbitals around a nucleus. An orbital is a mathematical function.

ELECTRONIC CONFIGURATIONS AND THE SCHRÖDINGER MODEL

The electronic configurations for the Schrödinger model contain more detail than the electronic configurations of the shell model. This is because the Schrödinger model specifies subshells that electrons occupy. The subshell electronic configuration of a sodium atom is shown in Figure 2.3.3.

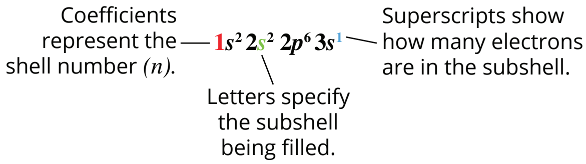


FIGURE 2.3.3 The subshell electronic configuration of a sodium atom. It shows that there are 2 electrons in the 1*s*-subshell, 2 electrons in the 2*s*-subshell, 6 electrons in the 2*p*-subshell and 1 electron in the 3*s*-subshell.

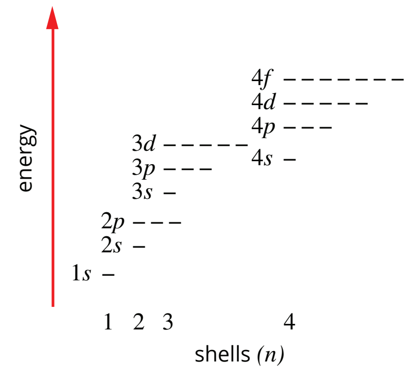


FIGURE 2.3.4 The relative energies of subshells. Each dashed line represents an orbital in a subshell. Each orbital can contain two electrons.

Sodium has 11 electrons. The subshell electronic configuration in Figure 2.3.3 indicates that sodium has:

- two electrons in the *s*-subshell in the first shell
- two electrons in the *s*-subshell of the second shell
- six electrons in the *p*-subshell of the second shell
- one electron in the *s*-subshell of the third shell.

Although the subshell electronic configurations may look complicated, the rules for constructing them are simple.

- The lowest energy subshells are always filled first.
- Each orbital contains a maximum of two electrons.

The order of energy levels of the subshells is:

$$1s < 2s < 2p < 3s < 3p < 4s < 3d \ldots$$

This is shown in Figure 2.3.4. In this diagram, each dash represents an orbital that can hold two electrons.

The geometric pattern shown in Figure 2.3.5 is a commonly used and convenient way of remembering the order in which the subshells are filled. Note that in this diagram the fourth shell starts filling before the third shell is completely filled. This is because the 4*s*-orbital is slightly lower in energy than the 3*d*-orbitals. As a result, the 4*s*-orbital accepts two electrons after the 3*s*- and 3*p*-orbitals are filled but before the 3*d*-orbital begins filling.

Figure 2.3.6 shows how the energy levels are filled in a neon atom, which has 10 electrons. The first two electrons fill the 1*s*-subshell, the second two electrons go into the next highest energy level, the 2*s*-subshell. The last six electrons then fill the next highest energy level, the 2*p*-subshell. The electronic configuration is written as $1s^2 2s^2 2p^6$.

Krypton has 36 electrons. According to the order of subshell filling, its electronic configuration is:

$$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$$

Although the 4*s*-subshell is filled before the 3*d*-subshell, the subshells of the third shell are usually grouped together when writing its electronic configuration. Therefore, the electronic configuration for a krypton atom is written as:

$$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6$$

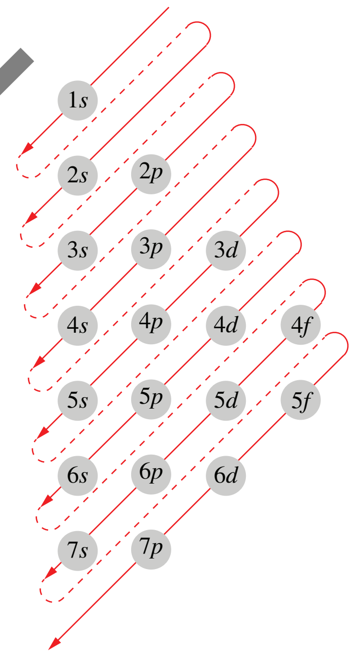


FIGURE 2.3.5 This geometric pattern shows the order in which the subshells are filled.

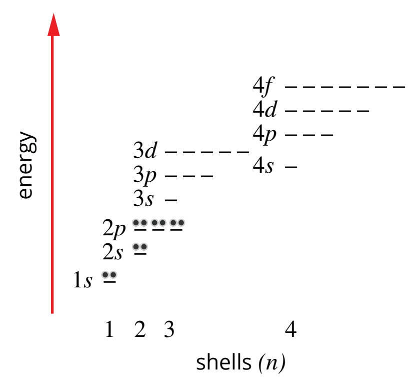


FIGURE 2.3.6 This electron subshell diagram shows how the orbitals in a neon atom are filled to give an electronic configuration.

Condensed electronic configuration

The electronic configuration of phosphorus is $1s^22s^22p^63s^23p^3$. This configuration can be written more simply as $[Ne]3s^23p^3$, where the [Ne] symbol signifies $1s^22s^22p^6$ because this is the electronic configuration of neon. A configuration like $[Ne]3s^23p^3$ is called a **condensed electronic configuration** or referred to as **noble gas notation**. When writing the condensed electronic configuration of an element, the symbol of the noble gas (with square brackets) in the period before the element is used. This means that the details for the inner shells are not written—they are implied by the use of the noble gas symbol. This is especially useful for elements with multiple shells containing electrons. Some other examples of condensed electronic configuration are shown in Table 2.3.2.

TABLE 2.3.2 Some examples of the use of condensed electronic configuration

Element	Atomic number	Electronic configuration (shell level)	Condensed electronic configuration (subshell level)
Si	14	2,8,4	$[Ne]3s^23p^2$
Zn	30	2,8,18,2	$[Ar]3d^{10}4s^2$
Br	35	2,8,18,7	$[Ar]3d^{10}4s^24p^5$

An advantage of the condensed notation is that it emphasises the valence shell electrons, which are fundamental for understanding how atoms interact. It also enables the detail of subshell notation to be displayed where there is limited space. Periodic tables, like the one on the back page of this book, often use these condensed electronic configurations.

Chromium and copper: exceptions

The electronic configurations for most elements follow the rules and pattern described above. There are two notable exceptions: element 24, chromium, and element 29, copper. Table 2.3.3 shows these exceptions.

TABLE 2.3.3 The electronic configurations for chromium and copper

Element	Electronic configuration predicted according to the pattern above	Actual electronic configuration
chromium, Cr	$1s^22s^22p^63s^23p^63d^44s^2$	$1s^22s^22p^63s^23p^63d^54s^1$
copper, Cu	$1s^22s^22p^63s^23p^63d^94s^2$	$1s^22s^22p^63s^23p^63d^{10}4s^1$

Each orbital can hold two electrons. In the *d*-subshell there are five orbitals and therefore 10 electrons that can be held within them. Scientists have determined that as a subshell fills, a single electron is placed in each orbital first. Then a second electron is entered into the orbitals until the filling process is complete.

Chemists calculate that there is very little difference in energy between 3*d*- and 4*s*-orbitals, and the $3d^54s^1$ configuration for chromium is slightly more stable than the $3d^44s^2$ configuration.

Similarly, for copper, the $3d^{10}4s^1$ arrangement with five completely filled *d*-orbitals is more stable than the $3d^94s^2$.

Worked example 2.3.1

WRITING ELECTRONIC CONFIGURATIONS USING THE SUBSHELL MODEL

Write the subshell electronic configuration for a manganese atom with 25 electrons.

<table>
<tr><th>Thinking</th><th>Working</th></tr>
<tr><td>Recall the order in which the subshells fill by listing them from lowest energy to highest energy and the number of orbitals in each.</td><td>1s, 1 orbital
2s, 1 orbital
2p, 3 orbitals
3s, 1 orbital
3p, 3 orbitals
4s, 1 orbital
3d, 5 orbitals</td></tr>
<tr><td>Fill the subshells by assigning two electrons per orbital, starting from the lowest energy subshells until you have reached the total number of electrons in your atom.</td><td><table>
<tr><th>Subshell</th><th>Electrons in subshell</th><th>Progressive total of electrons</th></tr>
<tr><td>1s</td><td>2</td><td>2</td></tr>
<tr><td>2s</td><td>2</td><td>4</td></tr>
<tr><td>2p</td><td>6</td><td>10</td></tr>
<tr><td>3s</td><td>2</td><td>12</td></tr>
<tr><td>3p</td><td>6</td><td>18</td></tr>
<tr><td>4s</td><td>2</td><td>20</td></tr>
<tr><td>3d</td><td>5</td><td>25</td></tr>
</table></td></tr>
<tr><td>Write the electronic configuration by writing each subshell with the number of electrons as a superscript. Remember to group subshells from the same shell.</td><td>$1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^2$</td></tr>
</table>

Worked example: Try yourself 2.3.1

WRITING ELECTRONIC CONFIGURATIONS USING THE SUBSHELL MODEL

Write the subshell electronic configuration for a vanadium atom with 23 electrons.

WS2
FPO

2.3 Review

SUMMARY

- The Bohr shell model of the atom was unable to fully explain the properties of atoms and a new model was needed to accurately describe the observed electron behaviour in atoms.
- The Schrödinger (quantum mechanical) model proposes that electrons behave as waves and occupy a three-dimensional space around the nucleus.
- The Schrödinger model describes electronic configurations in terms of shells, subshells and orbitals.
- An orbital can be regarded as a region of space surrounding the nucleus in which an electron may be found. An orbital can hold a maximum of two electrons.
  - Orbitals of similar energy are grouped in subshells that are labelled *s*, *p*, *d* and *f*.
- Each subshell has a different energy in an atom.
- Electrons fill the subshells from the lowest energy subshell to the highest energy subshell.
- The 4*s*-subshell is lower in energy than the 3*d*-subshell, so the fourth shell accepts two electrons before the third shell is completely filled.

Shell	Subshells	Orbitals in subshell	Maximum number of electrons in subshell
1	1*s*	1	2
2	2*s* 2*p*	1 3	2 6
3	3*s* 3*p* 3*d*	1 3 5	2 6 10
4	4*s* 4*p* 4*d* 4*f*	1 3 5 7	2 6 10 14

- Electronic configurations of atoms in the Schrödinger model are more detailed than the shell level configuration, as they specify the number of electrons in each subshell.
- Chromium and copper are exceptions to the predicted order of filling.

KEY QUESTIONS

Knowledge and understanding

1 Copy and complete the following table to write the electronic configuration of each of the atoms listed.

Element (atomic number)	Electronic configuration using the shell model	Electronic configuration using the subshell model
boron (5)	2,3	$1s^22s^22p^1$
lithium (3)		
chlorine (17)		
sodium (11)		
neon (10)		
potassium (19)		
scandium (21)		
copper (29)		
bromine (35)		

2 Identify the elements shown by the following electronic configurations.

a $1s^22s^22p^63s^23p^63d^{10}4s^24p^1$
b $1s^22s^22p^63s^23p^63d^{10}4s^24p^5$
c $1s^22s^22p^63s^23p^63d^{10}4s^1$
d $1s^22s^22p^63s^23p^63d^54s^2$
e $1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^{10}5s^25p^2$
f $[Kr]4d^{10}5s^25p^3$
g $[Kr]4d^{10}5s^25p^6$
h $[Ar]4s^1$

Analysis

3 Describe a limitation of the Bohr model of the atom that has been overcome by the Schrödinger model.

4 Describe the essential difference between the Bohr model and the Schrödinger model of the atom in terms of energy levels.

5 Write an electronic configuration of a sulfur atom that has been heated so that its electrons are no longer in the ground state.

2.4 The periodic table

As scientists' understanding of the atom improved and more elements were discovered, a way of organising this knowledge was needed.

The **periodic table** (Figure 2.4.1) is one of the most recognisable symbols of modern science, and more than 150 years after it was first developed by Dimitri Mendeleev, it is still an incredibly useful tool for chemists. It minimises the need to memorise isolated facts about different elements, and provides a framework on which to organise our understanding. By knowing the properties of particular elements and trends within the table, chemists are able to organise what would otherwise be an overwhelming collection of disorganised information.

FIGURE 2.4.1 This beautiful periodic table was commissioned by the Royal Australian Chemical Institute (RACI) to commemorate the 150-year anniversary of Mendeleev's original table. Each panel was a collaboration between an artist and a scientist, where the scientist provided information about the properties of the element, and the artist created a visual response to this information.

CHEMFILE

The power of a prediction

The first periodic table was developed by Russian chemist, Dimitri Mendeleev, in 1869 (see figure). When he first proposed his table, based on the observation that atomic properties seemed to vary in repeating patterns, he left spaces in the table for elements that had not yet been discovered. He even went so far as to predict the properties of these yet-undiscovered elements. He was ridiculed and largely ignored by the scientific establishment, and it was not until five years later, when gallium was discovered and matched his predicted properties almost exactly (see table), that the rest of the scientific community took his periodic table seriously.

A comparison between the properties predicted by Mendeleev, and the actual properties for element 31, gallium

Property	Mendeleev's predictions for 'eka-aluminium'	gallium
Atomic mass	68	69.723
Density (g/cm³)	6.0	5.91
Melting point (°C)	low	29.76

Russian chemist Dimitri Mendeleev (1834–1907) developed the first periodic table in 1869.

One of the remarkable things about the periodic table is that as our understanding of atomic theory has developed this periodic table remains relevant, even though it is a tool that was developed before protons, neutrons or electrons were even proposed. Each new development in atomic theory has revealed an underlying pattern in the table. In the next section you will see how the shells and valence electrons of Bohr's model correspond to the periods and groups, and you will see how Schrödinger's quantum mechanical model is reflected in the blocks.

THE MODERN PERIODIC TABLE

We now know that the number of protons (the atomic number) is what makes one element fundamentally different from another element. The elements in the modern periodic table are therefore arranged in rows in order of increasing atomic number.

Chemists use the number of electrons in the outer shell (the **valence electrons**), to organise the elements into columns. Shading is often used to highlight different aspects of the table.

The modern periodic table has several key features as can be seen in Figure 2.4.2.

- The periodic table is arranged in order of increasing atomic number.
- The horizontal rows are known as periods and are labelled 1–7.
- The vertical columns are known as groups and are labelled 1–18.
- **Main group elements** are elements in groups 1, 2 and 13–18.
- The elements in groups 3–12 are known as **transition metals**.

Some periodic tables also indicate other properties of the elements such as boiling point or whether the element is a solid, liquid or gas at room temperature.

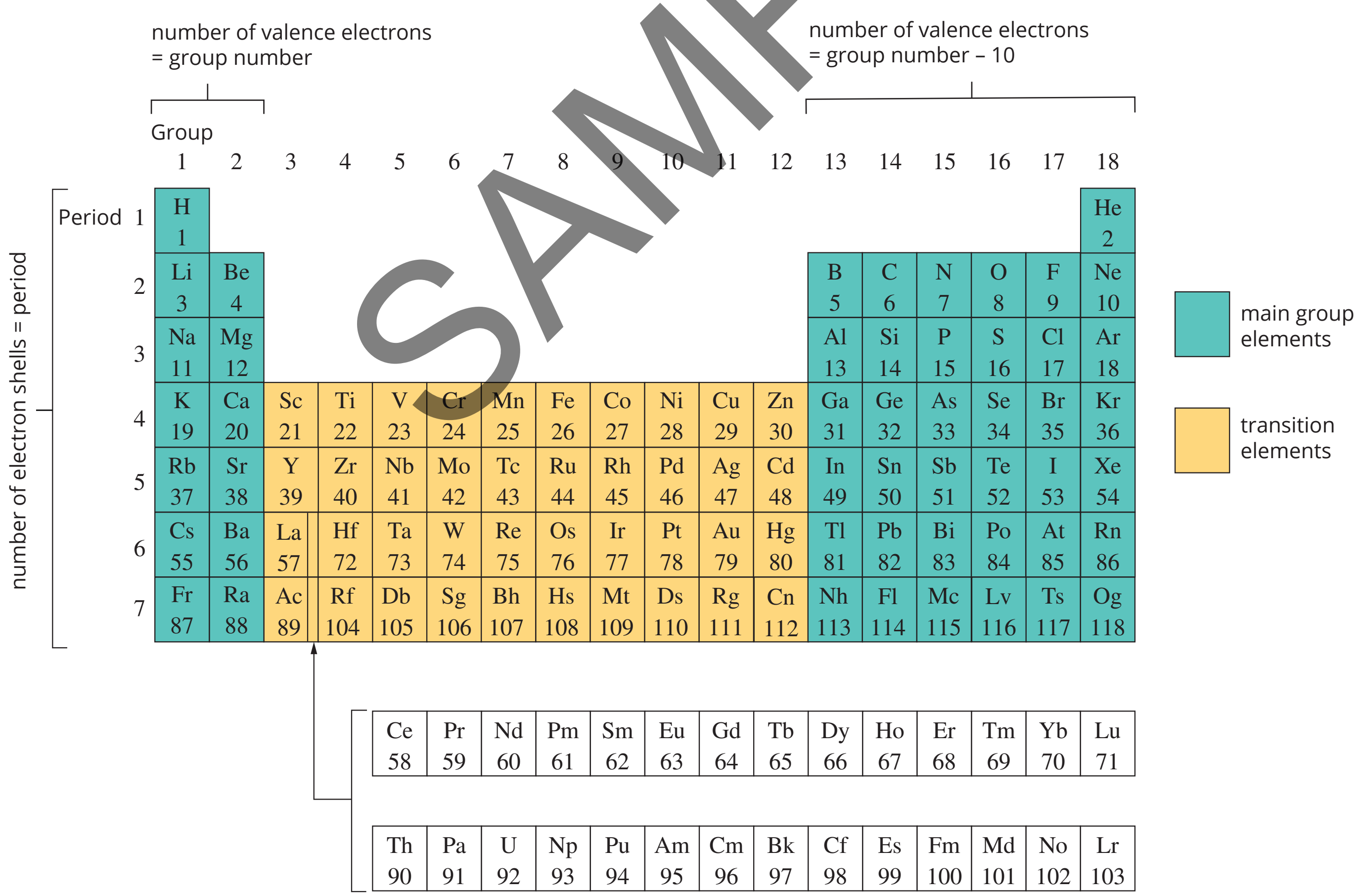


FIGURE 2.4.2 The periodic table with some key features highlighted

Groups

Elements in the periodic table are arranged into vertical columns called **groups**. For main group elements, the group number can be used to determine the number of valence electrons (outer-shell electrons) in an atom of the element.

In groups 1 and 2, the number of valence electrons is equal to the group number. For example, magnesium is in group 2 and therefore has two valence electrons.

In groups 13–18, the number of valence electrons is equal to the group number minus 10. For example, oxygen is in group 16 so oxygen has six outer-shell electrons. Similarly, neon is in group 18 so neon has eight valence electrons. Helium is an important exception. It is located in group 18 but only has two valence electrons. Helium is placed in group 18 because it is unreactive, like other group 18 elements. This information is summarised in Table 2.4.1.

TABLE 2.4.1 The number of valence electrons in elements belonging to each group

Group	Number of valence electrons
1	1
2	2
13	3
14	4
15	5
16	6
17	7
18	8*

*Helium has two valence electrons.

The electrons in the outer shell of an atom (the valence electrons) are the electrons that are involved in chemical reactions. As a consequence, the number of valence electrons determines the chemical properties that an element exhibits.

Because elements in the same group have the same number of valence electrons, elements in the same group have similar properties. For example, the **alkali metals** are elements in group 1 (with the exception of hydrogen). They are all relatively soft metals and are highly reactive with water and oxygen. Consider the electronic configurations of the atoms of the first three metals of this group:

Li $1s^2 2s^1$

Na $1s^2 2s^2 2p^6 3s^1$

K $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$

The valence shell of each atom of each element in group 1 contains one electron in an s-subshell. This similarity in the valence shell structure gives these elements similar chemical properties.

Fluorine, chlorine, bromine and iodine are **halogens** (group 17). They are all coloured and highly reactive (Figure 2.4.3). Their electronic configurations are:

F $1s^2 2s^2 2p^5$

Cl $1s^2 2s^2 2p^6 3s^2 3p^5$

Br $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^5$

I $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^5$

Notice how all these elements have a highest-energy subshell electronic configuration of s^2p^5.

The columns in the periodic table are known as groups. The group number can be used to determine the number of valence electrons in an atom of an element.

CHEMFILE

Helium supplies at risk

On the Earth, most helium is found under the Earth's crust with other natural gases. If a deposit of natural gas contains a commercially viable amount of helium, then it is extracted from the mixture. The largest commercial deposits of helium are currently found in the US, Qatar and Algeria.

We all know about the use of helium in balloons, but it has many other uses. For example, it is used in medical research and diagnostic equipment, to cool nuclear reactors and rockets and to provide an unreactive atmosphere for arc welding.

However, once helium is released into the atmosphere it is virtually impossible to recover. This means that we are using up the limited sources of helium, with some scientists estimating that we could run out within 25–30 years.

Helium is one of the most at-risk elements, with some scientists estimating that supplies could run out within 25–30 years.

FIGURE 2.4.3 Three examples of halogens. These conical flasks contain, from left to right, chlorine (Cl, pale green), bromine (Br, red-brown) and iodine (I, purple).

The **noble gases** (group 18) are a particularly interesting group. The noble gases have a very stable electron arrangement: helium and neon have full outer shells and the other members of this group have a stable octet of valence electrons (eight electrons). Chemical reactions involve the rearrangement of valence electrons to achieve a stable outer shell. Noble gases already have a stable electronic configuration, so they do not tend to lose or gain electrons. This means that the noble gases have very low reactivity.

The arrangement of electrons in atoms is responsible for the **periodicity** (periodic pattern) of element properties.

Figure 2.4.4 shows some of the special groups in the periodic table.

Period \ Group	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	H 1																	He 2
2	Li 3	Be 4											B 5	C 6	N 7	O 8	F 9	Ne 10
3	Na 11	Mg 12											Al 13	Si 14	P 15	S 16	Cl 17	Ar 18
4	K 19	Ca 20	Sc 21	Ti 22	V 23	Cr 24	Mn 25	Fe 26	Co 27	Ni 28	Cu 29	Zn 30	Ga 31	Ge 32	As 33	Se 34	Br 35	Kr 36
5	Rb 37	Sr 38	Y 39	Zr 40	Nb 41	Mo 42	Tc 43	Ru 44	Rh 45	Pd 46	Ag 47	Cd 48	In 49	Sn 50	Sb 51	Te 52	I 53	Xe 54
6	Cs 55	Ba 56	La 57	Hf 72	Ta 73	W 74	Re 75	Os 76	Ir 77	Pt 78	Au 79	Hg 80	Tl 81	Pb 82	Bi 83	Po 84	At 85	Rn 86
7	Fr 87	Ra 88	Ac 89	Rf 104	Db 105	Sg 106	Bh 107	Hs 108	Mt 109	Ds 110	Rg 111	Cn 112	Nh 113	Fl 114	Mc 115	Lv 116	Ts 117	Og 118

Ce 58	Pr 59	Nd 60	Pm 61	Sm 62	Eu 63	Gd 64	Tb 65	Dy 66	Ho 67	Er 68	Tm 69	Yb 70	Lu 71
Th 90	Pa 91	U 92	Np 93	Pu 94	Am 95	Cm 96	Bk 97	Cf 98	Es 99	Fm 100	Md 101	No 102	Lr 103

- alkali metals
- alkaline earth metals
- halogens
- noble gases
- lanthanoids
- actinoids

FIGURE 2.4.4 The periodic table highlighting some special groups

Periods

The horizontal rows in the periodic table are called **periods** and are numbered 1–7. The number of a period gives information about the electronic configuration of an element. The period an element is located within is equal to the number of occupied electron shells in the element's atoms. For example, the outer shell of magnesium and chlorine is the third shell and both of these elements are in period 3:

Mg $1s^2 2s^2 2p^6 3s^2$

Cl $1s^2 2s^2 2p^6 3s^2 3p^5$

Similarly, the elements in period 5 all have outer shell electrons in the fifth shell.

The rows of the periodic table are called periods. The period number tells you the number of occupied electron shells in the element's atoms.

Blocks

As understanding of atomic theory developed into Schrödinger's quantum mechanical model, a new pattern emerged in the periodic table. We can now see that it can be divided into four main **blocks**. The elements in each block all have the same type of subshell as the highest energy subshell (*s*, *p*, *d* or *f*).

Figure 2.4.5 shows which groups of elements fall into the four different blocks.

The blocks in the periodic table contain elements that have the same type of subshell as the highest energy subshell (*s*, *p*, *d* or *f*).

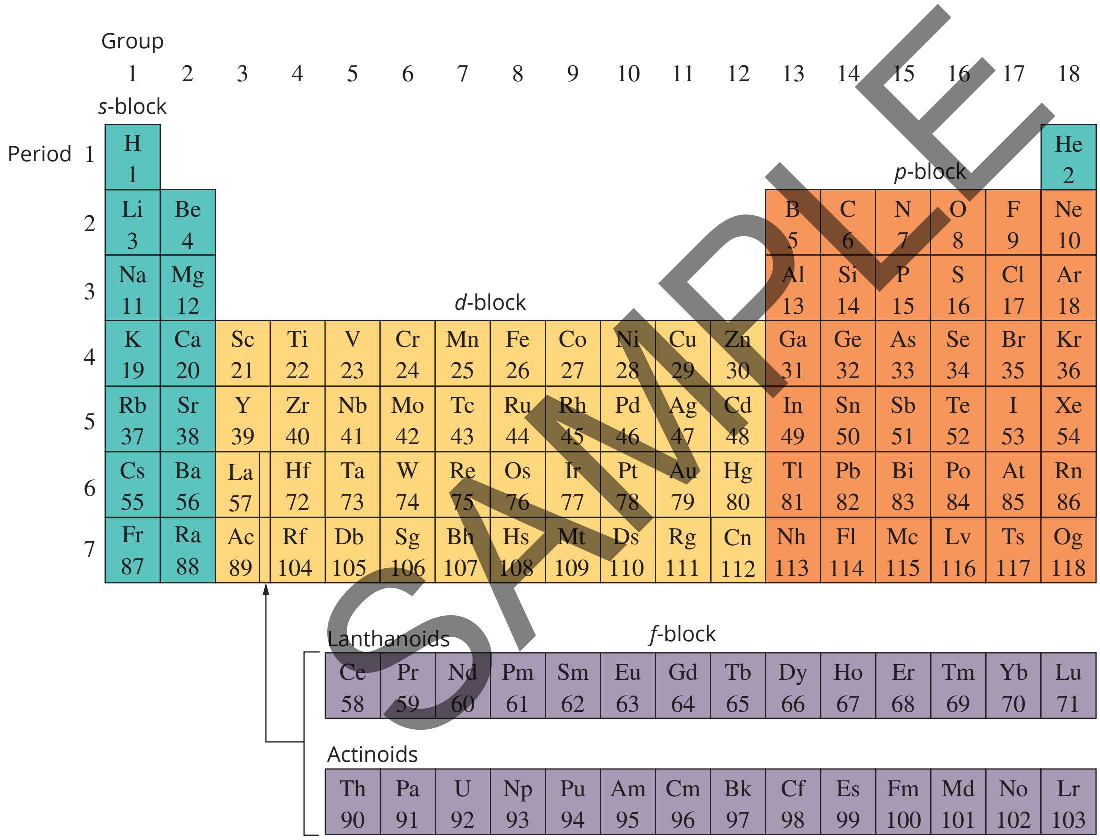


FIGURE 2.4.5 Colour is used to distinguish between the *s*, *p*, *d* and *f*-blocks of elements.

CASE STUDY ANALYSIS

Naming elements on the periodic table

Ancient elements

Twelve elements were known in ancient times. These included gold, silver, mercury and sulfur. The original names for these elements were derived from Latin. People at this time did not have an understanding of the modern definition of an element, so materials like soda (sodium oxide) and calomel (mercury chloride) were at first thought to be elemental.

Searching for new elements

Later in the eighteenth century and into the nineteenth century, scientists became able to extract gaseous elements from the air and isolate solid elements from the ground. Elements often took names from the Latin words for where they were found. Such elements included silicon, Si (found in sand, known in Latin as *silex*) and calcium, Ca (found in limestone, known in Latin as *calx*). Mercury, Hg, a silvery metal that is a liquid at room temperature, was given the name *hydrargyrum*, meaning silver water.

As analytical techniques improved and more elements were found, naming became more imaginative, with names being derived from Greek roots, detected characteristics, places or mythology: for example, the inert gas argon (from the Greek *ἀργόν*, inactive), rubidium (from the Latin *rubidius*, deep red), germanium (from the Latin *germania*, meaning Germany), and vanadium (after the Scandinavian goddess Vanadis).

By 1829, 55 elements were known. The first modern periodic table, constructed in 1869, included 64 elements, and by 1914 a total of 72 elements had been discovered.

Synthetic elements and atomic science

The first synthetic element, technetium, was made in 1936. New elements were discovered after 1940; of these, many were discovered during thermonuclear bomb tests (see Figure 2.4.6), and more recently in particle accelerators. This has brought the total number of known elements in the periodic table to 118. The synthetic elements have large, unstable nuclei and not much is known about their chemistry. Many, such as einsteinium or americium, were named after a scientist or place. The Cold War made naming of these elements contentious, but since 1999 the names of all new elements have been decided by a panel of the International Union of Pure and Applied Chemistry (IUPAC).

FIGURE 2.4.6 Thermonuclear bomb tests like this were the source of many of the first synthetic elements discovered.

Analysis

1. Using your knowledge of element symbols and the periodic table, deduce the common name of each of the following elements from its non-English name.
   - **a** *ferrum*
   - **b** *kalium*
   - **c** *wolfram*
   - **d** *plumbum*
   - **e** *hydrargyrum*
2. IUPAC has a set of guidelines for naming new elements. What are these guidelines? (You will need to research your answer.)
3. The periodic table was most recently updated by IUPAC in both 2012 and 2016. In these two updates, elements 113, 115, 117 ,118 (2016), 114 and 116 (2012) were added. Research the names of these elements and outline why they were given their particular names.

CRITICAL ELEMENTS

Of the 118 elements in the periodic table, many exist in only very small quantities on earth. In recent decades, some of these rare elements have been used in bulk for the first time to produce new electronics, medicines and catalysts. Unfortunately, our use of these materials is not sustainable, as there is currently very little recycling or recovery of these elements. As a consequence, over 40 elements have been identified as being '**endangered**', with supplies of some elements predicted to run out in less than 100 years. Figure 2.4.7 highlights elements that have been identified as endangered. (Note this classification is constantly changing as our demand for elements changes, and new deposits are found.)

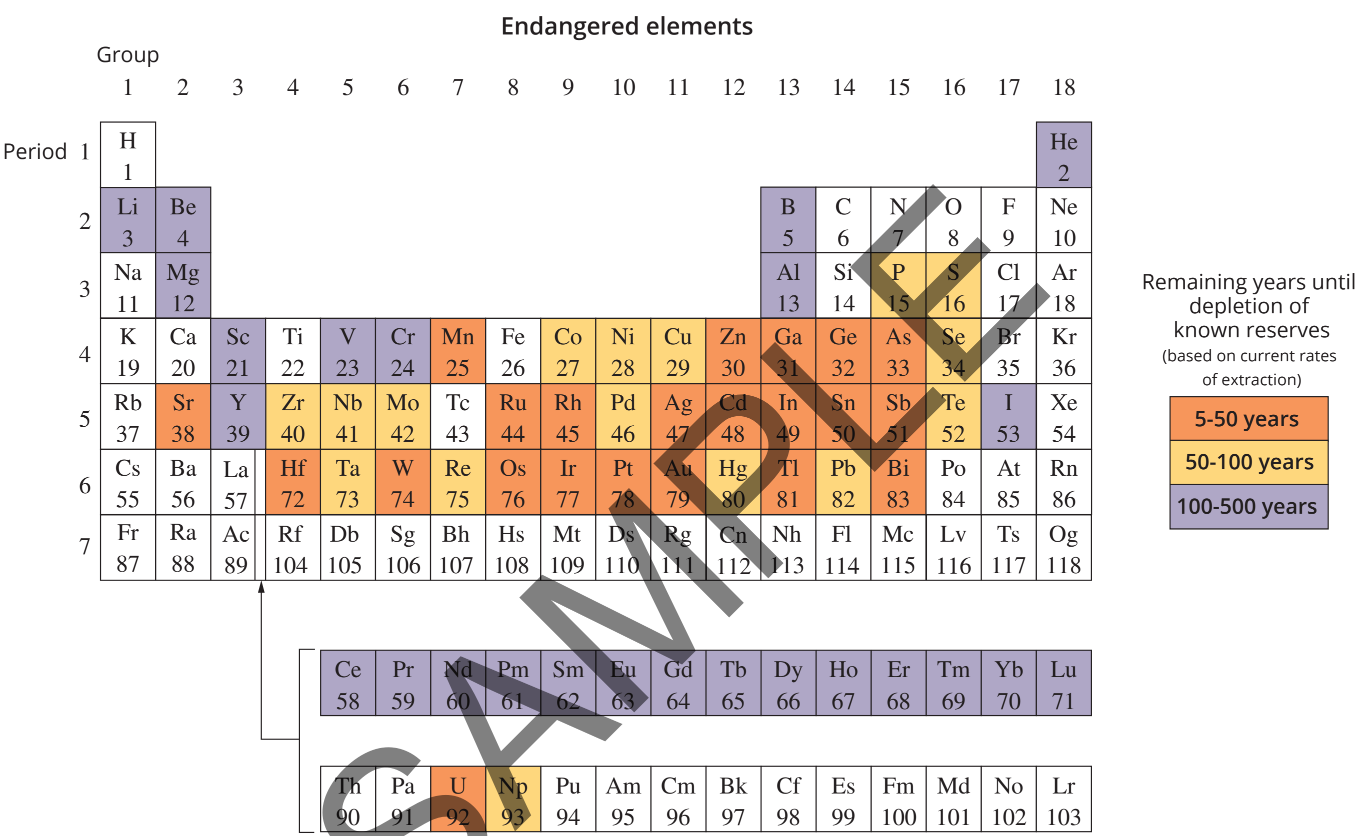


FIGURE 2.4.7 Some of the elements identified as endangered, with estimates for how long they will be available.

It is not just the reserves of an element that are important in ensuring sustainability. In recent years the concept of a critical element has developed. **Critical elements** are elements that are heavily relied on for industry and society in areas such as renewable energy, electronics, food supply and medicine. Some scientists and economists believe that unhindered access to these elements is required for the essential running of society. Critical elements face some form of supply uncertainty. Elements may be considered critical for a number of reasons:

- Only small deposits are available on earth, and these are fast disappearing (endangered elements). For example, iridium, platinum, osmium and palladium (see Figure 2.4.7).
- Supply of these elements is centred in places of war and conflict (conflict elements). **Conflict elements** include tin, gold, tungsten and tantalum, which are essential in mobile phone production (Figure 2.4.8a). These elements are mined in areas of war and conflict, and routinely use child-labour in their mining, which makes their use non-sustainable (Figure 2.4.8b).

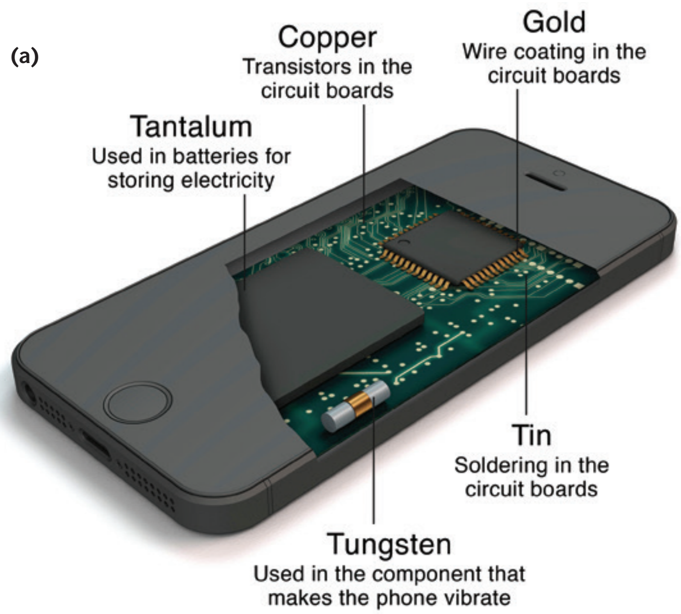


FIGURE 2.4.8 (a) Tin, gold, tungsten and tantalum are all used in mobile phone production. Their use is not sustainable as they are often mined in areas of war and conflict. (b) Child labour is routinely used in the mining of these elements. This young teenage worker is pictured outside a gold mine in Africa.

- There is little to no recycling and recovery of the element, so reserves are being used up. For example, the rare earth elements (Lanthanides). Figure 2.4.9 shows the current rates of recycling for a number of these elements.
- They have significant economic importance, there are no viable substitutes and deposits are concentrated in a small number of countries and so supply could be at risk (critical raw materials). For example, deposits of tungsten, antimony, molybdenum, germanium, gallium and indium are concentrated in China. Deposits of platinum, palladium, rhodium, ruthenium and vanadium are concentrated in South Africa.
- Expansion of new technologies in recent decades has greatly increased our use of these elements. For example, the use of helium in medical technologies and phosphorus in fertilisers.

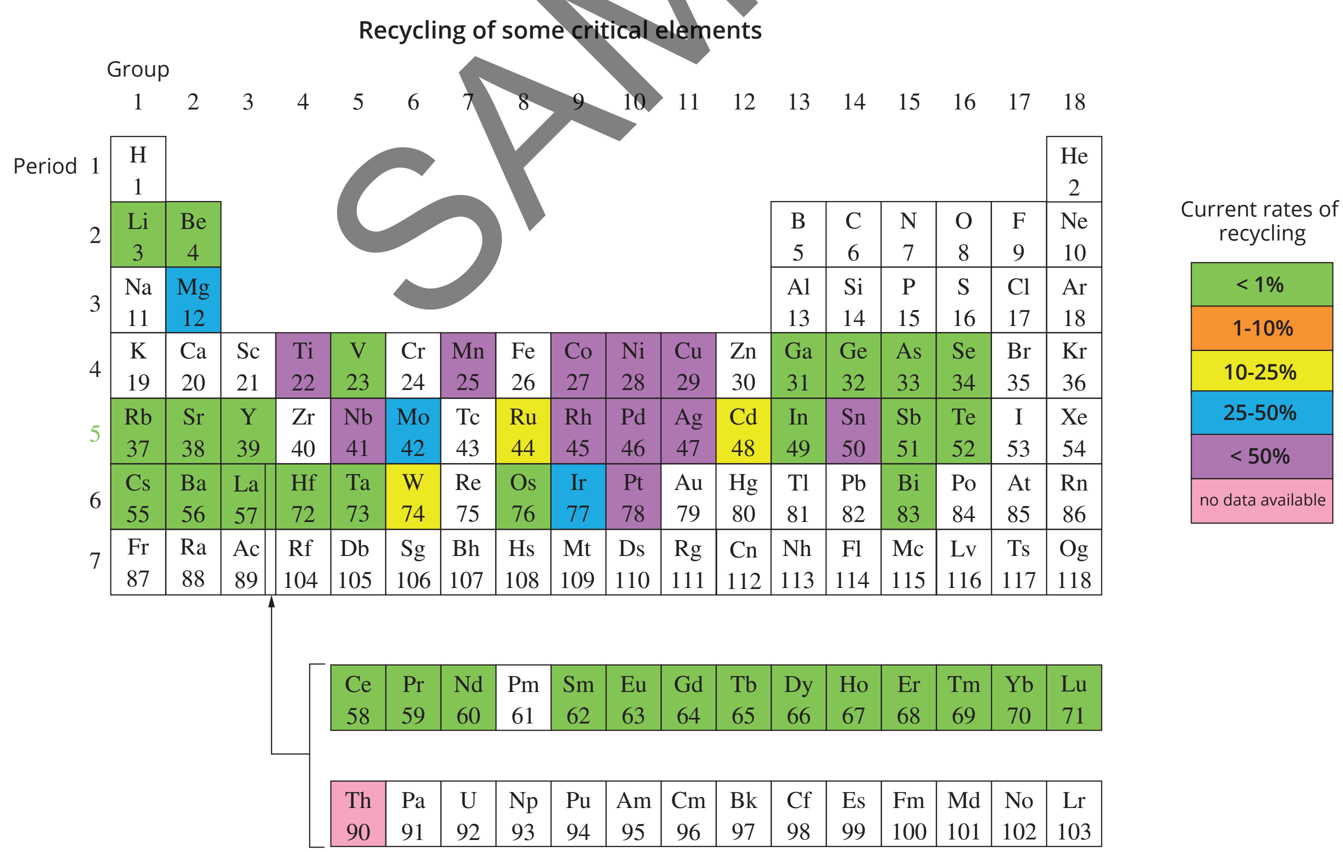


FIGURE 2.4.9 There is little to no recycling of many of the elements currently being used as raw materials.

The concept of critical elements draws attention to the non-sustainability of our consumption. Recycling campaigns have been developed globally to help make products more sustainable. One of the most well-known was in the lead up to the 2020 Tokyo Olympic Games, when the Japanese people collected approximately 79 000 tonnes of old mobile phones and other electronic waste. This was recycled to recover the precious metals inside to make 100% of the gold, silver and bronze Olympic medals presented to the athletes (Figure 2.4.10). The recycling of metals is discussed in more detail in Chapter 4.

Our use of many elements is not currently sustainable. New ways of recovering and recycling these elements are needed.

FIGURE 2.4.10 Emma McKeon of Australia, Sarah Sjöström of Sweden and Pernille Blume of Denmark show off their 2020 Tokyo Olympic medals, made from metal recovered from 79 000 tonnes of electronic waste collected by the Japanese people in the lead-up to the games.

A new approach to recycling

Phosphorus plays a crucial part in the production of the world's food. It is used to create fertilisers that enable greater crop yields (Figure 2.4.11). However, it is estimated that supply of phosphorus will become depleted sometime this century. Experts predict this could have very serious consequences for world food supplies. In order to maintain overall supply of phosphorus, it may even be necessary for it to be recovered from animal and human waste.

One suggested method of recovery is from human urine, using a toilet with a urine diversion and dehydration unit like that shown in Figure 2.4.12. The phosphorus can eventually be precipitated out of the urine as calcium phosphate.

Interestingly, there is a historical precedent to this method of phosphorus recovery. The element was first discovered in human urine by alchemist Hennig Brand. In 1669, Brand accidentally discovered phosphorus while searching for the mythical 'philosopher's stone'. His method reportedly began with almost 6000 litres of urine (Figure 2.4.13).

FIGURE 2.4.11 Phosphorus is a vital part of the fertilisers need to feed the population of the world.

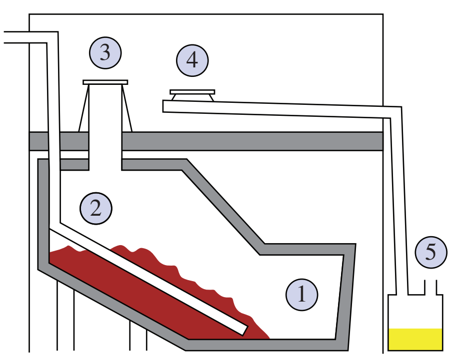


FIGURE 2.4.12 A diagram of a composting toilet with urine diversion and dehydration: 1. solid waste compartment; 2. ventilation pipe; 3. toilet; 4. urinal; 5. urine collection and dehydration

FIGURE 2.4.13 This painting, *The Alchemist in Search of the Philosopher's Stone* depicts the discovery of phosphorus, by alchemist Hennig Brand in 1669, in an experiment that reportedly began with almost 6000 litres of urine.

2.4 Review

SUMMARY

- The periodic table is a tool for organising elements according to their chemical and physical properties.
- The elements of the modern periodic table are arranged in order of increasing atomic number.
- Columns in the periodic table are known as groups and are numbered 1–18.
- The main group elements are in groups 1, 2 and 13–18 in the periodic table.
- The number of valence electrons a for main group element can be determined by the group in which it is located.
- Elements in between the main group elements (groups 3–12) are known as transition metals.
- Rows in the periodic table are known as periods and are numbered 1–7.
- The number of occupied electron shells of an atom of an element is equal to the number of the element's period.
- The periodic table has four main blocks of elements; the elements in each block have the same type of subshell (*s*, *p*, *d* or *f*) as their highest energy subshell.
- Our use of many elements is currently non-sustainable, which has led to them being described as critical elements.
- To ensure critical elements are available for future generations, methods of recycling and recovery must be developed.

KEY QUESTIONS

Knowledge and understanding

1 In the periodic table, group is to column as period is to what?

2 Define the terms 'period', 'group' and 'block'

3 What is the name given to elements in groups 1, 2 and 13–18?

4 State the IUPAC name for the following groups with reference to the periodic table.
- **a** 1
- **b** 2
- **c** 17
- **d** 18

5 How many valence electrons are in atoms of elements found in the following groups?
- **a** 2
- **b** 13
- **c** 15
- **d** 18

6 What is the subshell electronic configuration of an atom of an element in period 4 and group 1?

Analysis

Use the periodic table in Figure 2.4.2 on page XX to answer question 7.

7 **a** In which group of the periodic table will you find the following?
- **i** B
- **ii** Cl
- **iii** Na
- **iv** Ar
- **v** Si
- **vi** Pb

b Determine the name and symbol of the following elements. In addition, write the electronic configuration of each element.
- **i** second element in group 14
- **ii** second element in period 2
- **iii** element that is in group 18 and period 3

c Identify the period of the periodic table in which each of the following elements belongs.
- **i** K
- **ii** F
- **iii** He
- **iv** H
- **v** U
- **vi** P

8 Explain the term 'critical element'. Give three examples of elements that are considered critical elements.

2.5 Trends in the periodic table

The periodic table does not just provide information about an element's electronic configuration. It can also be used as a tool for summarising the relative properties of elements and explaining the trends observed in those properties.

You have already seen how the group number of an element indicates how many valence electrons an atom of that element has. The period indicates how many electron shells are occupied in an atom of an element. Properties such as atomic radius, **electronegativity** (the electron-attracting power of an atom) and **first ionisation energy** (the energy required to remove one electron from an atom) also show common trends in the periodic table.

Periodic trends were observed by Dimitri Mendeleev and formed the basis of the table of the elements that he first published in 1869. Mendeleev described the way the properties of the elements vary as the **periodic law**.

TRENDS IN ELECTRONIC CONFIGURATION

To understand the reason for the periodicity (repeating pattern) of element properties, consider the electronic configurations of the alkali metals (group 1). The elements in this group (lithium, sodium, potassium, rubidium and caesium) are all relatively soft metals and are highly reactive with water and oxygen.

Li $1s^22s^1$
Na $1s^22s^22p^63s^1$
K $1s^22s^22p^63s^23p^64s^1$
Rb $1s^22s^22p^63s^23p^63d^{10}4s^24p^65s^1$
Cs $1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^{10}5s^25p^66s^1$

These elements have similar valence shell electronic configurations – all have one electron in an *s*-subshell. This similarity in a group's arrangement of electrons gives elements similar properties and is responsible for the periodicity of element properties.

You can also see that the number of electron shells increases moving down the group. The increase in electron shells means that the valence electrons are in higher energy subshells and have a weaker attraction to the nucleus. The decrease in the attractive force between the valence electrons and the nucleus as you move down a group causes trends in properties to be observed within a group.

EFFECTIVE NUCLEAR CHARGE

The **effective nuclear charge** (sometimes referred to as core charge) of an atom is a measure of the attractive force felt by the valence shell electrons towards the nucleus. Effective nuclear charge can be used to predict the properties of elements and explain trends observed across a period in the periodic table.

Consider an atom of lithium, which has an atomic number of three. It has three protons in its nucleus, two electrons in the first shell and one electron in the second shell (Figure 2.5.1).

The valence shell electron is attracted to the three positive charges in the nucleus. This electron is also repelled by the two electrons in the inner shell. The electrons in the inner shell shield the valence shell electron from the attraction of the nucleus. The valence shell electron is effectively attracted to the nucleus as if there were a +1 nuclear charge. This atom is therefore said to have an effective nuclear charge of +1.

Similarly, an atom of chlorine has 17 protons and seven valence shell electrons; the number of electrons in the inner shells is 10. The effective nuclear charge of a chlorine atom is 17 – 10 = +7.

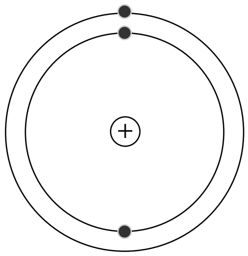


FIGURE 2.5.1 A lithium atom with one valence electron and two electrons in the inner shell. The atom has an effective nuclear charge of +1.

Effective nuclear charge = number of protons in the nucleus – number of total inner-shell electrons

Worked example 2.5.1

EFFECTIVE NUCLEAR CHARGE

Determine the effective nuclear charge of an atom of aluminium.

Thinking	Working
Determine the number of electrons in an atom of the element, using the periodic table as a reference.	The atomic number of aluminium is 13. Therefore, an atom of aluminium has 13 protons and 13 electrons.
Use the number of electrons to determine the electronic configuration.	With 13 electrons the electronic configuration is $1s^22s^22p^63s^23p^1$.
Determine the effective nuclear charge. Effective nuclear charge = number of protons – number of inner-shell electrons	The third shell is the valence shell in this atom. There are 10 inner-shell electrons, which in this atom are electrons in the first and second shell. Effective nuclear charge = 13 – 10 = +3

Worked example: Try yourself 2.5.1

EFFECTIVE NUCLEAR CHARGE

Determine the effective nuclear charge of an atom of fluorine.

TABLE 2.5.1 Effective nuclear charges of main group elements

Group	Effective nuclear charge
1	+1
2	+2
13	+3
14	+4
15	+5
16	+6
17	+7
18*	+8

*Helium has an effective nuclear charge of +2.

The effective nuclear charge experienced by the valence shell electrons in atoms of elements increases from left to right across a period, as you have seen for sodium and chlorine. The effective nuclear charge of an atom of a main group element is equivalent to the number of valence electrons in the atom, as summarised in the Table 2.5.1.

Table 2.5.2 summarises how the attraction between the nucleus and valence electrons changes in the periodic table.

TABLE 2.5.2 The changes in attraction between the nucleus and valence electrons within groups and periods of the periodic table

	Trend in effective nuclear charge	Trend in attraction between the nucleus and valence electrons
Down a group	remains constant	Effective nuclear charge stays constant down a group, but the valence electrons are held less strongly as they are further from the nucleus because there are more shells in the atom.
Left to right across a period	increases	The valence electrons are more attracted to the nucleus as the effective nuclear charge increases.

All of the observed trends in the periodic table can be related to these changes in attraction between the valence electrons and the nucleus.

ELECTRONEGATIVITY

Electronegativity is the ability of an atom to attract electrons towards itself when forming a chemical bond. As the positive pull in an atom comes from the nucleus, when the effective nuclear charge is greater, the electronegativity increases. Figure 2.5.2 on the following page shows the electronegativity of many of the main group elements. Electronegativity values for the noble gases (group 18) are not listed because the elements have a stable outer shell and do not readily form bonds with other atoms.

Electronegativity is the ability of an atom to attract electrons towards itself when forming a chemical bond. It is a measure of how strongly an atom pulls on the electrons of nearby atoms.

Electronegativity increases across a period.

Electronegativity decreases down a group.

1	2	13	14	15	16	17
Li 1.0	Be 1.6	B 2.0	C 2.6	N 3.0	O 3.4	F 4.0
Na 0.9	Mg 1.3	Al 1.6	Si 1.9	P 2.2	S 2.6	Cl 3.2
K 0.8	Ca 1.0	Ga 1.8	Ge 2.0	As 2.2	Se 2.6	Br 3.0
Rb 0.8	Sr 1.0	In 1.8	Sn 2.0	Sb 2.1	Te 2.1	I 2.7
Cs 0.8	Ba 0.9	Tl 2.0	Pb 2.3	Bi 2.0	Po 2.0	At 2.2
Fr 0.7	Ra 0.9					

FIGURE 2.5.2 The electronegativity of elements generally decreases down a group and increases across a period, from left to right.

The trends observed in the electronegativity of the elements are summarised in Table 2.5.3.

TABLE 2.5.3 Trends in electronegativity in groups and periods of the periodic table

	Trend in electronegativity	Explanation
Down a group	decreases	The effective nuclear charge stays constant and the number of shells increases down a group. Therefore, valence electrons are less strongly attracted to the nucleus as they are further from the nucleus. As a result, electronegativity decreases.
Left to right across a period	increases	The number of occupied shells in the atoms remains constant but the effective nuclear charge increases across a period. Therefore, the valence electrons become more strongly attracted to the nucleus. As a result, electronegativity increases.

Atomic radius decreases across a period. This is because the effective nuclear charge increases across a period, so the valence shell electrons are pulled in to the nucleus.

ATOMIC RADIUS

Atomic radius is a measurement used for the size of atoms. It can be regarded as the distance from the nucleus to the valence shell electrons. It is usually measured by halving the distance between the nuclei of two atoms of the same element that are bonded together. Figure 2.5.3 depicts the atomic radii of many of the main group elements. As you can see, the atomic radius decreases as you move across a period. This is because as the effective nuclear charge increases, the valence shell electrons are pulled in more tightly towards the nucleus.

Table 2.5.4 explains the trends in atomic radius in the periodic table.

TABLE 2.5.4 Trends in atomic radii in the periodic table

	Trend in atomic radius	Explanation
Down a group	increases	Effective nuclear charge stays constant and the number of shells increases as you move down a group. As a result, atomic radius increases.
Left to right across a period	decreases	As you move across a period, the number of occupied shells in the atoms remains constant but the effective nuclear charge increases. The valence electrons become more strongly attracted to the nucleus, so atomic radius decreases across a period.

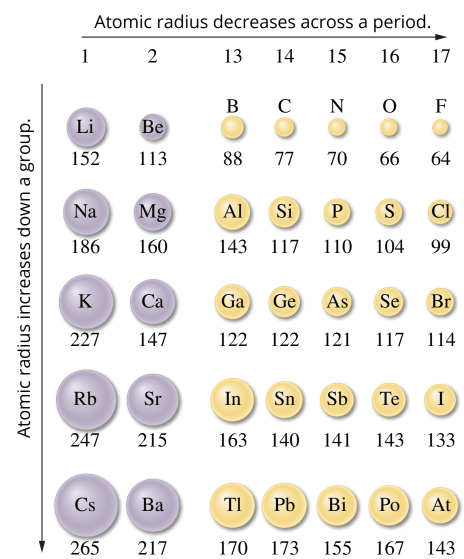


FIGURE 2.5.3 The relative sizes of atoms of selected main group elements. The atomic radii are given in picometres (pm). A picometre is 10^{-12} m.

FIRST IONISATION ENERGY

When an element is heated, its electrons can move to higher energy shells. If an atom is given sufficient energy, an electron can be completely removed from the atom. If this occurs, the atom will now have one less electron than the number of protons in the nucleus, and becomes a positively charged ion.

First ionisation energy is the energy required to remove one electron from an atom of an element in the gas phase. For example, the first ionisation energy of sodium is 494 kJ per mole of sodium atoms. (A mole is a way of counting atoms. You will learn how to use the mole in Chapter 8.)

First ionisation energy is the energy required to remove one electron from an atom of an element.

Figure 2.5.4 shows the first ionisation energies of most main group elements.

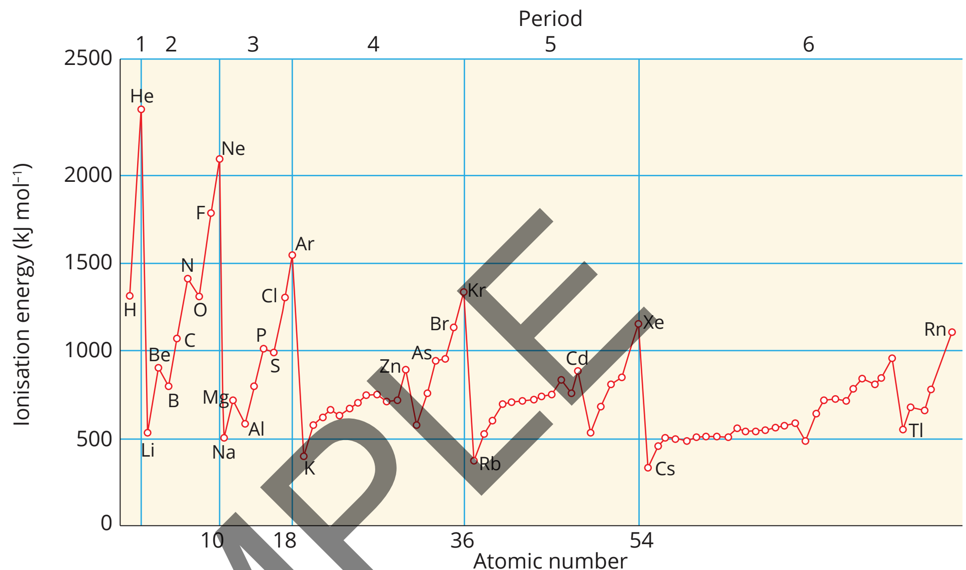


FIGURE 2.5.4 The first ionisation energy generally increases from left to right across a period.

The magnitude of the first ionisation energy reflects how strongly the valence electrons are attracted to the nucleus of the atoms. The more strongly the valence electrons are attracted to the nucleus, the more energy is required to remove one of them from the atom and the higher the first ionisation energy. When this data is compared to position on the periodic table, the trends described in Table 2.5.5 emerge.

TABLE 2.5.5 The trend in ionisation energies in groups and periods of the periodic table

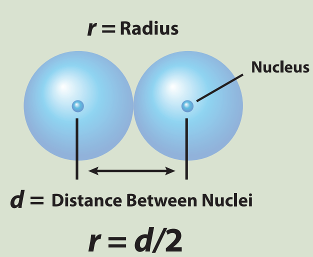
	Trend in first ionisation energy	Explanation
Down a group	decreases	Effective nuclear charge stays constant and the number of shells increases down a group. Therefore, the valence electrons are less attracted to the nucleus as they are further from the nucleus. As a result, the energy required to overcome the attraction between the nucleus and the valence electron is less, and the first ionisation energy decreases down a group.
Left to right across a period	increases	Effective nuclear charge increases and the number of occupied shells remains constant across a period. As a result, the valence electrons become more strongly attracted to the nucleus, and more energy is required to remove an electron. Therefore, first ionisation energy increases across a period.

The *s*- and *p*-block elements of the periodic table follow these patterns. As the effective nuclear charge increases across a period, so too does the ionisation energy. As atomic radius increases down a group (due to the additional electron shell), the electrons are further from the nucleus. Therefore, the valence electrons in elements lower in a group can be removed more easily, meaning the ionisation energy decreases.

CHEMFILE

Measuring atoms

Atoms do not have sharply defined boundaries and so it is not possible to measure their radii directly. One method of obtaining atomic radius, and therefore an indication of atomic size, is to measure the distance between nuclei of atoms of the same element in molecules. For example, in a hydrogen molecule (H_2) the two nuclei are 74 picometres (pm) apart. The radius of each hydrogen atom is assumed to be half of that distance, i.e. 37 pm. The figure below shows how the radius of atoms is determined.



The radius of an atom can be determined by measuring the distance between two nuclei of the same element in a molecule. Half this distance will be the atomic radius.

TRENDS IN METALS AND NON-METALS

One of the most useful distinctions that can be made between elements in the periodic table is the difference between metals, metalloids and non-metals, as can be seen in Figure 2.5.5. **Metalloids** share properties of both metals and non-metals.

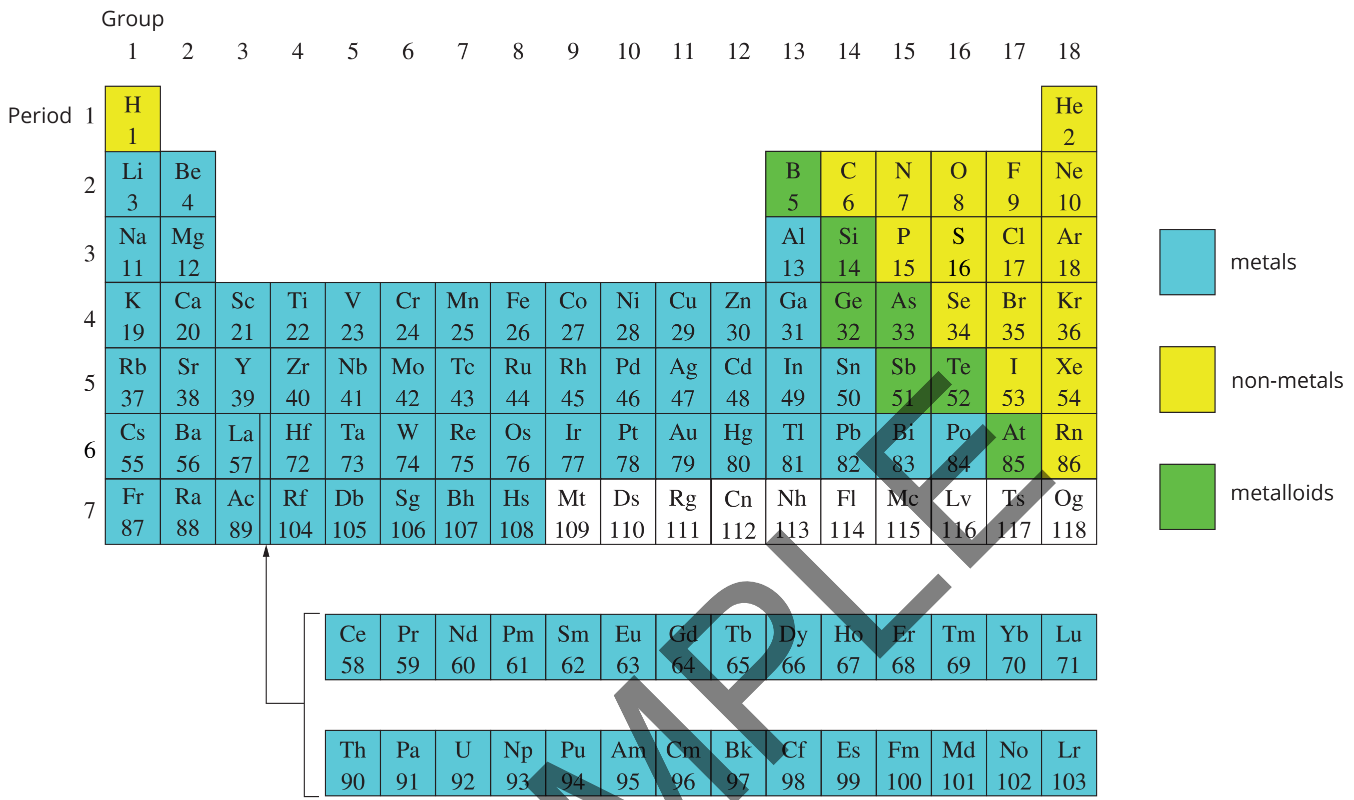


FIGURE 2.5.5 Periodic table highlighting the metals, metalloids and non-metals

As you will discover in the coming chapters, an understanding of whether an element is a metal or non-metal helps you to understand the way it will bond with other elements. The reactivity of both metals and non-metals is also something that follows a trend in the periodic table.

In general, the properties of metals arise from the fact that they readily lose electrons. As you go down a group, the valence shell electrons are further away from the pull of the nucleus and are therefore lost more easily. This means that as you go down a group, metals become more reactive.

Across a period, the opposite occurs, as the effective nuclear charge increases. This means the most reactive metals are in group 1.

Consequently, the most reactive metal in the periodic table is in the bottom left corner: francium.

In contrast, the properties of non-metals arise from the fact that they readily gain electrons. In general, non-metals become less reactive as you go down a group, as the valence electrons become further away from the nucleus. Non-metals become more reactive across a period, as the effective nuclear charge increases. The exception to this are the noble gases, which are inert.

Consequently, the most reactive non-metal is in the top right of the periodic table (excluding the noble gases): fluorine.

Knowing whether an element is a metal, non-metal or metalloid can help you to predict the way it will react and bond with other elements.

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2.5 Review

SUMMARY

- The effective nuclear charge of an atom is a measure of the attractive force felt by the valence electrons towards the nucleus.
- The effective nuclear charge is calculated by subtracting the total number of inner-shell electrons from the number of protons in the nucleus.
- Electronegativity is the ability of an element to attract electrons towards itself when forming a chemical bond.
- Atomic radius is a measurement used for the size of atoms. It can be regarded as the distance from the nucleus to the outermost electrons.
- The first ionisation energy is the energy required to remove one electron from an atom of an element in the gas phase.
- The table on the right summarises how these properties have specific trends within the groups and periods of the periodic table.

Property	Down a group	Across a period (left to right)
effective nuclear charge	no change	increases
electronegativity	decreases	increases
atomic radius	increases	decreases
first ionisation energy	decreases	increases
metal reactivity	increases	decreases
non-metal reactivity	decreases	increases

- Many trends in the physical properties of elements in the periodic table can be explained using two key ideas.
  - From left to right across a period, the effective nuclear charge of atoms increases, so the attractive force felt between the valence electrons and the nucleus increases.
  - Down a group, the number of shells in an atom increases so that the valence electrons are further from the nucleus and are held less strongly.

KEY QUESTIONS

Knowledge and understanding

1 Define the term 'effective nuclear charge' of an atom and determine the effective nuclear charge of an atom of carbon.

2 State the relationship between electronegativity and effective nuclear charge.

3 Identify whether the following elements are metals, metalloids or non-metals using the periodic table.

a K
b N
c Ge
d Cu
e Sn
f I

Analysis

4 Figure 2.5.2 gives electronegativity values for the elements in groups 1, 2 and 13–17 of the periodic table.

a Give the name and symbol of the element that has the:

i highest electronegativity
ii lowest electronegativity.

b In which group do you see the following?

i greatest change in electronegativity as you go down the group
ii smallest change in electronegativity as you go down the group

5 Based on your understanding of the trends in the periodic table, sort the following in order of increasing atomic radius:
N, B, Ga, Al, Cl

6 Explain why the number of subatomic particles in an atom increases across a period but the size of the atom decreases.

7 By referring to the periodic table, organise the following elements in order of increasing first ionisation energies:
Na, He, Al, K, S, Ca and P

8 Suggest a reason why caesium is more reactive with water than lithium.

Chapter review

02

KEY TERMS

alkali metals
atom
atomic number
atomic radius
atomic theory of matter
block (periodic table)
Bohr model
chemical symbol
compound
condensed electronic configuration
conflict element
critical element
effective nuclear charge
electron
electronic configuration
electron shell
electronegativity
electrostatic attraction
element
emission spectrum
endangered element
energy level
excited state
first ionisation energy
flame test
ground state
group (periodic table)
halogen
ion
isotope
main group element
mass number
matter
metalloid
model
neutron
noble gas notation
noble gas
nucleon
nucleus
nuclide notation
orbital
period
periodic law
periodic table
periodicity
proton
quantised
quantum mechanics
radioactive
Schrödinger model
subatomic particle
subshell
transition metal
valence electrons
valence shell

REVIEW QUESTIONS

Knowledge and understanding

1 The maximum number of electrons in the first three shells of an atom is:
A 8,8,18
B 2,8,12
C 2,8,8
D 2,8,18

2 Isotopes of an element contain different numbers of (more than one answer possible):
A protons
B neutrons
C nucleons
D electrons

3 Of the electronic configurations below, which one represents an atom in an excited state?
A $1s^2 2s^2 2p^6 3s^1$
B $1s^2 2s^2 2p^6 3s^2 3p^6$
C $1s^2 2s^2 2p^6 3s^2 3p^5 4s^2$
D $1s^2 2s^2 2p^3$

4 Select the response that best describes how the elements in the periodic table are arranged.
A by the number of neutrons in their nucleus
B by the number of protons in their nucleus
C by increasing mass number
D by increasing atomic mass

5 Identify the most reactive non-metal from the following elements.
A caesium
B chlorine
C fluorine
D oxygen

6 An atom of chromium can be represented by the symbol $^{52}_{24}Cr$
a State its atomic number and mass number.
b Determine the number of electrons, protons and neutrons in the chromium atom.

7 What is the name of the element that has an electronic shell configuration of 2,8,2?

8 Determine the Schrödinger model of electronic configuration that corresponds to the shell model electronic configuration 2,8,6.

9 Use the periodic table to determine the period and block of the following elements.
a Hydrogen
b Carbon
c Phosphorus
d Copper
e Uranium

10 Select the appropriate word from the following options to complete the sentences below: increase, increases, decrease, decreases

a The force of attraction between the nucleus and valence electrons ___________ in a period from left to right.

b Atomic radii of elements ___________ in a period from left to right.

c Atomic radii of elements ___________ in a group from top to bottom.

d Metallic character of elements ___________ from top to bottom in a group.

Application and analysis

11 Two atoms both have 20 neutrons in their nucleus. The first has 19 protons and the other has 20 protons. Are they isotopes? Why or why not?

12 Analyse the following list of atoms.

$^{27}_{13}A$ $^{40}_{20}B$ $^{32}_{16}C$ $^{14}_{7}D$ $^{8}_{3}E$ $^{15}_{7}F$ $^{19}_{9}G$ $^{20}_{9}H$ $^{4}_{2}I$

a Identify which pairs of atoms are isotopes of the same element.

b Identify which atoms have equal numbers of protons and neutrons in the nucleus.

c Identify which is an isotope of sulfur.

d Determine which has one more electron than a magnesium atom.

e Identify how many different elements are shown.

13 Determine which shell the 30th electron of an atom would go into according to the rules for determining the electronic shell configuration of an atom.

14 Determine the subshell electronic configuration of the following ions.

a Na^+

b S^{2-}

c Zn^{2+}

15 Explain how the Schrödinger model of the atom explains why the fourth electron shell begins filling before the third shell is completely filled.

16 **a** Write the subshell electronic configuration of nitrogen.

b What period and group does nitrogen belong to in the periodic table?

c How many valence electrons does nitrogen have?

d What is nitrogen's effective nuclear charge?

17 Determine the period and group of the elements with the following electronic configurations.

a $1s^22s^2$

b $1s^22s^22p^63s^23p^2$

c $1s^22s^22p^63s^23p^63d^{10}4s^24p^1$

18 Consider the elements in period 2 of the periodic table: lithium, beryllium, boron, carbon, nitrogen, oxygen, fluorine. Describe the changes that occur across the period. Consider:

a atom radius

b metallic character

c electronegativity.

19 Explain each of the following briefly.

a The atomic radius of chlorine is smaller than that of sodium.

b The first ionisation energy of fluorine is higher than that of lithium.

c The reactivity of beryllium is less than that of barium.

d There are two groups in the *s*-block of the periodic table.

20 Identify the correct chemical symbol for each of the following.

a the element that is in group 2 and period 4

b a noble gas with exactly three occupied electron shells

c an element from group 14 that is a non-metal

d an element that has exactly three occupied electron shells and is in the *s*-block

e the element in period 2 that has the largest atomic radius

f the element in group 15 that has the highest first ionisation energy

g the element in period 2 with the highest electronegativity

21 Identify the following 23 elements from the clues given, then sort them into periodic table groups.

- third element in group 15; second element in period 5; sixth element in group 17
- elements with the following symbols: C, Ca, Ga, Sb, Te, Rn
- elements with the following atomic numbers: 3,12,13,19,36,53,56
- the lightest noble gas; the third heaviest halogen; the heaviest alkali metal
- elements with the following electronic configurations:
  - $1s^22s^22p^63s^23p^4$
  - $1s^22s^22p^63s^23p^63d^{10}4s^24p^2$
  - $1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^{10}5s^25p^66s^1$
  - $1s^22s^22p^1$

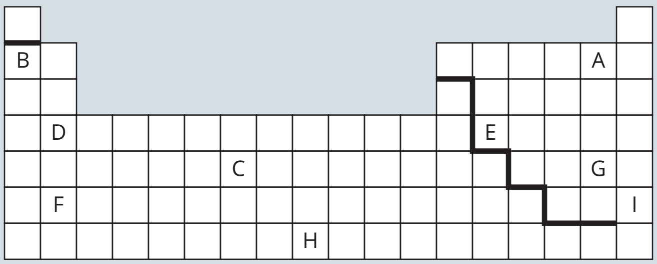
22 Compare the following pairs of elements to predict which requires more energy to remove a valence electron.

a phosphorus or magnesium

b fluorine or iodine

23 Predict which of the following pairs of elements would react together most vigorously. Explain your answer.
Li and Cl_2; Li and Br_2; K and Cl_2; K and Br_2

24 Identify the elements that meet the criteria below using the following representation of the periodic table. (Use the letters given.)



a have the same number of valence electrons

b are in the same period

c has smallest atomic radius

d has smallest ionic radius

e has lowest electronegativity

f is a noble gas

g is classified as a non-metal

h is classified as a metalloid