2 Atoms, elements and compounds

FOCUS POINTS

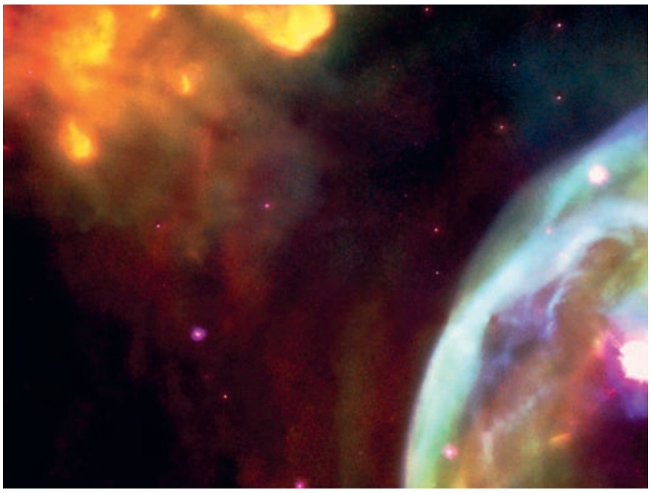
- ★ How are elements, molecules, ions, compounds and mixtures different from each other?
- ★ How do the properties of the particles in an atom lead to an atom's structure?
- ★ What do oxidation and reduction mean?
- ★ What is an isotope?

In Chapter 1, you saw that all matter is made up of particles. In this chapter you will look closely at these particles and see that they are made up of atoms. Atoms are the smallest part of elements. An element is made up of one type of atom and can be either a metal or a non-metal. Metals and non-metals have different properties.

You will look at how atoms of different elements can combine to form substances called compounds, and how this combining occurs in a chemical reaction. By the end of the chapter you should be able to write a simple word or symbol equation to represent these reactions.

You will see that although atoms are the smallest part of an element that shares the chemical properties of that element, they are made from even smaller particles. By learning about the properties and behaviour of these smaller particles (electrons, protons and neutrons), you will be able to see how they affect the chemical properties of elements and compounds.

The universe is made up of a very large number of substances (Figure 2.1), and our own part of the universe is no exception. When we examine this vast array of substances more closely, it is found that they are made up of some basic substances which were given the name **elements** in 1661 by Robert Boyle.



▲ **Figure 2.1** Structures in the universe, such as stars, planets and meteorites, are made of millions of substances. These are made up mainly from just 91 elements, all of which occur naturally on the Earth

In 1803, John Dalton suggested that each element was composed of its own kind of particles, which he called **atoms**. Atoms are much too small to be seen. We now know that about 20×10^6 of them would stretch over a length of only 1 cm.

2.1 Elements

As well as not being able to be broken down into a simpler substance, each element is made up of only one kind of atom. The word atom comes from the Greek word *atomos* meaning 'unsplittable'. For example, aluminium is an element which is made up of only aluminium atoms. It is not possible to obtain a simpler substance chemically from the aluminium atoms. You can only combine it with other elements to make more complex substances, such as aluminium oxide, aluminium nitrate or aluminium sulfate.

One hundred and eighteen elements have now been identified. Twenty of these do not occur in nature and have been made artificially by scientists. They include elements such as curium and flerovium. Ninety-eight of the elements occur naturally and range from some very reactive gases, such as fluorine and chlorine, to gold and platinum, which are unreactive elements. A physical property is any characteristic of a substance that we can measure. The elements have different properties that we can measure, and we can then classify them according to those properties.

All elements can be classified according to their various properties. A simple way to do this is to classify them as **metals** or **non-metals** (Figures 2.2 and 2.3). Table 2.1 shows the physical property data for some common metallic and non-metallic elements. You will notice from Table 2.1 that many metals have high densities, high melting points and high boiling points, and that most non-metals have low densities, low melting points and low boiling points. Table 2.2 summarises the different properties of metals and non-metals.

a Gold is very decorative

b Aluminium has many uses in the aerospace industry

c These coins contain nickel

▲ **Figure 2.2** Some metals

▼ **Table 2.1** Physical data for some metallic and non-metallic elements at room temperature and pressure

Element	Metal or non-metal	Density/ g cm^{-3}	Melting point/°C	Boiling point/°C
Aluminium	Metal	2.70	660	2580
Copper	Metal	8.92	1083	2567
Gold	Metal	19.29	1065	2807
Iron	Metal	7.87	1535	2750
Lead	Metal	11.34	328	1740
Magnesium	Metal	1.74	649	1107
Nickel	Metal	8.90	1453	2732
Silver	Metal	10.50	962	2212
Zinc	Metal	7.14	420	907
Carbon	Non-metal	2.25	Sublimes at 3642	
Hydrogen	Non-metal	0.07[a]	−259	−253
Nitrogen	Non-metal	0.88[b]	−210	−196
Oxygen	Non-metal	1.15[c]	−218	−183
Sulfur	Non-metal	2.07	113	445

Source: Earl B., Wilford L.D.R. Chemistry data book. Nelson Blackie, 1991 a: at −254°C; b: at −197°C; c: at −184°C.

The elements also have chemical properties, which are characteristics or behaviours that may be observed when the substance undergoes a chemical change or reaction. A discussion of the chemical properties of some metals and non-metals is given in Chapters 9 and 10.

▼ **Table 2.2** How the properties of metals and non-metals compare

Property	Metal	Non-metal
Physical state at room temperature	Usually solid (occasionally liquid)	Solid, liquid or gas
Malleability	Good	Poor – usually soft or brittle
Ductility	Good	
Appearance (solids)	Shiny (lustrous)	Dull
Melting point	Usually high	Usually low
Boiling point	Usually high	Usually low
Density	Usually high	Usually low
Conductivity (thermal and electrical)	Good	Very poor

Test yourself

1 Using Tables 2.1 and 2.2, pick the 'odd one out' in the following group and explain why it is different from the others.
zinc copper oxygen lead

2 Using Tables 2.1 and 2.2, pick the 'odd one out' in the following group and explain why it is different from the others.
carbon nitrogen iron sulfur

3 Using Tables 2.1 and 2.2, pick the 'odd one out' in the following group of properties of metals and explain why it is different from the others.
- high melting point
- high density
- soft or brittle
- good electrical conductivity

a A premature baby needs oxygen

b Artists often use charcoal (carbon) to produce an initial sketch

c Neon is used in advertising signs

▲ **Figure 2.3** Some non-metals

Atoms – the smallest particles

Everything is made up of billions of extremely small atoms. The smallest atom is hydrogen, and we represent each hydrogen atom as a sphere having a diameter of 0.000 000 07 mm (or 7×10^{-8} mm) (Table 2.3). Atoms of different elements have different diameters as well as different masses.

▼ **Table 2.3** Sizes of atoms

Atom	Diameter of atom/mm	Masses/g
Hydrogen	7×10^{-8}	1.67×10^{-24}
Oxygen	12×10^{-8}	2.66×10^{-23}
Sulfur	20.8×10^{-8}	5.32×10^{-23}

Chemists use shorthand symbols to label the elements and their atoms. The symbol consists of one, two or three letters, the first of which is always a capital. The initial letter of the element's name is often used and, where several elements have the same initial letter, another letter from the name is added. For example, **C** is used for **carbon**, **Ca** for **calcium** and **Cl** for **chlorine**. Some symbols seem to have no relationship to the name of the element, for example, **Na** for **sodium** and **Pb** for **lead**. These symbols come from their Latin names: **na**trium for sodium and **pl**um**b**um for lead. A list of some common elements and their symbols is given in Table 2.4.

▼ **Table 2.4** Some common elements and their symbols. The Latin names of some of the elements are given in brackets

Element	Symbol	Physical state at room temperature and pressure
Aluminium	Al	Solid
Argon	Ar	Gas
Barium	Ba	Solid
Boron	B	Solid
Bromine	Br	Liquid
Calcium	Ca	Solid
Carbon	C	Solid
Chlorine	Cl	Gas
Chromium	Cr	Solid
Copper (Cuprum)	Cu	Solid
Fluorine	F	Gas
Germanium	Ge	Solid
Gold (Aurum)	Au	Solid
Helium	He	Gas
Hydrogen	H	Gas
Iodine	I	Solid
Iron (Ferrum)	Fe	Solid
Lead (Plumbum)	Pb	Solid
Magnesium	Mg	Solid
Mercury (Hydragyrum)	Hg	Liquid
Neon	Ne	Gas
Nitrogen	N	Gas
Oxygen	O	Gas
Phosphorus	P	Solid
Potassium (Kalium)	K	Solid
Silicon	Si	Solid
Silver (Argentum)	Ag	Solid
Sodium (Natrium)	Na	Solid
Sulfur	S	Solid
Tin (Stannum)	Sn	Solid
Zinc	Zn	Solid

Molecules

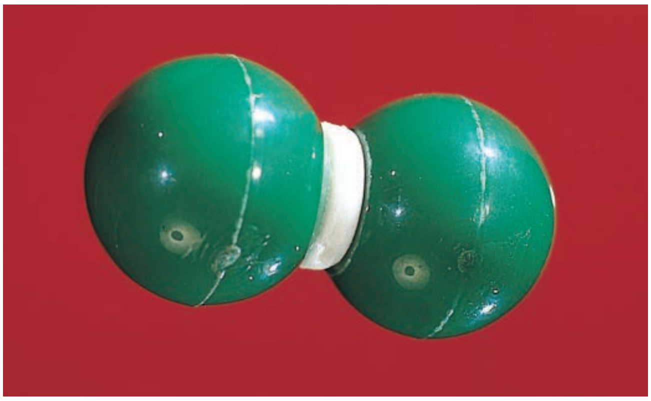
The atoms of some elements are joined together in small groups. These small groups of atoms are called **molecules**. The atoms of some elements are always joined in pairs, for example, hydrogen, oxygen, nitrogen, fluorine, chlorine, bromine and iodine. They are known as **diatomic** molecules. In chemical shorthand the molecule of chlorine shown in Figure 2.4 is written as Cl_2. The atoms of some other elements, such as phosphorus and sulfur, join in larger numbers, four and eight respectively, which we write as P_4 and S_8.

The complete list of the elements with their corresponding symbols is shown in the **Periodic Table** on p. 135.

The gaseous elements helium, neon, argon, krypton, xenon and radon (which are all gases at 0°C at sea level and atmospheric pressure) are composed of separate, individual atoms. When an element exists as separate atoms, then the molecules are said to be **monatomic**. In chemical shorthand these monatomic molecules are written as He, Ne, Ar, Kr, Xe and Rn respectively.

Cl—Cl

a Represented by a letter-and-stick model



b Represented by a space-filling model

▲ **Figure 2.4** A chlorine molecule

Molecules are not always formed by atoms of the same type joining together as elemental molecules. Most molecules consist of atoms of different elements, for example, water exists as molecules containing oxygen and hydrogen atoms. We will learn more about these in the next section.

Test yourself

4 How many atoms of hydrogen would have to be placed side by side along the edge of your ruler to fill just one of the 1 mm divisions?

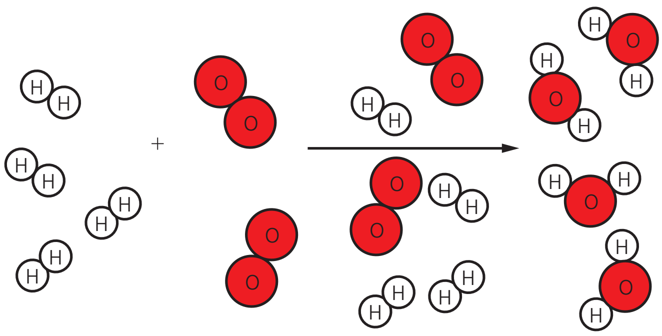
5 How would you use chemical shorthand to write a representation of the molecules of iodine and fluorine?

6 Using the Periodic Table on p. 135, write down the symbols for each of these elements and give their physical state at room temperature.
  - **a** chromium
  - **b** krypton
  - **c** osmium

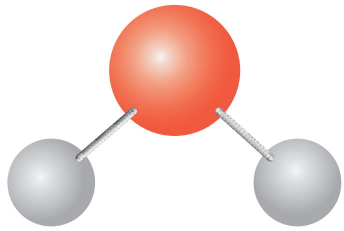
2.2 Compounds

Compounds are pure substances which are formed when two or more elements chemically combine together. A **pure substance** is a material that has a constant composition (is homogeneous) and has consistent properties throughout. Water is a simple compound formed from the elements hydrogen and oxygen (Figure 2.5). This combining of the elements can be represented by a word equation:

hydrogen +	oxygen	→	water
Hydrogen a pure element	Oxygen a pure element	Hydrogen and oxygen mixed together	Water a pure compound formed from hydrogen burning in oxygen



a The element hydrogen reacts with the element oxygen to produce the compound water



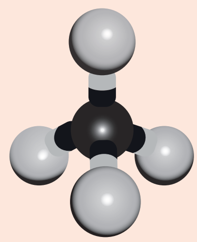
b A model of water showing 2 H atoms and one O atom. Models such as this can be built to show what a compound looks like

▲ **Figure 2.5**

Water molecules contain two atoms of hydrogen and one atom of oxygen, and water has the **chemical formula** H_2O. If there is only one atom of an element in the molecule, no number is required in the formula, only its symbol, as in the case of oxygen in the water molecule H_2O.

Test yourself

7 What is the formula for the molecule shown in the diagram which contain carbon (black sphere) and hydrogen (white spheres)?



Elements other than hydrogen will also react with oxygen to form compounds called oxides. For example, magnesium reacts violently with oxygen gas to form the white powder magnesium oxide (Figure 2.6). This reaction is accompanied by a release of energy as new chemical bonds are formed.

▲ **Figure 2.6** Magnesium burns brightly in oxygen to produce magnesium oxide

When a new substance is formed during a chemical reaction, a **chemical change** has taken place.

magnesium + oxygen → magnesium oxide

When substances such as hydrogen and magnesium combine with oxygen in this way, they are said to have been oxidised, and this process is known as **oxidation**.

Reduction is the opposite of oxidation. In this process oxygen is removed rather than added.

Key definitions

Oxidation is gain of oxygen.

Reduction is loss of oxygen.

Any chemical process that involves reduction and oxidation is known as a **redox** reaction. For example, to extract iron from iron(III) oxide, the oxygen has to be removed. The reduction of iron(III) oxide can be done in a **blast furnace** using carbon monoxide. The iron(III) oxide loses oxygen to the carbon monoxide and is reduced to iron.

Carbon monoxide is the **reducing agent**. A reducing agent is a substance that reduces another substance during a redox reaction. In the reaction, carbon monoxide is oxidised to carbon dioxide by the iron(III) oxide. In this process, the iron(III) oxide is the **oxidising agent**. An oxidising agent is a substance which oxidises another substance during a redox reaction.

We can write the redox reaction as:

iron(III) oxide + carbon monoxide → iron + carbon dioxide

Test yourself

8 Zinc is extracted from its ore zinc blende in a furnace by a redox reaction. What does the term 'redox reaction' mean?

9 Identify the oxidising and reducing agents in each of the following reactions:
   a copper(II) oxide + hydrogen → copper + water
   b tin(II) oxide + carbon → tin + carbon dioxide
   c $PbO(s) + H_2(g) \rightarrow Pb(s) + H_2O(l)$.

For a further discussion of oxidation and reduction see Chapter 3 (p. 31) and Chapter 5 (p. 71).

Key definitions

Redox reactions involve simultaneous oxidation and reduction.

An **oxidising agent** is a substance that oxidises another substance and is itself reduced.

A **reducing agent** is a substance that reduces another substance and is itself oxidised.

Practical skills

Heating copper

For safe experiments/demonstrations which are related to this chapter, please refer to the *Cambridge IGCSE Chemistry Practical Skills Workbook*, which is also part of this series.

Safety

- Eye protection must be worn.
- Take care when handling hot apparatus.
- Handle the copper with tongs or tweezers, not your fingers.

A student wants to find out what happens when copper is heated in air. In order to do this, they carried out the following experiment and recorded their results.

- First, they found the mass of an empty crucible (a suitably prepared beer-bottle top (metal) is an alternative to a porcelain crucible).
- They added a piece of copper to the crucible and found the mass again.
- They then heated the crucible strongly for approximately two minutes.
- After they allowed it to cool, they then found the mass after heating.

Mass of crucible = 12.90 g
Mass of crucible + copper = 14.18 g

Mass of copper = __________ g
Mass of crucible + contents after heating = 14.30 g

Colour of contents after heating = black

1. Draw a labelled diagram of the experimental set-up used in this experiment.
2. Calculate the change in mass that has taken place during the heating.
3. Explain what has caused the change in mass.
4. What is the black substance left on the copper after heating?
5. Write a word and balanced chemical equation to show the process that has taken place.
6. a How could you modify the experiment to ensure there was no loss of substance taking place during the heating process?
   b What are the other possible sources of error?
7. Predict what would happen, in terms of mass change and colour change, if calcium were heated in air in the same way as the copper.

Formulae

The formula of a compound is made up from the symbols of the elements that make up the compound and numbers that show the ratio of the different atoms the compound is made from. Carbon dioxide has the formula CO_2, which tells you that it contains one carbon atom for every two oxygen atoms. The 2 in the formula indicates that there are two oxygen atoms present in each molecule of carbon dioxide.

Worked example

Write the ratio of atoms in sodium sulfate – Na_2SO_4.

Substance	Formula	Ratio of atoms
Sodium sulfate	Na_2SO_4	Na : S : O 2 : 1 : 4

Test yourself

10 Write down the ratio of the atoms present in the formula for each of the compounds shown in Table 2.5.

Table 2.5 Names and formulae of some common compounds

Compound	Formula
Ammonia	NH_3
Calcium hydroxide	$Ca(OH)_2$
Carbon dioxide	CO_2
Copper sulfate	$CuSO_4$
Ethanol (alcohol)	C_2H_5OH
Glucose	$C_6H_{12}O_6$
Hydrochloric acid	HCl
Nitric acid	HNO_3
Sodium carbonate	Na_2CO_3
Sodium hydroxide	$NaOH$
Sulfuric acid	H_2SO_4

The ratio of atoms within a chemical compound is usually constant. Compounds are made up of fixed proportions of elements: they have a fixed composition. Chemists call this the **Law of constant composition**.

For further discussion of formulae, see p. 35.

Balancing chemical equations

Word equations are a useful way of representing chemical reactions, but a better and more useful way of seeing what happens during a chemical reaction is to produce a balanced chemical equation. This type of equation gives the formulae of the substances that are reacting, the reactants, and the new substances formed during the chemical reaction, the products, as well as showing the relative numbers of each of the particles involved.

Balanced equations often include symbols that show the physical state of each of the reactants and products:

(s) = solid
(l) = liquid
(g) = gas
(aq) = aqueous (water) solution

We can use the reaction between iron and sulfur as an example. The word equation to represent this reaction is:

$$\text{iron} + \text{sulfur} \xrightarrow{\text{heat}} \text{iron(II) sulfide}$$

When we replace the words with symbols for the reactants and the products, and include their physical state symbols, we get:

$$Fe(s) + S(s) \xrightarrow{\text{heat}} FeS(s)$$

Since there is the same number of each type of atom on both sides of the equation, this is a balanced chemical equation.

Worked example

Write, for the reaction between magnesium and oxygen producing magnesium oxide:

a the word equation
b the balanced chemical equation.

a The word equation is:

$$\text{magnesium} + \text{oxygen} \xrightarrow{\text{heat}} \text{magnesium oxide}$$

b When we replace the words with symbols for the reactants and the products and include their physical state symbols, it is important to remember that oxygen is a diatomic molecule:

$$Mg(s) + O_2(g) \xrightarrow{\text{heat}} MgO(s)$$

In the equation there are two oxygen atoms on the left-hand side (O_2) but only one on the right (MgO). We cannot change the formula of magnesium oxide, so to produce the necessary two oxygen atoms on the right-hand side, we will need 2MgO – this means 2 × MgO formula units. The equation now becomes:

$$Mg(s) + O_2(g) \xrightarrow{\text{heat}} 2MgO(s)$$

There are now two atoms of magnesium on the right-hand side and only one on the left. To balance the equation, we place a 2 in front of the magnesium, and obtain the following balanced chemical equation:

$$2Mg(s) + O_2(g) \xrightarrow{\text{heat}} 2MgO(s)$$

This balanced chemical equation now shows us that two atoms of magnesium react with one molecule of oxygen gas when heated to produce two units of magnesium oxide.

Test yourself

11 Write the word and balanced chemical equations for the reactions which take place between:
a calcium and oxygen
b copper and oxygen.

Going further

Instrumental techniques

Elements and compounds can be detected and identified by a variety of instrumental methods. Scientists have developed instrumental techniques that allow us to probe and discover which elements are present in a substance as well as how the atoms are arranged within the substance.

Many of the instrumental methods that have been developed are quite sophisticated. Some methods are suited to identifying elements. For example, atomic absorption spectroscopy allows the element to be identified and also allows the quantity of the element that is present to be found.

Some methods are particularly suited to the identification of compounds. For example, infrared spectroscopy is used to identify compounds by showing the presence of particular groupings of atoms.

Infrared spectroscopy is used by the pharmaceutical industry to identify and discriminate between drugs that are similar in structure, for example, penicillin-type drugs. Used both with organic and inorganic molecules, this method assumes that each compound has a unique infrared spectrum. Samples can be solid, liquid or gas and are usually tiny. However, Ne, He, O_2, N_2 or H_2 cannot be used.

This method is also used to monitor environmental pollution and has biological uses in monitoring tissue physiology including oxygenation, respiratory status and blood-flow damage.

Forensic scientists make use of both these techniques because they are very accurate, but they only require tiny amounts of sample – often only small amounts of sample are found at crime scenes. Other techniques utilised are nuclear magnetic resonance spectroscopy and ultraviolet/visible spectroscopy.

2.3 Mixtures

Many everyday things are not pure substances: they are **mixtures**. A mixture contains more than one substance, which could be elements and/or compounds. Examples of common mixtures are:

- sea water (Figure 2.7)
- air, which is a mixture of elements such as oxygen, nitrogen and neon, and compounds such as carbon dioxide (see Chapter 11, p. 178)
- alloys such as brass, which is a mixture of copper and zinc (for a further discussion of alloys, see Chapter 10, p. 165).

▲ **Figure 2.7** Sea water is a common mixture. It is a water solution of substances such as sodium chloride as well as gases such as oxygen and carbon dioxide

The difference between mixtures and compounds

There are differences between compounds and mixtures, which can be seen by looking at the reaction between iron filings and sulfur. A mixture of iron filings (powdered iron) and sulfur (Figure 2.8 bottom right), looks different from either of the individual elements (Figure 2.8 top). This mixture has the properties of both iron and sulfur; for example, a magnet can be used to separate the iron filings from the sulfur (Figure 2.9).

▲ **Figure 2.8** The elements sulfur and iron at the top of the photograph, and (below) black iron(II) sulfide on the left and a mixture of the two elements on the right

Substances in a mixture have not undergone a chemical reaction and it is possible to separate them, provided that there is a suitable difference in their physical properties. If the mixture of iron and sulfur is heated, a chemical reaction occurs and a new substance is formed. The product of the reaction is iron(II) sulfide (Figure 2.8 bottom left), and the word equation for this reaction is:

$$\text{iron} + \text{sulfur} \xrightarrow{\text{heat}} \text{iron(II) sulfide}$$

▲ **Figure 2.9** A magnet will separate the iron from the mixture

During the reaction, heat energy is given out as new chemical bonds are formed. This is called an **exothermic reaction** and accompanies a chemical change (Chapter 6, p. 92). The iron(II) sulfide formed has very different properties to the mixture of iron and sulfur (Table 2.6); for example, iron(II) sulfide would not be attracted towards a magnet. Some chemical reactions take in heat during the reaction, which is called an **endothermic reaction** (Chapter 6, p. 92). You will learn more about the different types of reactions in Chapter 6.

▼ **Table 2.6** Different properties of iron, sulfur, an iron/sulfur mixture and iron(II) sulfide

Substance	Appearance	Effect of a magnet	Effect of dilute hydrochloric acid
Iron	Dark grey powder	Attracted to it	Very little action when cold. When warm, a gas is produced with a lot of bubbling (effervescence)
Sulfur	Yellow powder	None	No effect when hot or cold
Iron/sulfur mixture	Dirty yellow powder	Iron powder attracted to it	Iron powder reacts as above
Iron(II) sulfide	Black solid	No effect	A foul-smelling gas is produced with some effervescence

In iron(II) sulfide, FeS, one atom of iron has combined with one atom of sulfur. In a mixture of iron and sulfur, no such ratio exists as the atoms have not chemically combined. Table 2.7 compares mixtures and compounds. Some common mixtures are discussed in Chapter 10 (p. 165) and Chapter 11 (p. 177).

▼ **Table 2.7** The major differences between mixtures and compounds

Mixture	Compound
It contains two or more substances.	It is a single substance.
The composition can vary.	The composition is always the same.
No chemical change takes place when a mixture is formed.	When the new substance is formed it involves chemical change.
The properties are those of the individual elements/compounds.	The properties are very different to those of the component elements.
The components may be separated quite easily by physical means.	The components can only be separated by one or more chemical reactions.

Test yourself

12 Make a list of some other common mixtures and then use your research to find out and state what they are mixtures of.

13 Which of the following are not mixtures: milk, tin, sulfur, cough linctus, brass, gold?

Going further

Other types of mixtures

There are mixtures which are formed by mixing two substances (or phases) which cannot mix. Gels, sols, foams and emulsions are all examples of just such mixtures. Look closely at the substances in Figure 2.10, which shows examples of these different types of mixture.

a

b

c

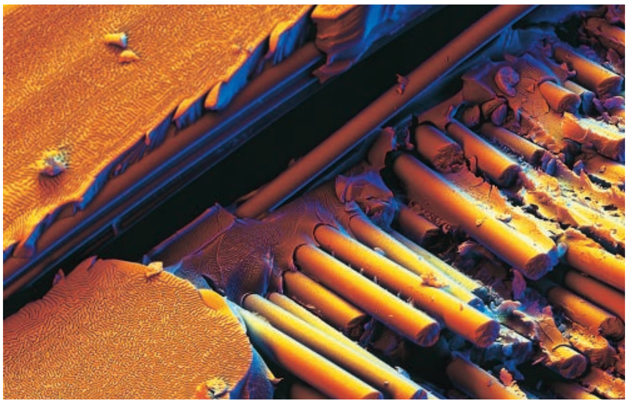
d

a This jelly is an example of a 'gel'
b Emulsion paint is an example of a 'sol'
c These foams have been formed by trapping bubbles of gas in liquids
d Emulsions are formed by mixing immiscible liquids

▲ **Figure 2.10**

Composite materials are those that combine the properties of two constituents in order to get the exact properties needed for a particular job. Glass-reinforced fibre is an example of a composite material combining the properties of two different materials. It is made by embedding short fibres of glass in a matrix of plastic.

The glass fibres give the plastic extra strength so that it does not break when it is bent or moulded into shape. The finished material has the strength and flexibility of the glass fibres as well as the lightness of plastic (Figure 2.11).



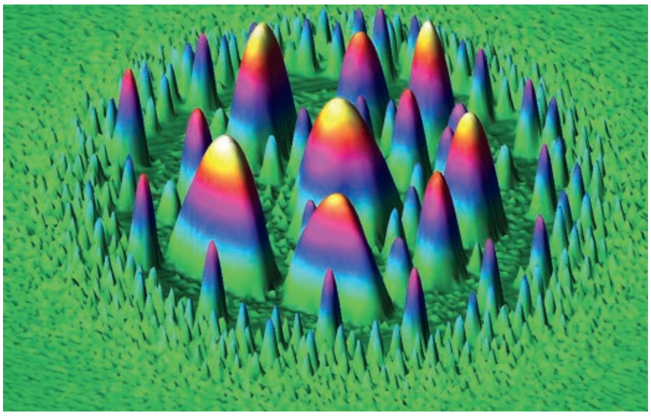
a Glass-reinforced plastic (GRP) consists of glass fibres (rod shapes) embedded in plastic, in this case polyester

b The glass-reinforced plastic used to make boats like this is a composite material

▲ **Figure 2.11**

2.4 Inside atoms

Everything you see around you is made out of tiny particles, which we call atoms (Figure 2.12). When John Dalton developed his atomic theory, over 200 years ago, he stated that the atoms of any one element were identical and that each atom was 'indivisible'. Scientists in those days also believed that atoms were solid particles, like marbles.

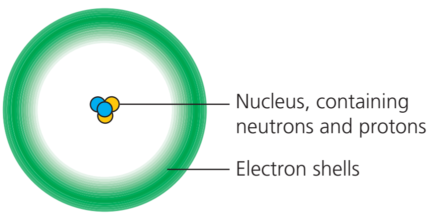


▲ **Figure 2.12** An electron micrograph of atoms taken using a very powerful microscope called an electron microscope

However, in the last hundred years or so it has been proved by great scientists, such as Niels Bohr, Albert Einstein, Henry Moseley, Joseph Thomson, Ernest Rutherford and James Chadwick, that atoms are in fact made up of even smaller 'sub-atomic' particles. Seventy sub-atomic particles have now been discovered, and the most important of these are **electrons**, **protons** and **neutrons**.

These three sub-atomic particles are found in distinct and separate regions of the atom. The protons and neutrons are found in the centre of the atom, which is called the **nucleus**. Neutrons have no charge and protons are positively charged. The nucleus occupies only a very small volume of the atom and is very dense.

The rest of the atom surrounding the nucleus is where electrons are found. Electrons are negatively charged and move around the nucleus very quickly at specific distances from the nucleus in **electron shells** or **energy levels**. The electrons are held in the shells within the atom by an **electrostatic force of attraction** between themselves and the positive charge of protons in the nucleus (Figure 2.13). Each shell can contain only a fixed number of electrons: the first shell can hold up to two electrons, the second shell can hold up to eight electrons, the third shell can hold up to 18, and so on.



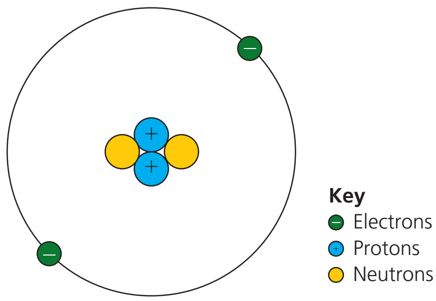
▲ **Figure 2.13** Diagram of an atom

Electrons are tiny and relatively light; approximately 1837 electrons are equal in mass to the mass of one proton or one neutron. A summary of each type of particle, its mass and relative charge is shown in Table 2.8. You will notice that the masses of all these particles are measured in atomic mass units (amu). This is because they are so light that their masses cannot be measured usefully in grams.

▼ **Table 2.8** Characteristics of a proton, a neutron and an electron

Particle	Symbol	Relative mass/ amu	Relative charge
Proton	p	1	+1
Neutron	n	1	0
Electron	e	1/1837	–1

Although atoms contain electrically charged particles, the atoms themselves are electrically neutral (they have no overall electric charge). This is because atoms contain equal numbers of electrons and protons. For example, Figure 2.14 represents the atom of the non-metallic element helium. The atom of helium possesses two protons, two neutrons and two electrons. The electrical charge of the protons in the nucleus is, therefore, balanced by the opposite charge of the two electrons.



▲ **Figure 2.14** An atom of helium has two protons, two electrons and two neutrons

Proton number and mass number

The number of protons in the nucleus of an atom is called the **proton number** (or atomic number) and is given the symbol Z. The helium atom in Figure 2.14, for example, has a proton number of 2, since it has two protons in its nucleus. Each element has its own proton number and no two elements have the same proton number. For example, the element lithium has a proton number of 3 since it has three protons in its nucleus.

Neutrons and protons have a similar mass, whereas electrons possess very little mass. So the mass of any atom depends on the number of protons and neutrons in its nucleus. The total number of protons and neutrons found in the nucleus of an atom is called the **mass number** (or nucleon number) and is given the symbol *A*.

Key definitions

Proton number or atomic number is the number of protons in the nucleus of an atom.

Mass number or nucleon number is the total number of protons and neutrons in the nucleus of an atom.

$$\underset{(A)}{\text{mass number}} = \underset{(Z)}{\text{proton number}} + \text{number of neutrons}$$

The helium atom in Figure 2.14 has a mass number of 4, since it has two protons and two neutrons in its nucleus. If we consider the metallic element lithium, it has three protons and four neutrons in its nucleus. It therefore has a mass number of 7.

The proton number and mass number of an element are usually written in the following shorthand way:

$$\begin{array}{l} \text{Mass number } (A) \searrow \\ \qquad\qquad\qquad\qquad {}^{4}_{2}\text{He} \leftarrow \text{Symbol of the element} \\ \text{Proton number } (Z) \nearrow \end{array}$$

The number of neutrons present can be calculated by rearranging the relationship between the proton number, mass number and number of neutrons to give:

$$\text{number of neutrons} = \underset{(A)}{\text{mass number}} - \underset{(Z)}{\text{proton number}}$$

Worked example

1 What is the number of neutrons in one atom of ${}^{24}_{12}Mg$?

number of neutrons = mass number − proton number

$12 = 24(A) - 12(Z)$

2 What is the number of neutrons in one atom of ${}^{207}_{82}Pb$?

number of neutrons = mass number − proton number

$125 = 207(A) - 82(Z)$

Table 2.9 shows the number of protons, neutrons and electrons in the atoms of some common elements.

▼ **Table 2.9** Number of protons, neutrons and electrons in some elements

Element	Symbol	Proton number	Number of electrons	Number of protons	Number of neutrons	Mass number
Hydrogen	H	1	1	1	0	1
Helium	He	2	2	2	2	4
Carbon	C	6	6	6	6	12
Nitrogen	N	7	7	7	7	14
Oxygen	O	8	8	8	8	16
Fluorine	F	9	9	9	10	19
Neon	Ne	10	10	10	10	20
Sodium	Na	11	11	11	12	23
Magnesium	Mg	12	12	12	12	24
Sulfur	S	16	16	16	16	32
Potassium	K	19	19	19	20	39
Calcium	Ca	20	20	20	20	40
Iron	Fe	26	26	26	30	56
Zinc	Zn	30	30	30	35	65

Ions

An **ion** is an electrically charged particle. When an atom loses one or more electrons, it is no longer electrically neutral and becomes a positively charged ion. This is called a cation.

For example, when potassium is involved in a chemical reaction, each atom loses an electron to form a positive ion, K^+.

$$
{}_{19}K^+ \quad
\begin{array}{rl}
\text{19 protons} = & 19+ \\
\underline{\text{18 electrons} =} & \underline{18-} \\
\text{Overall charge} = & 1+
\end{array}
$$

When an atom gains one or more electrons, it becomes a negatively charged ion. This is called an anion. For example, in some of the chemical reactions involving oxygen, each oxygen atom gains two electrons to form a negative ion, O^{2-}.

$$
{}_{8}O^{2-} \quad
\begin{array}{rl}
\text{8 protons} = & 8+ \\
\underline{\text{10 electrons} =} & \underline{10-} \\
\text{Overall charge} = & 2-
\end{array}
$$

The process of gaining or losing electrons is known as **ionisation**.

Table 2.10 shows some common ions. You will notice that:

- some ions contain more than one type of atom, for example NO_3^-
- an ion may possess more than one unit of charge (either negative or positive), for example, Al^{3+}, O^{2-} and SO_4^{2-}.

Worked example

1 Show how the sodium ions shown in Table 2.10 are formed.

$$
{}_{11}Na^+ \quad
\begin{array}{rl}
\text{11 protons} = & 11+ \\
\underline{\text{10 electrons} =} & \underline{10-} \\
\text{Overall charge} = & 1+
\end{array}
$$

2 Show how the sulfide ions shown in Table 2.10 are formed.

$$
{}_{16}S^{2-} \quad
\begin{array}{rl}
\text{16 protons} = & 16+ \\
\underline{\text{18 electrons} =} & \underline{18-} \\
\text{Overall charge} = & 2-
\end{array}
$$

Table 2.10 Some common ions

Name	Formula
Lithium ion	Li^+
Sodium ion	Na^+
Potassium ion	K^+
Magnesium ion	Mg^{2+}
Calcium ion	Ca^{2+}
Aluminium ion	Al^{3+}
Zinc ion	Zn^{2+}
Ammonium ion	NH_4^+
Fluoride ion	F^-
Chloride ion	Cl^-
Bromide ion	Br^-
Hydroxide ion	OH^-
Oxide ion	O^{2-}
Sulfide ion	S^{2-}
Carbonate ion	CO_3^{2-}
Nitrate ion	NO_3^-
Sulfate ion	SO_4^{2-}

Test yourself

14 a State three differences between an electron and a proton.

b State two similarities between a neutron and a proton.

c Why are atoms electrically neutral?

15 Copy and complete the following table by writing the symbol for the element shown from the given number of particles present. The first one is done for you.

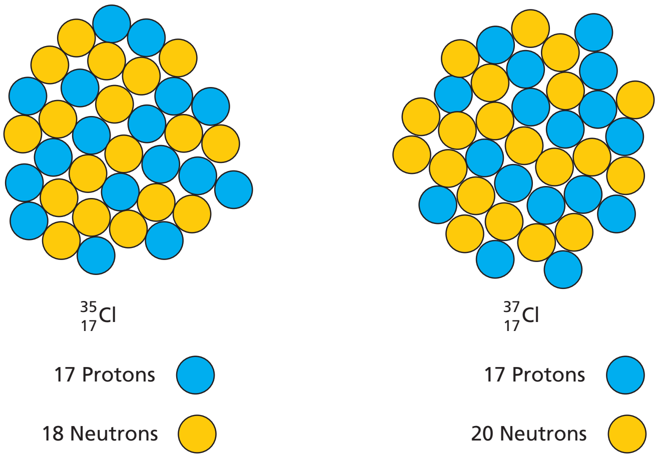
Element	Symbol	Particles present
Nitrogen	${}^{14}_{7}N$	7p, 7n, 7e
Aluminium		13p, 14n, 13e
Potassium		19p, 20n, 19e
Argon		18p, 22n, 18e

16 Copy and complete the following table giving the charge on the ions shown.

Element	Proton number	Number of electrons	Charge on ion
Sodium	11	10	
Fluorine	9	10	
Magnesium	12	10	

Isotopes

In some elements, not all of the atoms in a sample of the element are identical. Some atoms of the same element can contain different numbers of neutrons and so have different mass numbers. Atoms of the same element which have different neutron numbers are called **isotopes**. The two isotopes of chlorine are shown in Figure 2.15.



▲ **Figure 2.15** The nuclei of two isotopes of chlorine

Key definition

Isotopes are atoms of the same element that have the same number of protons but different numbers of neutrons.

Isotopes of the same element have the same chemical properties because they have the same number of electrons and therefore the same electronic configuration (p. 26).

The only effect of the extra neutron is to alter the mass of the atom and any properties that depend on it, such as density. Other examples of atoms and their isotopes are shown in Table 2.11.

There are two types of isotopes: those which are stable and those which are unstable. The isotopes which are unstable, as a result of the extra neutrons in their nuclei, are **radioactive** and are called **radioisotopes**. The nuclei of these atoms break up spontaneously with the release of not only large amounts of energy, but also certain types of dangerous radiations that can in some cases be useful to society. For example, uranium-235 is used as a source of power in nuclear reactors in nuclear power stations and cobalt-60 is used in radiotherapy treatment in hospitals (Figure 2.16); these are both radioisotopes.

▼ **Table 2.11** Some atoms and their isotopes

Element	Symbol	Particles present
Hydrogen	$^{1}_{1}H$	1e, 1p, 0n
(Deuterium)	$^{2}_{1}H$	1e, 1p, 1n
(Tritium)	$^{3}_{1}H$	1e, 1p, 2n
Carbon	$^{12}_{6}C$	6e, 6p, 6n
	$^{13}_{6}C$	6e, 6p, 7n
	$^{14}_{6}C$	6e, 6p, 8n
Oxygen	$^{16}_{8}O$	8e, 8p, 8n
	$^{17}_{8}O$	8e, 8p, 9n
	$^{18}_{8}O$	8e, 8p, 10n
Strontium	$^{86}_{38}Sr$	38e, 38p, 48n
	$^{88}_{38}Sr$	38e, 38p, 50n
	$^{90}_{38}Sr$	38e, 38p, 52n
Uranium	$^{235}_{92}U$	92e, 92p, 143n
	$^{238}_{92}U$	92e, 92p, 146n

▲ **Figure 2.16** Cobalt-60 is used in radiotherapy treatment. A beam of gamma rays produced by the radioactive isotope is directed into the patient's body to kill tumour tissue.

Going further

The mass spectrometer

How do we know isotopes exist? They were first discovered by scientists using apparatus called a mass spectrometer (Figure 2.17). The first mass spectrometer was built by the British scientist Francis Aston in 1919 and enabled scientists to compare the relative masses of atoms accurately for the first time.

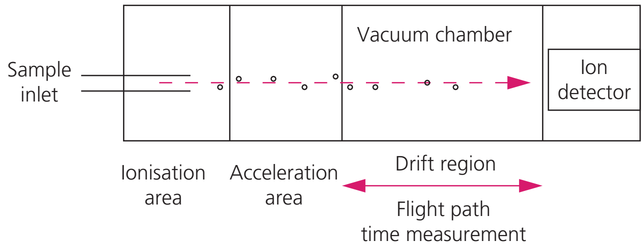


Figure 2.17 A diagram of a mass spectrometer

A vacuum exists inside the mass spectrometer. A sample of the vapour of the element is injected into the ionisation chamber where it is bombarded by electrons. The collisions which take place between these electrons and the injected atoms cause an electron to be lost from the atom, which becomes a positive ion with a +1 charge. These positive ions are then accelerated towards a negatively charged plate, in the acceleration area.

The spectrometer is set up to ensure that when the ions leave the acceleration area they all have the same kinetic energy, regardless of the mass of the ions. This means that the lighter ions travel faster than the heavier ones, and effectively separate the ions according to their mass. Having left the acceleration area, the time for the ions to reach the detector is recorded. The detector counts the number of each of the ions which fall upon it and so a measure of the percentage abundance of each isotope is obtained. A typical mass spectrum for chlorine is shown in Figure 2.18.

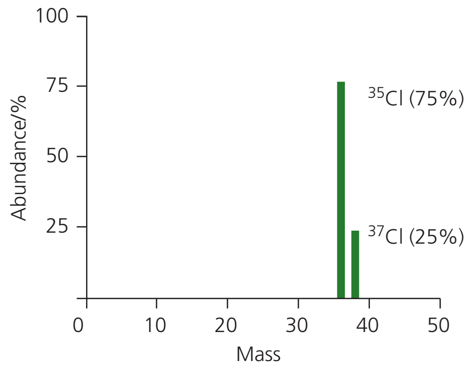


Figure 2.18 The mass spectrum for chlorine

Relative atomic mass

The average mass of a large number of atoms of an element is called its **relative atomic mass (A_r)**. This quantity takes into account the percentage abundance of all the isotopes of an element which exist.

Key definition

Relative atomic mass, A_r, is the average mass of the isotopes of an element compared to 1/12th of the mass of an atom of ^{12}C.

In 1961 the International Union of Pure and Applied Chemistry (IUPAC) recommended that the standard used for the A_r scale was carbon-12. An atom of carbon-12 was taken to have a mass of 12 amu. The A_r of an element is the average mass of the naturally occurring atoms of an element on a scale where ^{12}C has a mass of exactly 12 units.

$$A_r = \frac{\text{average mass of isotopes of the element}}{\frac{1}{12} \times \text{mass of one atom of carbon-12}}$$

Note: $\frac{1}{12}$ of the mass of one carbon-12 atom = 1 amu.

Worked example

What is the relative atomic mass of chlorine? Chlorine has two isotopes:

	$^{35}_{17}Cl$	$^{37}_{17}Cl$
% abundance	75	25

Hence the 'average mass' or A_r of a chlorine atom is:

$$\frac{(75 \times 35) + (25 \times 37)}{100} = 35.5$$

$$A_r = \frac{35.5}{1}$$

$$= 35.5 \text{ amu}$$

Test yourself

17 Calculate the number of neutrons in the following atoms:
a $^{27}_{13}Al$ **b** $^{31}_{15}P$ **c** $^{91}_{40}Zr$ **d** $^{190}_{76}Os$.

18 Given that the percentage abundance of $^{20}_{10}Ne$ is 90% and that of $^{22}_{10}Ne$ is 10%, calculate the A_r of neon.

The arrangement of electrons in atoms

The nucleus of an atom contains the heavier sub-atomic particles – the protons and the neutrons. The electrons, the lightest of the sub-atomic particles, move around the nucleus at great distances from the nucleus relative to their size. They move very fast in electron shells, very much like the planets orbit the Sun.

It is not possible to give the exact position of an electron in an electron shell. However, we can state that electrons can only occupy certain, definite electron shells and that they cannot exist between them. Also, as mentioned earlier, each of the electron shells can hold only a certain number of electrons.

- First shell holds up to two electrons.
- Second shell holds up to eight electrons.
- Third shell holds up to 18 electrons.

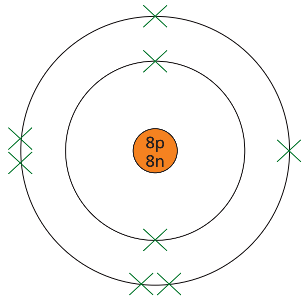
There are further electron shells which contain increasing numbers of electrons.

The third electron shell can be occupied by a maximum of 18 electrons. However, when eight electrons have occupied this shell, a certain stability is given to the atom and the next two electrons go into the fourth electron shell, and then the remaining ten electrons complete the third shell.

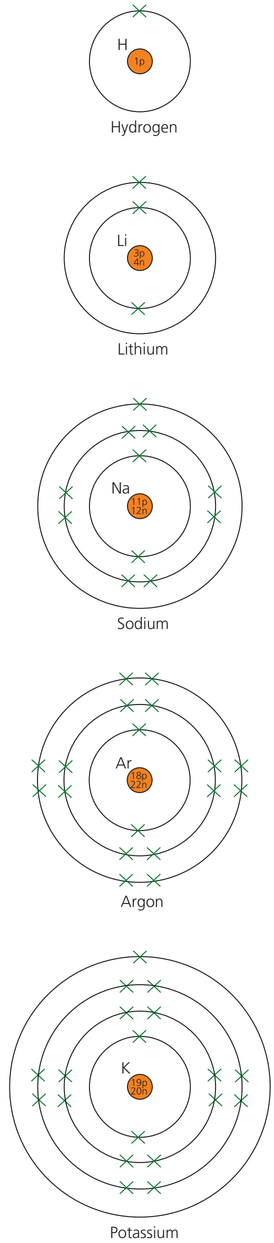
The electrons fill the electron shell starting from the shell nearest to the nucleus, which has the lowest energy. When this is full (with two electrons), the next electron goes into the second electron shell. When this shell is full with eight electrons, then the electrons begin to fill the third and fourth shells as stated above.

For example, a $^{16}_{8}O$ atom has a proton number of 8 and therefore has eight electrons. Two of the eight electrons enter the first electron shell, leaving six to occupy the second electron shell, as shown in Figure 2.19. The electronic configuration for oxygen can be written in a shorthand way as 2,6.

There are 118 elements, and Table 2.12 shows the way in which the electrons are arranged in the first 20 of these elements. The way in which the electrons are distributed is called the **electronic configuration** (electronic structure). Figure 2.20 shows the electronic configuration of a selection of atoms.



▲ **Figure 2.19** Arrangement of electrons in an oxygen atom



▲ **Figure 2.20** Electronic configurations of hydrogen, lithium, sodium, argon and potassium

▼ **Table 2.12** Electronic configuration of the first 20 elements

Element	Symbol	Proton number	Number of electrons	Electronic configuration
Hydrogen	H	1	1	1
Helium	He	2	2	2
Lithium	Li	3	3	2,1
Beryllium	Be	4	4	2,2
Boron	B	5	5	2,3
Carbon	C	6	6	2,4
Nitrogen	N	7	7	2,5
Oxygen	O	8	8	2,6
Fluorine	F	9	9	2,7
Neon	Ne	10	10	2,8
Sodium	Na	11	11	2,8,1
Magnesium	Mg	12	12	2,8,2
Aluminium	Al	13	13	2,8,3
Silicon	Si	14	14	2,8,4
Phosphorus	P	15	15	2,8,5
Sulfur	S	16	16	2,8,6
Chlorine	Cl	17	17	2,8,7
Argon	Ar	18	18	2,8,8
Potassium	K	19	19	2,8,8,1
Calcium	Ca	20	20	2,8,8,2

From Table 2.12 you can see helium has a full outer shell of two electrons, and neon has a full outer shell of 8 electrons. These two elements are very unreactive for this reason. In Chapter 9 we will see that the number of outer shell electrons is related to the position in the Periodic Table. Helium and neon are part of a group of elements known as the noble or inert gases and they are generally very stable and unreactive (p. 145). This is linked to their full outer shells. When elements react to form compounds, they do so to achieve full electron shells, and this idea forms the basis of the electronic theory of chemical bonding, which we will discuss further in the next chapter.

Test yourself

19 How many electrons may be accommodated in each of the first three electron shells?

20 What is the same about the electronic configurations of:

- **a** lithium, sodium and potassium?
- **b** beryllium, magnesium and calcium?

21 An element X has a proton number of 13. What is the electronic configuration of X?

Revision checklist

After studying Chapter 2 you should be able to:

- ✔ Describe the differences between elements, compounds and mixtures.
- ✔ Interpret and use symbols for given atoms.
- ✔ State the formulae of the elements and compounds you have dealt with.
- ✔ Define the molecular formula of a compound as the number and type of different atoms in one molecule.
- ✔ Deduce the formula of a simple compound from the relative numbers of atoms present in a model or a diagrammatic representation of the compound.
- ✔ Construct word equation and symbol equations to show how reactants form products, including state symbols.
- ✔ Define redox reactions as involving both oxidation and reduction.
- ✔ Define oxidation as oxygen gain and reduction as oxygen loss.
- ✔ Identify redox reactions as reactions involving gain and loss of oxygen.

- ✔ Define an oxidising agent and a reducing agent.

- ✔ Describe the structure of the atom as a central nucleus containing neutrons and protons surrounded by electrons in shells.
- ✔ State the relative charges and relative masses of a proton, a neutron.
- ✔ Define proton number and atomic number as well as mass number.
- ✔ Determine the electronic configuration of elements with the proton number 1 to 20.
- ✔ Describe the formation of positive ions, known as cations, and negative ions, known as anions.
- ✔ State what isotopes are.

- ✔ State that isotopes of the same element have the same electronic configuration and so have the same chemical properties.
- ✔ Calculate the relative atomic mass (A_r) of an element from given data of the relative masses and abundance of their different isotopes.

Exam-style questions

1 a Define the terms:
- **i** proton [3]
- **ii** neutron [2]
- **iii** electron. [3]

b An atom **X** has a proton number of 19 and relative atomic mass of 39.
- **i** How many electrons, protons and neutrons are there in atom **X**? [3]
- **ii** How many electrons will there be in the outer electron shell of atom **X**? [1]
- **iii** What is the electronic configuration of atom **X**? [1]

2 a $^{69}_{31}Ga$ and $^{71}_{31}Ga$ are isotopes of gallium. Use this example to explain what you understand by the term isotope. [3]

b A sample of gallium is 60% $^{69}_{31}Ga$ atoms and 40% $^{71}_{31}Ga$ atoms. Calculate the relative atomic mass of this sample of gallium. [2]

3 Define the following terms using specific examples to help with your explanation:
- **a** element [2]
- **b** metal [3]
- **c** non-metal [2]
- **d** compound [2]
- **e** molecule [2]
- **f** mixture. [2]

4 State which of the substances listed below are:
- **a** metallic elements
- **b** non-metallic elements
- **c** compounds
- **d** mixtures.

silicon, sea water, calcium, argon, water, air, carbon monoxide, iron, sodium chloride, diamond, brass, copper, dilute sulfuric acid, sulfur, oil, nitrogen, ammonia [17]

5 State, at room temperature and pressure (r.t.p.), which of the substances listed below is/are:
- **a** a solid element
- **b** a liquid element
- **c** a gaseous mixture
- **d** a solid mixture
- **e** a liquid compound
- **f** a solid compound.

bromine, carbon dioxide, helium, steel, air, oil, marble, copper, water, sand, tin, bronze, mercury, salt [11]

6 a How many atoms of the different elements are there in the formulae of these compounds?
- **i** nitric acid, HNO_3 [3]
- **ii** methane, CH_4 [2]
- **iii** copper nitrate, $Cu(NO_3)_2$ [3]
- **iv** ethanoic acid, CH_3COOH [3]
- **v** sugar, $C_{12}H_{22}O_{11}$ [3]
- **vi** phenol, C_6H_5OH [3]
- **vii** ammonium sulfate, $(NH_4)_2SO_4$ [4]

b Balance the following equations.
- **i** $Zn(s) + O_2(g) \rightarrow ZnO(s)$ [2]
- **ii** $Fe(s) + Cl_2(g) \rightarrow FeCl_3(s)$ [3]
- **iii** $Li(s) + O_2(g) \rightarrow Li_2O(s)$ [2]
- **iv** $H_2(g) + O_2(g) \rightarrow H_2O(g)$ [2]
- **v** $Mg(s) + CO_2(g) \rightarrow MgO(s) + C(s)$ [2]