

MODULE 1

the chemical
earth

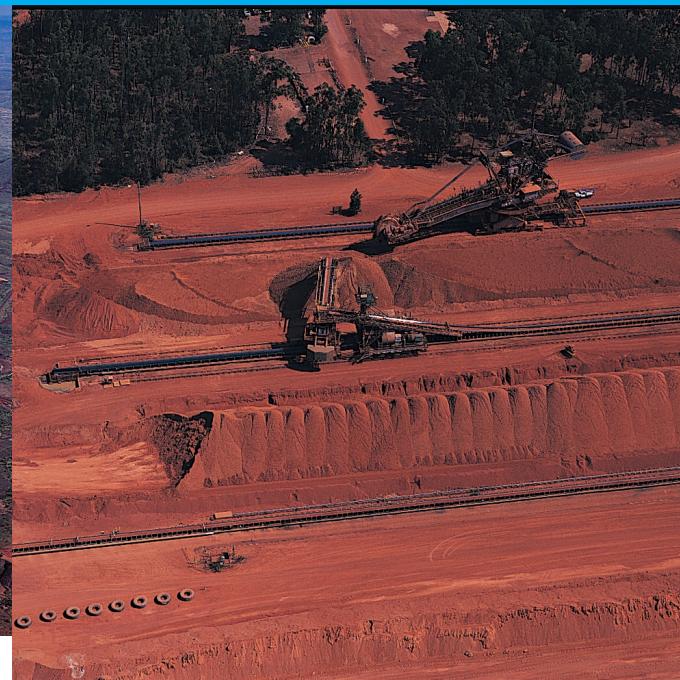
The land, seas and atmosphere of planet Earth provide us humans with all the substances (or resources) we need for our survival, comfort and pleasure. Some of these resources we get from living matter—such as food and clothing from plants and animals, and timber for shelter from trees—but many of the resources we use come from the non-living Earth itself. This module will focus mainly upon the non-living resources.

Some of the common substances (resources) we get from the Earth are:

- coal, oil and natural gas for our energy needs
- metals such as iron, aluminium and copper for making structures, machines and containers
- gypsum for making plaster for houses
- rock phosphate and sulfur for making fertilisers
- sand and limestone for making glass and cement
- silicon for making computer chips and solar cells
- gold, diamonds and other gemstones for personal decoration
- argon from the atmosphere for filling light bulbs.

The photos below show some of these substances.

The part of the earth we live in, the biosphere, consists of mixtures of many substances. Extracting the substances we want involves separating mixtures—extracting the required substance from the unwanted material. Often we need the substances we use to be pure, for example copper for electrical wiring and silicon for



making solar cells or computer chips. Hence we need to be able to purify substances and use properties of the substances to determine whether or not they are pure.

Because there are so many pure substances in the Earth, we frequently *classify* them into various categories. We can classify them into metals or non-metals, or we can group them into families of substances with similar properties (in what we call the Periodic Table).

In this module then, we shall consider the ways in which people in general and chemists in particular separate mixtures into pure substances, and look at the common properties and uses of some of these pure substances. We shall survey the range of substances (elements and compounds) that make up the Earth and discuss some methods of classifying or organising them in Chapter 1.

Most of the substances of the Earth are compounds rather than free elements. We shall therefore explore why and how elements combine to form compounds. This will require a look at the structure of atoms and a treatment of what is called chemical bonding, which is covered in Chapter 2.

Many of the substances we want are not the pure compounds we extract from the Earth but rather substances we make from those raw materials—iron from haematite, aluminium from aluminium oxide, chlorine from salt. We get these wanted substances by what are called chemical reactions, introduced in Chapter 3. Because there are so many compounds in the Earth we need a system of naming them which tells us exactly what the compound is—in particular which tells us its chemical formula. Chapter 3 will also introduce systematic naming.



HARVESTING THE RESOURCES OF THE EARTH

FROM LEFT: iron ore, bauxite for the production of aluminium, salt from sea water and sulfur for making sulfuric acid.



Mixtures, elements and compounds in the Earth

IN THIS CHAPTER

Mixtures and pure substances

Elements and compounds

The ‘spheres’ of planet Earth

The Earth as mixtures of substances

Separation of mixtures of:

— solids of different sizes

— solids and liquids (soluble and miscible)

— miscible liquids

— immiscible liquids

— gases

Properties for identifying substances:

— colour and physical state

— melting and boiling points

— density

Gravimetric analysis

Elements occurring as free elements

Why most elements occur as compounds

Metals and non-metals

Physical properties and uses of metals

The Periodic Table

In this module we will be looking at various mixtures, elements and compounds that make up planet Earth. The discussion will use several basic concepts of chemistry that should be familiar to you from your earlier studies of science. However, we shall review these basic concepts as you need them, so let us start by revising mixtures, elements and compounds.

1.1 MIXTURES AND PURE SUBSTANCES

Table 1.1 explains the differences between a mixture and a pure substance.

Perhaps the meaning of two words used in it should be explained.

Homogeneous means of uniform composition throughout, for example pure water, sugar, aluminium, petrol or whisky. **Heterogeneous** means having non-uniform composition where we can recognise small pieces of the material which are different from other pieces, for example strawberry jam, wood, or water with ice in it.

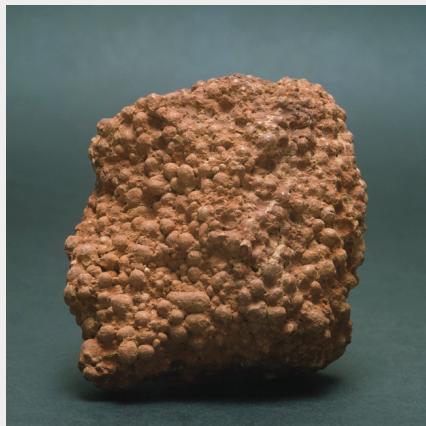
Impure substance

An **impure substance** is one substance contaminated with small amounts of one or more other substances.

An impure substance is therefore a **mixture**.

TABLE 1.1 Differences between a mixture and a pure substance

A mixture	A pure substance
can be separated into two or more pure substances by physical or mechanical means such as filtering, boiling or using a magnet or tweezers	cannot be separated into two or more substances by physical or mechanical means
may be homogeneous (tap water, air) or heterogeneous (fruit cake, concrete)	is homogeneous (crystals of sugar, piece of copper)
displays the properties (characteristics) of the pure substances making it up (different parts of the mixture show different properties)	has properties (characteristics) such as appearance, colour, density, melting and boiling points, which are constant throughout the whole sample
has properties that can change as the relative amounts of the substances present are changed	has properties that do not change regardless of how it is prepared or how many times it is subjected to purification procedures
has a variable composition; that is, the relative amounts of each pure substance present can be varied	has a fixed composition, no matter how it is made or where it comes from
Examples: sea water, air, coffee, milk, petrol, whisky, brass, and 'silver' coins	Examples: table salt, sugar, copper, aluminium, diamond, gold, polyethylene and alcohol



The substances described in Table 1.2: bauxite (left), a mixture, aluminium oxide (centre), a compound separated from it, and aluminium (right), an element obtained from the compound

1.2 ELEMENTS AND COMPOUNDS

Some **pure substances** can be decomposed into simpler substances while others cannot. Those that cannot be decomposed are called elements; those that can be decomposed are called compounds.

An **element** is a pure substance which cannot be decomposed into simpler substances.

Some common elements are aluminium, copper, carbon (diamond), oxygen, gold, nitrogen and mercury.

A **compound** is a pure substance which can be decomposed into simpler substances, for example into elements.

Some common compounds are table salt (sodium chloride), sugar (sucrose), water, sodium carbonate (washing soda), sulfate of ammonia (ammonium sulfate, a common fertiliser), alcohol (ethanol) and aspirin (painkiller).

A compound:

- is made up of two or more elements
- always has the elements present in the same ratio by mass
- has properties that are quite different from those of the elements that make it up

Table 1.2 presents a specific example that highlights the differences between mixtures, compounds and elements.

TABLE 1.2 Contrasting properties of bauxite, aluminium oxide and aluminium

Bauxite (mixture)	Aluminium oxide (compound)	Aluminium (element)
red pebbly solid	crystalline white solid	silvery lustrous solid
no definite melting point	melting point is 2045°C	melting point is 660°C
composition varies from mine to mine	constant composition by mass (52.9% aluminium)	
can be <i>separated</i> into aluminium oxide, iron(III) oxide and dirt	can be <i>decomposed</i> (by electrolysis) into aluminium and oxygen	cannot be decomposed into simpler substances
density varies with composition	density is 4.0 g/mL	density is 2.7 g/mL
fairly easily ground into a fine powder	the small crystals are hard and brittle	fairly soft, but malleable and ductile

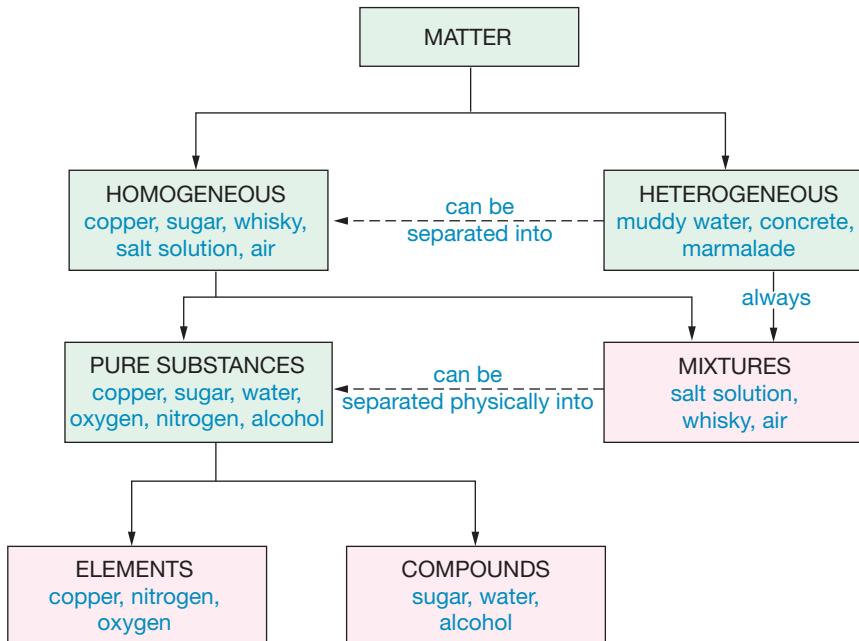


The elements copper (centre), sulfur (yellow, directly below copper) and carbon (black, anticlockwise from sulfur), and some compounds of copper: (anticlockwise from carbon) copper sulfate, copper carbonate, copper(I) oxide, copper sulfide and copper chloride

There are about 90 naturally occurring elements, and these are the basic building materials of all matter. In addition, scientists have made a further dozen or so artificial elements (which have quite short lifetimes). From this mere 90 or so elements, nature and scientists have made millions of different compounds.

FIGURE 1.1
A classification of matter

Figure 1.1 shows the classification of matter into mixtures, elements and compounds.



An example will help clarify these ideas.

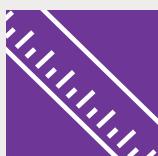
When iron filings and powdered yellow sulfur are mixed in a small basin, we can still identify the individual substances. We can vary the relative amounts of the two substances present and separate them by using a magnet to remove the iron filings. This is a *mixture*. However if this mixture is placed in a test tube and heated strongly, a black substance forms. Any excess sulfur burns off (as sulfur dioxide). The black solid is homogeneous and is not affected by a magnet. Its properties are completely different from those of sulfur and iron. It has a fixed composition: regardless of the actual proportions of sulfur and iron in the original mixture, the black solid always contains 36% sulfur. This black solid is a *compound*. Heating has brought about a chemical reaction in which iron and sulfur have reacted to produce a new substance, the compound iron sulfide.



WEBSITE

<http://www.chem4kids.com>

(select *matter* and *elements*; an attractively presented website, treats the material very simply and quite informally)



Exercises

1 Classify the following as *homogeneous* or *heterogeneous*:

- | | | |
|--------------|-------------|--------------|
| a a diamond | d mercury | g concrete |
| b soil | e blood | h dry ice |
| c white wine | f an orange | i toothpaste |

- 2** **a** Classify the homogeneous items in Exercise 1 as *mixtures* or *pure substances*.
b In terms of mixture or pure substance, what can you say about *all* the heterogeneous substances in Exercise 1?
- 3** Which of the following are pure substances?
- | | | |
|---------------------------|-------------------------------------|--------------------------|
| a soft drink | d aluminium foil | g deodorant spray |
| b a silver pendant | e polyethylene wrapping film | h copper pipe |
| c exhaled breath | f a brick | i lead sheeting |
- 4** For the mixtures in Exercise 3, which are *homogeneous* and which *heterogeneous*?
- 5** A chemist had a red powder that was homogeneous and had all the characteristics of a pure substance. When a sample of it was heated strongly, drops of a heavy silvery liquid and a colourless gas formed. When the mixture of liquid and gas was cooled, there was no change. The chemist was unable to re-form the red powder. Is the red powder an element or a compound? Explain why.
- 6** A 1.00 g sample of a shiny grey solid was heated. At 327°C it changed to a silvery liquid; the temperature did not increase until all the solid had turned to liquid. When the liquid was cooled, it changed back to a grey solid; the mass was still 1.00 g. Is this substance an element or a compound? Explain why.
- 7** When 1.00 g of a particular pale blue solid was heated, it changed to a black solid. The mass of the black solid was 0.64 g. Suggest why there has been an apparent loss of mass. Is the original blue solid an element or a compound? Why?

Before beginning our survey of the thousands of substances in planet Earth that we use as resources for our everyday lives, let us identify the different zones of the Earth.

1.3 THE 'SPHERES' OF PLANET EARTH

Planet Earth is made up of several distinct layers as shown in Figure 1.2: the inner and outer core, mantle, crust and atmosphere. Some parts of Earth have the special names shown in Table 1.3.

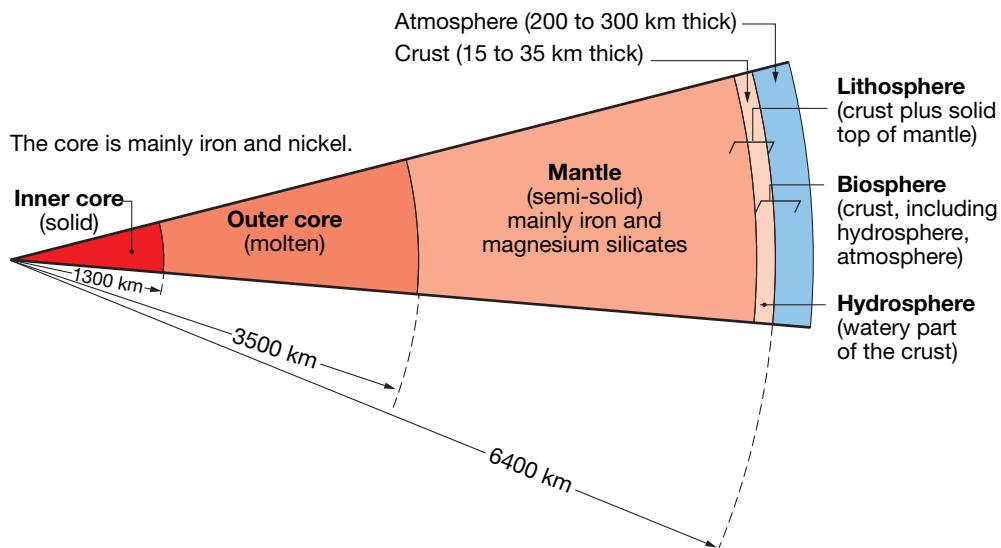


FIGURE 1.2
The various zones of planet Earth

TABLE 1.3 The various ‘spheres’ of Earth

<i>lithosphere</i>	crust plus the top portion of the mantle
<i>hydrosphere</i>	the water of the Earth’s crust: salt water of the oceans, fresh water of rivers and lakes, and ground water (water stored under the Earth’s surface)
<i>atmosphere</i>	layer of gas about 200 to 300 km thick that surrounds the planet (75% of the mass of the atmosphere is in the lower 15 km and 99.997% in the lower 90 km)
<i>biosphere</i>	the portion of Earth inhabited and used by living matter: the biosphere consists of the atmosphere, hydrosphere and lithosphere

1.4 THE EARTH AS MIXTURES OF SUBSTANCES

Planet Earth—or at least that portion of it accessible to humans—is made up almost entirely of mixtures. The nature and composition of the mixtures vary greatly from region to region.

The atmosphere

The **atmosphere** is a mixture of gases. It is predominantly a mixture of the *elements* nitrogen, oxygen and argon, though it contains small amounts of gaseous *compounds* such as water, carbon dioxide, nitrogen dioxide, sulfur dioxide and carbon monoxide.

The hydrosphere

The **hydrosphere** also consists of different mixtures. The major component of the mixtures in the hydrosphere is the *compound* water. Rivers and lakes are predominantly water that has dissolved in it small amounts of the *elements* oxygen and nitrogen and *compounds* such as carbon dioxide, and sodium, calcium and magnesium chlorides and sulfates. The total amount of dissolved matter in ‘fresh water’ is generally less than 0.05%. If only these dissolved substances are present in fresh water then it is a solution (that is, homogeneous). However, suspended solids and microscopic bacteria, algae and plant life are often present as well and these make the mixtures heterogeneous.

Sea water is a mixture of water and sodium chloride (about 3.5%) along with smaller amounts of other *compounds* such as magnesium and calcium chlorides and various bromides, iodides and sulfates.

Ground water (water stored under the Earth’s surface) is a mixture similar to river water, except that it generally contains larger amounts of dissolved solids such as chlorides and sulfates.

Apart from dissolved oxygen and nitrogen, the mixtures of the hydrosphere contain only *compounds*.

The lithosphere

The **lithosphere** contains an extremely diverse range of mixtures. There are:

- *rocks* which are mixtures of various silicates (compounds of silicon, oxygen and various metals)
- *sand* which is mainly silicon dioxide with variable amounts of ground-up shells or finely divided dirt

- *soils* which are mixtures of various aluminosilicates (clays, compounds containing aluminium, silicon, oxygen and metals), sand, and decomposing animal and vegetable matter
- *mineral ores* which are mainly oxides, sulfides, carbonates, sulfates and chlorides of metals mixed with various silicates or aluminosilicates, and
- *coal, oil and natural gas* which are mixtures of compounds of carbon formed from decayed plant and animal matter.

The mixtures of the lithosphere contain predominantly *compounds*. There are some mixtures (fairly rare) which contain uncombined *elements* such as gold, silver, platinum and carbon (as diamond).

Living matter

Living matter such as animals, plants, algae and bacteria consists of mixtures of carbon-containing *compounds* (along with water and small amounts of minerals); the main types are carbohydrates (or sugars), proteins, fats and vitamins. Many of the carbohydrates and proteins are insoluble (cellulose, skin, hair) and make up the structure of the organisms, while others are soluble and so cause the functioning of the organism. Free elements (apart from dissolved oxygen and nitrogen) are extremely rare in living matter.

1.5 USING THE RESOURCES OF THE EARTH

A few of the mixtures we obtain from the Earth are used ‘as is’; that is, without being separated into their pure components. Examples are coal (for fuel), sand (for making concrete and glass), sandstone (for paving and facing buildings), and various clays (for earthenware pots and pipes and bricks).

However, most of the mixtures we extract from the Earth have to be separated into pure substances before we can use them. For example, the compound aluminium oxide has to be separated from the mixture bauxite before we can extract the useful element aluminium from it; the important compound sodium chloride (common salt) has to be separated from sea water (a mixture); the element oxygen (used medically and in oxyacetylene welding) has to be separated from the mixture air before it can be used.

Because separating mixtures into their components plays such an important part in our use of Earth resources, we need to look at some common methods of doing this.

We make many useful items from substances obtained from the Earth: bricks, ceramic tiles and clay pots from various clays, glass from sand, limestone and sodium carbonate, and plaster (the white powder) from gypsum (the pale brown powder)



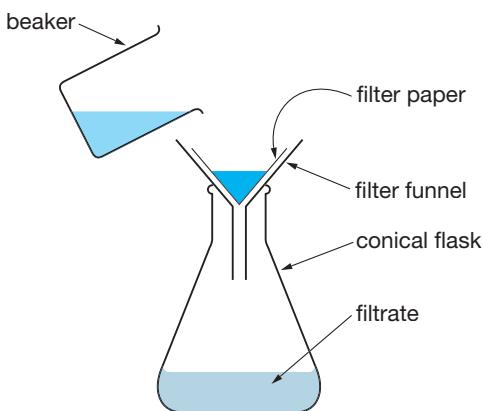
1.6 SEPARATION OF SOLIDS OF DIFFERENT SIZES

Mixtures in which the particles of the different substances have different sizes can be separated by **sieving**.

In the home kitchen we often use a sieve (piece of gauze held in a metal or plastic frame) to separate out lumps from a powdery substance such as flour. Similarly in the laboratory and in industry sieves are used to separate small particles from large ones. At quarries fine sand, needed for making concrete and mortar, is separated from the coarser material (sold as gravel) by sieving.

1.7 SEPARATING SOLIDS AND LIQUIDS

Mixtures of solids and liquids are commonly separated by **filtration** as shown in Figure 1.3. The liquid or solution passes through the paper while the suspended solid remains on the top of the filter paper. The liquid or solution that passes through the filter paper is called a **filtrate**. Sand can be separated from sea water in this way.



Sometimes if the solid is present as coarse or very dense particles (such as coarse sand in water), sedimentation and decantation can be used. **Sedimentation** is the process in which solids settle to the bottom of a container. **Decanting** or **decantation** is the process of carefully pouring off the liquid and leaving the solid undisturbed at the bottom of the container. Pouring tea off tea leaves is decantation.

1.8 SEPARATING DISSOLVED SOLIDS IN LIQUIDS

When a solid is dissolved in a liquid (that is, when we have a *solution*), the solid and liquid can be separated by **vapourising** off the liquid (called the *solvent*). We can do this either by boiling the solution (visible bubbles of vapour form) or by just evaporating it. In **evaporation** the liquid is heated to a temperature below its boiling point (or just let sit in an open vessel on the laboratory bench) so that some of the particles 'escape' from the surface of the liquid into the air and get blown away. Evaporation is a slower process than boiling.

Chemists frequently use the expression **evaporate to dryness**. This means heating a solution in an evaporating basin to drive off all the solvent (the

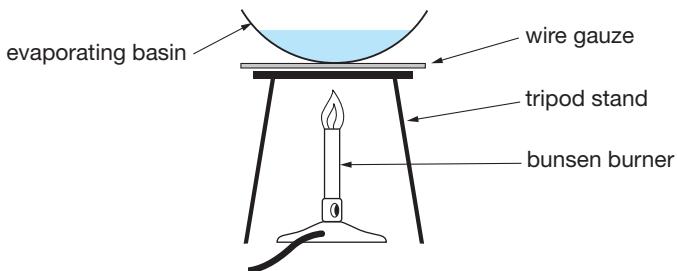


FIGURE 1.4

Evaporating a solution to dryness

liquid), and is illustrated in Figure 1.4. Evaporating to dryness is a common way of obtaining a solid from a solution.

Boiling or evaporating off the liquid is effective if it is the solid that we want (as in getting salt from sea water). However if it is the liquid that we want (as in getting fresh water from sea water), or if we want both the solid and the liquid, then we have to use distillation.

1.9 DISTILLATION

Distillation is a method of separating two or more liquids from one another or of separating the liquid from the solids in a solution; this latter process can be thought of as purifying the liquid.

Distillation is the process in which a solution or mixture of liquids is boiled with the vapour formed being condensed back to a liquid in a different part of the apparatus and so separated from the mixture.

The arrangement shown in Figure 1.5 is commonly used for distillation. The mixture or impure liquid, for example sea water, is placed in the flask and heated to boiling. The liquid changes to vapour, rises up the neck of the flask and diffuses down the side arm and into the water-cooled condenser, where the vapour is cooled and condensed back to a liquid, which is collected in the beaker. The liquid collected from a distillation is called a **distillate**.

If the impurities are non-volatile (as is the case with sea water), pure liquid is collected (pure water from sea water). Similarly, if some dirty mineral turps, which had been used to clean a paintbrush and which was highly coloured, were distilled[†], colourless turpentine would be obtained since the oils and pigments of the paint are non-volatile (compared with turps).



Collecting salt by evaporating sea water. In Vietnam people do the hard work while in Australia we use machines. Note that the sea water is not evaporated to dryness: it is evaporated until a large proportion of the salt has crystallised out. In this way, most of the impurities stay in the solution

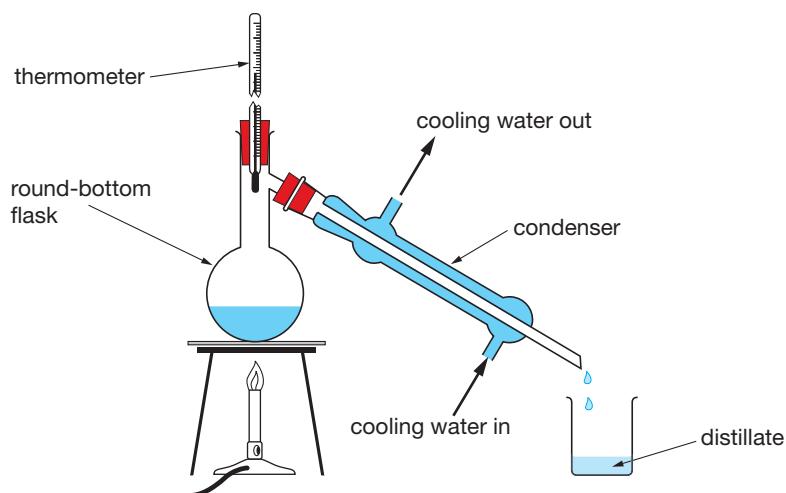


FIGURE 1.5
Simple distillation apparatus

[†] Since turps is flammable, this distillation must not be performed using an open flame; rather a hot plate or heating mantle must be used.

1.10 SEPARATING LIQUIDS

A mixture of two or more liquids can be separated by distillation if the boiling points of the liquids are sufficiently different (say by about 40 or 50°C). For example, distillation of a mixture of water and ethylene glycol (motor car anti-freeze) produces pure water as the distillate and leaves the glycol behind in the distillation flask. This happens because the boiling points are so different: 100°C for water and 198°C for ethylene glycol.

If we distil a mixture of two liquids (such as alcohol and water) which do not have greatly different boiling points (78°C and 100°C), then the distillate is not a pure substance. Instead it is a mixture, but generally the distillate is richer in the lower boiling point component. A mixture of 15% ethanol (ordinary alcohol) with 85% water, when distilled, produces a distillate of about 40% ethanol and 60% water. This is how brandy is made from wine, or whisky from fermented grain solution.

We often use the word *volatile* in these contexts: it means ‘able to be converted to a vapour’. The *more volatile* of two liquids is the one with the lower boiling point: ethanol is more volatile than water.

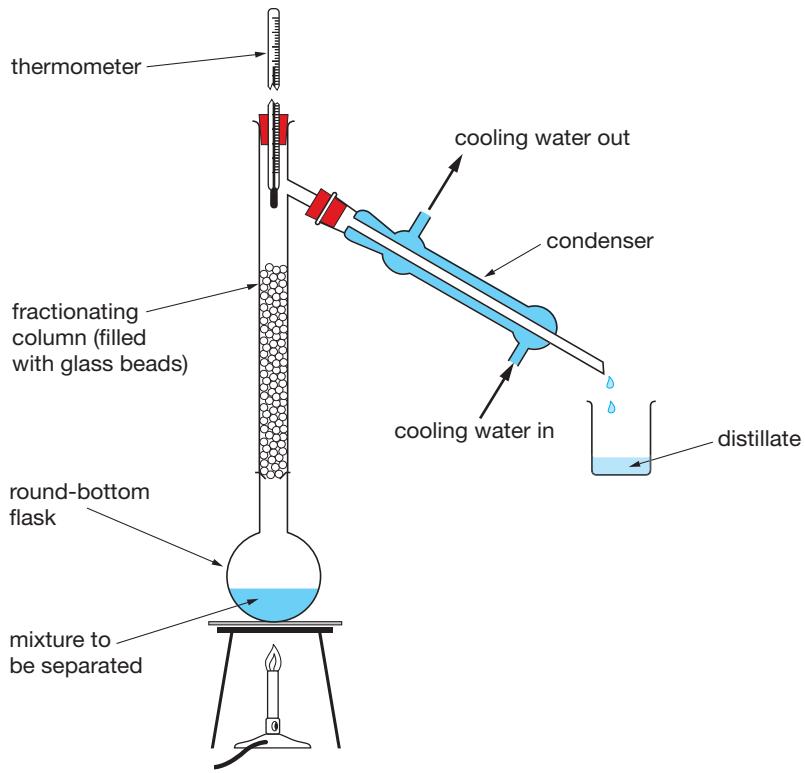
If a mixture of a volatile liquid with non-volatile impurities (solid or liquid) is distilled, the distillate is pure liquid. If a mixture of two liquids of comparable volatility (similar boiling points) is distilled, the distillate is generally richer in the more volatile liquid.

Despite what has just been said, it is possible to use distillation to separate liquids with similar boiling points if we use a process called fractional distillation.

Fractional distillation

To separate liquids by distillation when their boiling points are fairly close together we have to use many distillation steps. This can be done in specially

FIGURE 1.6
Fractionating column in the laboratory (commonly about 400 to 500 mm high)



designed equipment: the process is called **fractional distillation**. A typical laboratory fractionating column is shown in Figure 1.6. This arrangement allows for repeated condensations and vaporisations up the column, effectively giving many separate distillations (typically 10 to 100). This means that eventually a pure sample of the more volatile substance in the original mixture emerges from the top of the column. If a mixture of the two petrol-like liquids benzene and cyclohexene (boiling points 80°C and 83°C respectively) is fractionally distilled the final distillate is pure benzene. If ethanol–water mixtures such as wines are fractionally distilled the final distillate is 98% ethanol (with 2% water).

Fractional distillation is widely used in industry. Examples of this are:

- separation of crude oil into various commercial products
- separation of ethanol (for use as a fuel additive) from fermented solutions of sugar molasses or grain mashes
- production of liquid nitrogen and argon gas from liquefied air.

Exercises

- 8 A careless teacher inadvertently mixed the laboratory supplies of aluminium oxide and lead pellets. Although small, the lead pellets were larger than the fine crystals of aluminium oxide. Aluminium oxide is much less dense than lead. Suggest a way of separating these two substances.
- 9 Suppose you were given a solution of sucrose (sugar) in water. How would you separate this mixture if your aim was:
- a to obtain pure sucrose b to obtain pure water?
- 10 Use Table 1.7 on page 21 to answer the following:
- a If a mixture of alcohol (ethanol) and glycerol (boiling point 290°C) was distilled, what would be the composition of the distillate? Explain why.
- b Would distillation of a solution of acetic acid in water, for example vinegar, produce a distillate more concentrated or less concentrated (in acetic acid) than the starting solution? Give your reason.
- c If a mixture of chloroform and ethanol was distilled, the distillate would be richer (than the starting material) in which substance?
- d How would you obtain a pure sample of one liquid from the mixture in (c)?
- e If you distilled a mixture of hexane and heptane (boiling point 98°C), what would the distillate be?



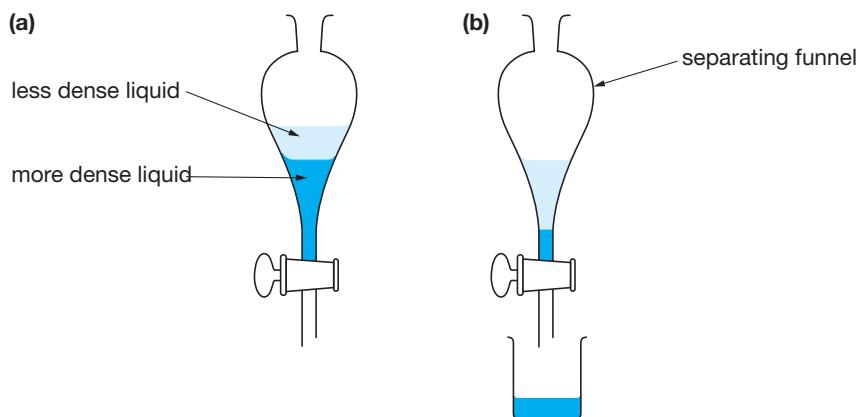
1.11 SEPARATING IMMISCIBLE LIQUIDS

Two liquids are said to be **immiscible** if, when they are mixed, they do not form a homogeneous liquid, but instead stay as drops of one liquid dispersed through the other liquid. If left standing for some time immiscible liquids separate into two distinct layers, one on top of the other. Examples are water and kerosene, water and cooking oil. Liquids that mix to form a homogeneous liquid are said to be **miscible**. Alternatively, we say that one liquid dissolves in the other.

Immiscible liquids are generally separated by using a **separating funnel**. This pear-shaped piece of apparatus, shown in Figure 1.7, tapers to a narrow tube just above the stopcock. This shape allows us to run out the bottom liquid without getting it contaminated with any of the top one. A mixture of petrol and water can be separated in this way.

FIGURE 1.7

- (a) Two immiscible liquids in a separating funnel,
- (b) just before all of the 'heavier' (more dense) liquid has been run out



1.12 SEPARATION BASED ON SOLUBILITY

Mixtures of solids can be easily separated if one solid is soluble in a particular solvent while the others are not. Sufficient solvent is added to the mixture to dissolve the soluble component; then the insoluble component(s) is(are) filtered off. The soluble solid is recovered by evaporating the filtrate (solution of that solid) to dryness. This is shown in Figure 1.8. A mixture of salt and sand can be separated this way by adding sufficient water to dissolve the salt, then filtering off the sand.

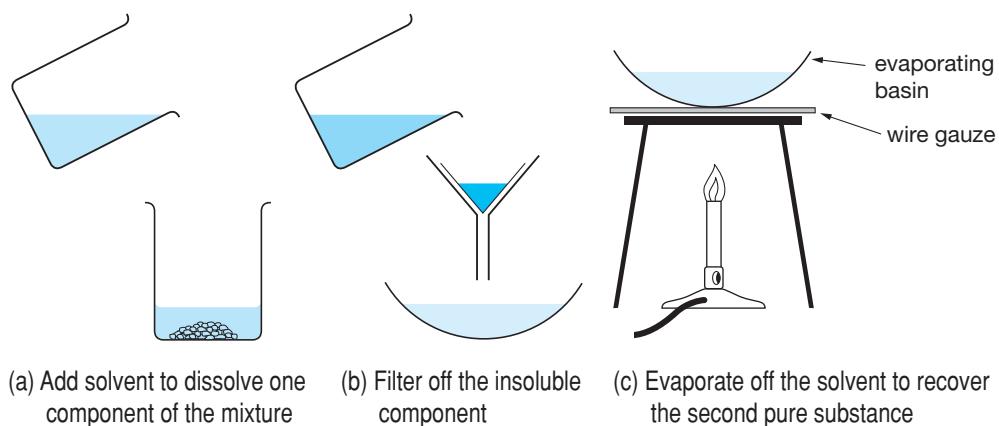
Solvents other than water can be used. For example, a mixture of iodine and charcoal (carbon) can be separated by adding hexane, which dissolves the iodine but not the carbon.

We will clarify the meaning of some terms before proceeding.

A **suspension** is a dispersion of particles through a liquid with the particles being sufficiently large that they settle out on standing.

FIGURE 1.8

Separating a mixture by filtration followed by evaporation



(a) Add solvent to dissolve one component of the mixture

(b) Filter off the insoluble component

(c) Evaporate off the solvent to recover the second pure substance

A suspension is heterogeneous in that the dispersed particles can be seen either by eye or by using a microscope. Examples are sand in water, milk and paint.

A **solution** is a homogeneous mixture in which the dispersed particles (molecules or ions) are so small that they never settle out and cannot be seen by a microscope.

Examples are salt or sugar in water, iodine in alcohol, and brandy. In a solution the substance that is dissolved is called the **solute**. The liquid that does

the dissolving is called the **solvent**. Sugar, common salt and caustic soda are solutes that readily dissolve in the solvent water. As solutes, iodine and cooking oil dissolve in the solvent hexane.

1.13 SEPARATING GASES

Gas mixtures are generally separated by using either differences in boiling points or differences in solubilities in liquids such as water. Boiling points of some common gases are shown in Table 1.4.

Argon for filling light bulbs and oxygen for medical uses and for oxyacetylene welding are obtained from air by fractional distillation. Fractional distillation rather than simple distillation is used because the boiling points of nitrogen, oxygen and argon are so close together (Table 1.4). First the air is liquefied (by cooling it to below -196°C), then the mixture is fractionally distilled. In principle the method is as in Figure 1.6, although of course at extremely low temperatures. Nitrogen, having the lowest boiling point, comes off first, followed by argon and finally oxygen.

TABLE 1.4 Boiling points of some common gases (at standard atmospheric pressure)

Gas	Boiling point ($^{\circ}\text{C}$)	Gas	Boiling point ($^{\circ}\text{C}$)
ammonia	-33 S ^a	hydrogen sulfide	-60 SS
argon	-186	methane	-161.5
butane	-0.5	nitric oxide	-152
carbon dioxide	-78 SS ^a	nitrogen	-196
carbon monoxide	-192	nitrogen dioxide	21 S
chlorine	-35 SS	oxygen	-183
ethane	-88.6	propane	-42
hydrogen	-253	sulfur dioxide	-10 S

^a S means soluble in water, SS means slightly soluble, all other gases are insoluble.

Natural gas contains carbon dioxide, hydrogen sulfide and water vapour in addition to the ‘useful’ substances, methane, ethane, propane and butane. It can be separated into ‘unwanted gases’ (carbon dioxide, hydrogen sulfide and water), propane and butane (for liquefied petroleum gas, LPG) and town gas (methane and ethane) by passing the mixture through a bubbler and pair of U-tubes as in Figure 1.9. The bubbler A contains sodium hydroxide solution (which dissolves

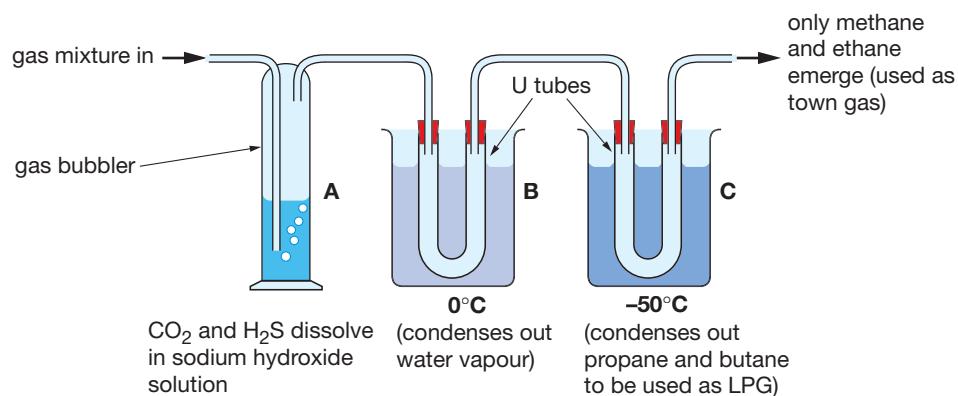


FIGURE 1.9
Separation of natural gas into unwanted gases, LPG and town gas

carbon dioxide and hydrogen sulfide more readily than does pure water). Tube B at a temperature of about 0°C condenses out water vapour while Tube C at about –50°C condenses out the propane and butane: methane and ethane pass through as gases to be used as town gas (check boiling points in Table 1.4).

1.14 SUMMARY OF METHODS OF SEPARATION

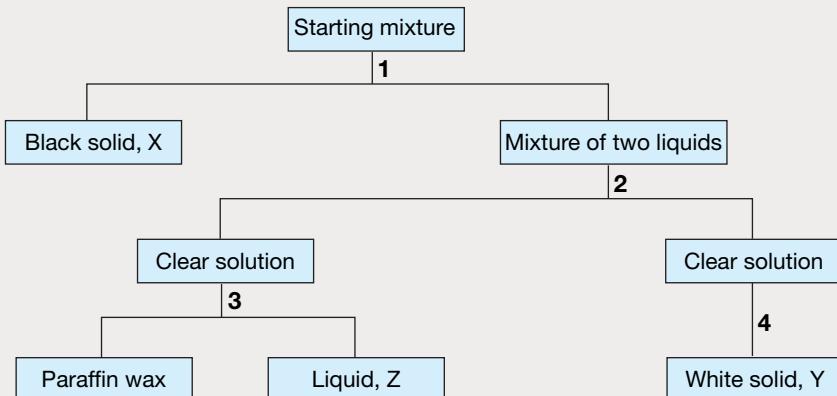
Table 1.5 summarises the methods discussed here for separating mixtures and lists the physical properties these methods depend upon. Because purification of a substance is removal of unwanted substances (impurities), these separation methods are also methods of purifying substances.

TABLE 1.5 Separation methods and the properties they depend upon

Separation method	Property used in the separation
sieving	particle size
vapourisation (evaporation or boiling)	liquid has a much lower boiling point than the solid
distillation	big difference in boiling points
fractional distillation	significant but small difference in boiling points
filtration	one substance a solid, the other a liquid or solution
adding a solvent, then filtration	one substance is soluble in the chosen solvent, while the others are insoluble
using a separating funnel	components are immiscible liquids

Exercises

- 11 In Figure 1.9 LPG (liquefied petroleum gas) was separated from natural gas. LPG is a mixture of propane and butane. How would you separate LPG into these two constituents? What difference in property are you using?
- 12 A school's supply of crystalline magnesium sulfate became contaminated with some barium sulfate. Draw a flow chart for the method you would use to separate these two substances, ending up with dry pure crystals of each substance. Magnesium sulfate is soluble in water; barium sulfate is not.
- 13 A pair of students were given a mixture of powdered charcoal, kerosene, paraffin wax, sodium sulfate and water. They were asked to obtain pure samples of the charcoal, kerosene, wax and sodium sulfate from this mixture. A flow chart of the procedure they followed is shown below.



- a** Name the separation procedures, 1, 2, 3 and 4, that they used. Draw a diagram showing how each of these procedures would be performed in the laboratory.
- b** Identify the solids X and Y and the liquid Z.
- c** How would you vary their procedure in order to recover a pure sample of water as well?
- ***14** Draw a flow chart to show how you would separate a mixture of carbon, iodine and ammonium chloride into pure samples of the three substances. Ammonium chloride is soluble in water, but insoluble in hexane. Iodine is slightly soluble in water but readily soluble in hexane. Carbon is insoluble in both solvents.
- 15** A river bed is a mixture of sand and small and large stones. A quarry operator wants to separate this mixture into fine sand (for making concrete), coarse sand (for use in gardens and under paving), small river pebbles (for landscaping) and large stones (for crushing into aggregate for concrete making). Suggest how this separation could be done. What differences in what property are you using?
- Use data in Table 1.4 for Exercises 16 and 17.
- 16** When ammonia is made industrially, not all of the nitrogen and hydrogen used to make it are used up. How would you separate the ammonia from the other gases?
- 17** The exhaust gas from a motor car contained the gases nitrogen, carbon dioxide, carbon monoxide, nitrogen dioxide and nitric oxide. How would you separate carbon dioxide and nitrogen dioxide individually from a sample of this gas? How would you then separate the three remaining gases?
- 18** Crude oil is a resource we collect from the Earth. It is a complex mixture of liquids of varying volatility and dissolved non-volatile solids. The main commercial products we want from it are petrol (highly volatile), kerosene and diesel fuel (less volatile), heating oil (even less volatile) and tar (non-volatile solid). How would you separate crude oil into these four fractions? What property is your separation based on, and how does this property vary among the components of the mixture?
- ***19** The tar obtained in Exercise 18 contains waxes also (such as paraffin wax). These waxes are soluble in hydrocarbon liquids (hydrocarbons are compounds of hydrogen and carbon) whereas tar is not. How would you separate such wax from tar? What difference in property are you using?
- 20** Bauxite is a naturally occurring ore from which aluminium is extracted. Bauxite is a mixture of aluminium oxide, iron(III) oxide and dirt. The first step in extracting aluminium is to separate aluminium oxide from the other substances. Aluminium oxide is soluble in hot concentrated sodium hydroxide solution but much less soluble in cold solution; iron(III) oxide and dirt are insoluble under all conditions. How would you separate aluminium oxide from bauxite?

FOR INVESTIGATION

Select an industry (agriculture, mining or manufacturing) that involves the separation of a mixture obtained from the lithosphere, hydrosphere or atmosphere, or from living matter. Describe fully that separation procedure, including at least one diagram. Identify the property or properties used to bring about the separation. Name the products of the separation and give the uses of the ‘useful’ ones. Discuss any waste disposal problems associated with the process.



1.15 PROPERTIES USED TO IDENTIFY PURE SUBSTANCES

We have been talking about separating mixtures into pure substances. Let's now look at some of the properties or characteristics that are used to identify pure substances. Common properties are colour, physical state, melting and boiling points, density, electrical conductivity, solubility in different liquids and mechanical properties.

1.16 COLOUR

Some substances have quite distinct colours. Examples are deep brown for liquid bromine, reddish brown for solid copper, pale blue for copper sulfate pentahydrate, yellow for sulfur, red for iron(III) oxide and deep purple for potassium permanganate. However, we should be cautious about identifying substances solely on their colour. This is because two substances often have very similar colours and so can be mistaken for each other. Nevertheless colour can be a good first clue.

1.17 PHYSICAL STATE AT ROOM TEMPERATURE

By physical state we mean whether the substance is a solid, liquid or gas. These **states of matter** as they are called have the characteristics given in Table 1.6.

Changes from solid to liquid to gas and vice versa are called **changes of state**.

The names for the various changes of state are also shown in Table 1.6.

TABLE 1.6 The three states of matter

Solid	Liquid	Gas
has a definite volume	has a definite volume	expands to fill the volume available to it
has a definite shape (bars or sheets of Al or Cu) or is made up of small pieces (crystals) that have definite shapes (sugar, salt, sand)	takes the shape of the container it is placed on	takes the shape of the container it is placed in
difficult to compress	difficult to compress	easily compressed
solid <i>melting, fusion</i> <i>freezing, solidification</i>	liquid	gas <i>vaporisation, evaporation, boiling</i> <i>condensation, liquefaction</i> <i>sublimation</i>

1.18 MELTING AND BOILING POINTS

The **melting point** of a solid is the lowest temperature at which the solid changes to a liquid.

It is the temperature at which solid and liquid can co-exist indefinitely. Each pure

substance has a definite melting point at atmospheric pressure, and it can be used to help identify the solid.

If a solid is a pure substance, its melting point is quite sharp. All of a sample of the solid melts at the one temperature. If the solid is a mixture (that is, an impure substance), then it melts over a range of temperatures. In addition, *an impure substance starts to melt at a lower temperature than does the pure substance*. Hence for solids, melting point can be used as a test of purity.

If the melting point is sharp and if it does not increase after submitting the solid to a further purification process, then we know that the substance is pure. The value of the melting point can be used as one piece of evidence for identifying the substance.

Melting points for some common substances are given in Table 1.7.

The reverse process of converting a liquid to a solid is called freezing. *The freezing point of a liquid is the same temperature as the melting point of the solid*. It is the *highest* temperature at which the liquid can be converted to a solid. A liquid can be converted back to a solid at any temperature below the freezing (melting) point.

There are two ways of vaporising a liquid (converting a liquid to a gas or vapour). One is to leave the liquid in an open container. Liquid slowly *evaporates*. The other is to heat the liquid until it boils.

TABLE 1.7 Melting and boiling points and densities for some common substances

Substance	Melting point (°C)	Boiling point (°C)	Density (g/mL ^a)	Substance	Melting point (°C)	Boiling (°C)	Density (g/mL ^a)
aluminium	660	2450	2.7	ethanol	-114	78	0.79
copper	1083	2600	9.0	ethyl acetate	-84	77	0.90
zinc	420	610	7.1	ethylene glycol	-16	198	1.11
lead	327	1740	11.4	acetic acid	16.7	118	1.04
mercury	-39	357	13.6	chloroform	-64	62	1.48
sulfur	114	444	2.0	hexane	-95	69	0.66
phosphorus	44	280	1.8	oxygen	-219	-183	
carbon (graphite)	3730	4830	2.3	nitrogen	-210	-196	
sodium	98	892	0.97	hydrogen	-259	-253	
water	0.0	100.0	1.00				

^a at 25°C

Boiling is the process of converting a liquid to a vapour (gas) by heating the liquid until visible bubbles of vapour form throughout the whole bulk of the liquid and quickly rise to the surface. We call the lowest temperature at which this occurs the **boiling point**.

Boiling is easily detected by simply observing the bubbles that form.

Pure substances boil at fixed temperatures (at constant atmospheric pressure), so the value of the boiling point can be used as one step towards identifying a liquid substance. Boiling points do vary with the prevailing atmospheric pressure so the term **normal boiling point** is introduced, meaning the

boiling point at standard atmospheric pressure. However, the ‘normal’ is often omitted when there is little doubt about what is intended. Table 1.7 gives the normal boiling points of some common substances.

If a substance A has a lower boiling point than substance B, we say that A is *more volatile* than B. A is more easily converted to a vapour.

Pure substances have sharp boiling points; that is, all of the sample boils at the one temperature. Mixtures boil over a range of temperatures. A non-volatile impurity in a liquid raises the boiling point of the liquid: a solution of salt in water boils at a slightly higher temperature than does pure water. An impurity that is more volatile than the liquid lowers the boiling point: solutions of alcohol in water boil at lower temperatures than does pure water.

Boiling point can therefore be used as a test of purity. If a liquid has a sharp and constant boiling point, and if it does not change after putting the liquid through a further purification process, then the liquid is a pure substance and the value of the boiling point can be used to help identify the substance.

Note that impurities can raise or lower the boiling point; they can only lower the melting point.

The process of *converting a gas (vapour) to a liquid is called condensation* or **liquefaction**. Condensation of a vapour back to a liquid can occur at any temperature less than or equal to the boiling point of the liquid.

Exercises

21 Melting and boiling points for some substances are shown in the table.

Substance	Melting point (°C)	Boiling point (°C)	Substance	Melting point (°C)	Boiling point (°C)
sulfur	114	444	sulfur hexafluoride	-51	-63
gallium	30	2400	lead bromide	373	914
neon	-249	-246	magnesium	650	1110
bromine	-7	58	phosphorus trichloride	-91	74
carbon disulfide	-111	46	carbon tetrabromide	91	190
hydrogen peroxide	-2	158	sulfur dioxide	-73	-10

- a Which of these substances would be:
- i solids ii liquids at room temperature (20°C)?
- b Arrange the solids in order of increasing melting point.
- c Arrange the liquids in order of increasing boiling point.
- d Which of these substances would *not* be solids at -60°C?
- 22 A small portion of a white solid was slowly heated. It started to melt at 88°C but had not completely changed to liquid until 93°C. The rest of the white solid was then put through a purification procedure. A small portion of the ‘purified’ solid melted sharply at 91°C. A further ‘purification’ did not change the melting point. Was the original white solid pure or not? Why? Was the final white solid pure? Why? Use Table 1.7 or the one in Exercise 21 to identify the solid.

1.19 DENSITY

Density is defined as mass per unit volume.

Common units for density are grams per millilitre, g/mL, or kilograms per cubic metre, kg/m³.

Density of a liquid can be determined by measuring the mass of a known volume of the liquid, for example from a pipette. For a geometrically shaped piece of a solid, density can be determined by measuring the dimensions of the solid, calculating its volume, then measuring its mass. In each case:

$$\text{density} = \frac{\text{mass}}{\text{volume}} \quad \dots (1.1)$$

Density is a characteristic property of pure substances and can be used to help identify substances. Densities of some common substances are shown in Table 1.7.

Exercises

23 Calculate the density of:

- a a block of aluminium measuring 0.50 cm × 1.20 cm × 0.80 cm which has a mass of 1.30 g
- b 5.0 mL chloroform which has a mass of 7.40 g
- c a cylinder of zinc, 3.0 mm in diameter and 3.0 cm long, which has a mass of 1.506 g

24 Using density values from Table 1.7, calculate the mass of:

- a 5.0 mL ethanol
- b a cube of copper having a side of 0.80 cm
- c 25.0 mL hexane

25 Using density values from Table 1.7, determine what volumes of the following substances would have the following masses:

- a 10 g chloroform b 25 g ethylene glycol c 250 g copper



Other properties used to identify substances, such as electrical conductivity, mechanical properties and solubility in different liquids, will be discussed later (in Sections 1.23, 4.2 and 7.13 respectively).

1.20 GRAVIMETRIC ANALYSIS

We have talked about separating mixtures. Sometimes we want to know the quantitative composition of mixtures; that is, to know what mass of each substance is present in a given quantity of the mixture, or alternatively to know the per cent composition of the mixture. Determining the quantities (masses) of substances present in a sample is called **gravimetric analysis**—analysis by weight or by mass.

Reasons for wanting to know the per cent composition of a mixture include:

- to decide whether a newly discovered mineral deposit contains a sufficiently high percentage of the required compound to make its extraction from that deposit economically viable

- to determine the composition of soil in a particular location to see if it is suitable for growing a certain crop
- to determine the amounts of particular substances present in water or air to decide how polluted the samples are
- to decide whether a particular commercial mixture has the same percentage composition as a similar mixture being marketed by a rival company.

To carry out a complete gravimetric analysis on a sample we need to determine the mass of each component present in the mixture. An example will illustrate.

Many commercial products such as those shown here give their composition, either in percentage or mass forms, on their labels



Example

A team of geologists discovered a new mineral in a remote desert location; it was a mixture of barium sulfate and magnesium sulfate. Its composition was determined as follows. They first ground up a 3.61 g sample with water; magnesium sulfate dissolves, barium sulfate does not. The barium sulfate was filtered off, dried and its mass determined to be 1.52 g. They evaporated the filtrate to dryness to recover the magnesium sulfate, and determined its mass to be 2.07 g. Calculate the percentage composition of the sample.

First we should check that the sample does in fact contain only barium and magnesium sulfates. The total mass of the two sulfates is $1.52 + 2.07 = 3.59$ g. This is within experimental error of 3.61 g (less than 1% discrepancy), so the sample contains no other substance.

$$\text{Percentage of barium sulfate} = \frac{\text{mass of barium sulfate present}}{\text{total mass of the sample}} \times 100$$

$$= \frac{1.52}{3.61} \times 100$$

(It does not matter whether we use 3.61 or 3.59; there is some experimental error.)

$$= 42\%$$

Since there are only two components in the mixture:

$$\begin{aligned}\text{percentage of magnesium sulfate} &= 100 - 42 \\ &= 58\%\end{aligned}$$

The sample consists of 42% barium sulfate and 58% magnesium sulfate, with accuracies of about $\pm 1\%$.

We could have calculated:

$$\text{percentage of magnesium sulfate} = \frac{2.07}{3.59} \times 100 = 58\%$$

but that is more work!

Gravimetric analysis can be used to determine the per cent composition of compounds also. This will be treated in Sections 5.7 and 5.8.

Exercises

- 26 The suitability of water for irrigating crops or for animals to drink depends upon the amount of dissolved solids present. To assess the suitability of water from a particular bore, a grazier took 500 g of the water, evaporated it to dryness, then determined the mass of solids remaining to be 3.63 g. Calculate the percentage of dissolved solids in this water. In addition, express the concentration of solids in parts per million, ppm; that is, grams of solids per million grams of the water.
- 27 Upon analysis a 3.67 g sample of a certain NPK fertiliser (nitrogen, phosphorus, potassium) was found to contain 1.79 g urea and 0.81 g ammonium phosphate with the rest being potassium chloride. Calculate the per cent composition of this fertiliser.
- 28 The product of a certain industrial process was a mixture of three alcohols: ethanol (ordinary alcohol), ethylene glycol (motor car anti-freeze) and glycerol (used in sweets and cosmetics). The boiling points are 78°C, 198°C and 290°C respectively. 18.33 g of this mixture was carefully distilled. 3.79 g of distillate was collected while the thermometer in the apparatus registered 78°C. The temperature rose to 198°C and 8.64 g of a second distillate was collected. When the temperature started to rise above 200°C the distillation was stopped. Assuming that what was left in the distillation flask was glycerol, calculate the per cent composition of the original mixture.



Most of the elements occur on Earth as compounds; very few are present as free (or combined) elements. Let us look more closely at this.

1.21 ELEMENTS OCCURRING ON EARTH AS FREE ELEMENTS

Oxygen and nitrogen occur as uncombined elements in the atmosphere. By mass the atmosphere (excluding water vapour) is 21% oxygen and 78% nitrogen.

Most of the remaining 1% of the atmosphere is argon, the most abundant of the noble gases. Trace amounts of the other noble gases, helium, neon, krypton, xenon and radon, are also present in the atmosphere. These **noble gases** are so called because they form virtually no compounds. Helium is also found in natural gas wells, which are its major commercial source. Radon is formed by the radioactive decay of uranium and so is released to the atmosphere when uranium is mined. Like uranium, radon is radioactive.

Other elements that occur in some places as uncombined elements are sulfur and the metals, gold, silver and platinum (and to a small extent copper).

1.22 WHY MOST ELEMENTS ON EARTH OCCUR AS COMPOUNDS

Apart from the few elements mentioned in the previous section, virtually all the rest of the 90 or so elements occur as compounds. This is because most elements are chemically reactive; that is, when they come into contact with certain other elements they react to form compounds. Despite some anomalies, as a general rule:

the more reactive an element is, the less chance there is of finding it in the Earth as an uncombined element.

Sodium, potassium, calcium, magnesium, fluorine and chlorine which are all very reactive elements are never found as free elements. Copper and sulfur which have moderate reactivity do exist naturally in some locations as uncombined elements, although they are more commonly found as compounds. Gold, platinum, and the noble gases such as argon and helium, which are extremely unreactive, occur naturally as uncombined elements.

1.23 METALS AND NON-METALS

Chemists have found it convenient to classify the elements into what are called metals and non-metals.

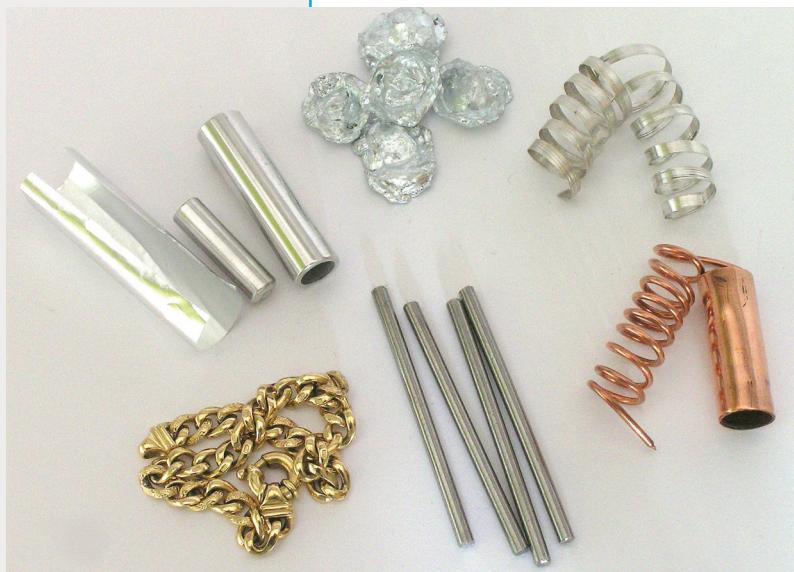
Metals are elements which:

- are solids at room temperature
- have a shiny or lustrous appearance
- are good conductors of heat and electricity
- are malleable (able to be rolled into sheets) and ductile (able to be drawn into wires).

Most other elements are called **non-metals**.

On this basis, aluminium, cobalt, copper, gold, iron, lead, magnesium, nickel, potassium, sodium, silver, tin and zinc are clearly metals. Argon, bromine, chlorine, hydrogen, iodine, nitrogen, oxygen, phosphorus and sulfur are non-metals.

There are however some problem elements. Mercury has a shiny appearance and is a good conductor of electricity, but it is a liquid. We choose to classify



mercury as a metal because, on balance, its properties are closer to those of the other metals than the non-metals. Carbon in the form of graphite is a fair conductor of electricity and is a solid. Despite this we classify it as a non-metal because, considering all its properties, it resembles the non-metals more closely than the metals.

As with any classification into two categories, there are some elements which are difficult to classify as metals or non-metals because they have some properties of both classes. Boron, silicon, germanium, arsenic, antimony and tellurium do not fall clearly into either category. We therefore call them **semi-metals**.

This classification of the elements into metals and non-metals is useful because very frequently, as we shall find in later chapters, the metals have one property while the non-metals have a different one.

1.24 PHYSICAL PROPERTIES AND USES OF ELEMENTS

Physical properties often influence the choice of an element for a particular purpose.

The main uses for common metals such as iron, aluminium, copper and lead are for building materials, cars, planes, trains, machinery, electrical wiring, domestic appliances and household goods. It is the physical properties of the metals that most often determine which metal will be used for a particular product. The physical properties most commonly involved are melting point, density, electrical conductivity, hardness and tensile strength.

- Aluminium is used for making aircraft primarily because of its low density (combined with adequate mechanical strength).
- Iron (or steel, which is more than 98% iron) is used to make motor cars and trains because it has high tensile strength. Because weight is less of a consideration here, the moderately high density of iron is of little concern.
- Copper is used for electrical wiring in houses and other buildings because of its high electrical conductivity. However, for high voltage transmission lines aluminium is preferred (because it has similar electrical conductivity but has a much lower density; the mass of cables is important in determining the spacing of the expensive transmission towers).
- Tungsten is used for filaments in electric light bulbs because of its high melting point.

Although physical properties are important, often other factors such as chemical reactivity (or resistance to corrosion) and cost are significant also. Properties and uses of metals will be discussed more fully in Sections 4.1 to 4.3.

Only a few non-metals are used in elemental (uncombined) form and again physical properties are often important in selecting the element for a particular purpose.

- Carbon as graphite is used as electrodes in common dry cells (batteries) because of its significant electrical conductivity and as a dry lubricant because of its slippery (or soft) nature.
- Carbon in the form of diamond is used in jewellery because it is extremely hard (and so resists scratching) and it scatters light very effectively (it has a high refractive index).
- Liquid nitrogen is used as a cooling agent (for extremely low temperatures) because of the suitability of its freezing (melting) and boiling points.

Chemical properties are often more important in selecting non-metals for particular uses. For example, argon is used to fill electric light bulbs because of its extreme inertness (lack of reactivity) and helium is used to fill balloons because it is far less chemically reactive than hydrogen, which has an even lower density.

The major uses of non-metals are as compounds. Wide ranges of compounds of non-metals are used in fuels, fertilisers, building materials, furniture, plastics, synthetic fibres, drugs, pesticides, paints, detergents, adhesives and cosmetics.

1.25 THE PERIODIC TABLE

Because there are so many elements (about 90) with such a wide variety of properties, chemists have devised a chart which organises the elements into groups with similar properties. It is called the **Periodic Table**. A typical Periodic Table is shown on the inside front cover of this book.

In this table the elements are arranged so that those with similar properties fall into the same vertical column. Lithium, Li, sodium, Na, potassium, K, rubidium, Rb, and caesium, Cs, all have very similar properties; they fall into the extreme left-hand column of the table. Beryllium, Be, magnesium, Mg, calcium, Ca, strontium, Sr, and barium, Ba, have similar properties; they fall into another vertical column—the one second from the left. Fluorine, F, chlorine, Cl, bromine, Br, and iodine, I, have similar properties; they are also in the same vertical column, the one second from the right.

The Periodic Table uses *symbols* for the elements. These will be discussed in Section 2.4.

The vertical columns are called **groups**. The main groups of the table, the ones with at least five elements in them, are numbered from 1 to 7 with the right-hand group being numbered either 0 or 8. Group 1 contains Li, Na, K, Rb, Cs. The elements Be, Mg, Ca, Sr and Ba belong to Group 2, while F, Cl, Br and I belong to Group 7. The ten groups of elements between Groups 2 and 3 are not normally numbered. The elements in these groups are called **transition elements**. The other elements (in Groups 1 to 7 and 0) are called **main-group elements**.

The horizontal rows are called **periods** and they too are numbered from 1 to 7. Period 1 contains only two elements, hydrogen, H, and helium, He. Periods 2 and 3 each contain eight elements—lithium, Li, to neon, Ne, and sodium, Na, to argon, Ar, respectively. Periods 4 and upwards contain larger numbers of elements (because they contain transition elements).

There is often a gradual change in properties as we go across any one period. While periods are useful for seeing trends in properties, they are less useful than the groups for working out actual properties of particular elements.

A big advantage of the Periodic Table is that it allows us to deduce the properties of elements we may not be familiar with. For example, if we know nothing about rubidium, Rb, we note that it occurs in the same group of the table as sodium, Na. Hence the properties of rubidium and the compounds it forms are very similar to those of sodium. If we know that sodium reacts with water, then the table tells that rubidium also reacts with water.

Similarly we can use the table to deduce that strontium will undergo the same reactions as magnesium and calcium, or that bromine and iodine will form the same type of compounds as chlorine.

As our study of chemistry progresses, we will see that the Periodic Table provides much more information than this. For present purposes it is sufficient

to appreciate that the table tells us which elements have similar properties; that is, those that are in the same group.

Metals and non-metals in the Periodic Table

On the Periodic Table the non-metals occur near the top and right of the table. Most of the other elements are metals. A diagonal band of semi-metals separates the metals from the non-metals. This division is shown in Table 1.8.

TABLE 1.8 Division of the elements into metals and non-metals

H																		He			
Li	Be															B	C	N	O	F	Ne
Na	Mg															Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br		Kr			
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I		Xe			
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At		Rn			
Fr	Ra	Ac	Ru	Db																	

Elements to the left of the blue shaded zone are metals; elements to the right (in the pink shaded zone) are non-metals. Elements within the blue shaded zone have properties intermediate between metals and non-metals and are therefore hard to categorise; they are called **semi-metals**. Hydrogen is usually classed as a non-metal



The only two elements that are liquid at room temperature, bromine and mercury

Physical states of the elements

Of the naturally occurring elements, at room temperature:

- *two* are liquids: mercury and bromine
- *eleven* are gases: oxygen, nitrogen, hydrogen, helium, neon, argon, krypton, xenon, radon, fluorine and chlorine. All of these are colourless except fluorine which is pale yellow and chlorine which is greenish yellow.
- the rest are solids.



WEBSITE

<http://www.webelements.com>

(presents a Periodic Table; click on any element to see a list of its properties and a photograph of a sample of the element: a very useful source of data about elements)

Exercises

- 29** Which of the elements A, B, C, D, E in the table below would you class as metals?
Give your reasons.

Element	Melting point (°C)	Relative electrical conductivity ^a
iron	1540	100
sulfur	113	<0.01
A	50	50
B	10	0.04
C	800	0.02
D	1450	140
E	3400	200

^a Relative to iron arbitrarily set at 100.

- *30 a** Give the names of *five* metals that are used in your home and state a use for each.
Which property makes the metal particularly suitable for that use and why?
b List *four* items of jewellery that you or your family own and that are made of predominantly pure elements. Name the elements. Which properties make these elements particularly suitable for jewellery?
c Name six compounds that are purchased as virtually pure compounds for use in the home. State what each is used for.

For Exercises 31 to 33 you will need to consult sources of information other than this book[†].

- *31 a** Draw up a table of the following properties for the elements listed below: melting point, density, electrical conductivity, hard or soft, tensile strength (high, moderate, low), brittle or malleable. Include in your table one common use of each element.
cadmium, iodine, lithium, nickel, platinum, phosphorus, sulfur, titanium
b Classify each element in your table as a metal, semi-metal or non-metal, giving your reasons.
c Arrange the metals from **b** in decreasing order of
i melting point **ii** density **iii** electrical conductivity.
- 32 a** For *five* non-metals that are solids at room temperature, draw up a table of element, melting point and density.
***b** Arrange these elements in order of decreasing (i) melting point (ii) density.
- 33 a** For *eight* elements that are gases at room temperature, draw up a table with the headings Element, Boiling point, Gas at -190°C (yes or no).
***b** List the elements in (a) in the order in which they would condense to liquid as the temperature was gradually lowered from room temperature to -270°C.

[†] A useful source is G Aylward, & T Findlay, *S.I. Chemical Data*, 5th edition, Wiley, 2002.

Important new terms

You should know the meaning of the following terms:

atmosphere (p. 10)
biosphere (p. 10)
boiling (p. 21)
boiling point (p. 21)
change of state (p. 20)
compound (p. 6)
condensation (p. 22)
decantation (p. 12)
density (p. 23)
distillate (p. 13)
distillation (p. 13)
element (p. 6)
evaporate to dryness (p. 12)
evaporation (p. 12)
filtrate (p. 12)
filtration (p. 12)
fractional distillation (p. 14)
gravimetric analysis (p. 23)
groups (p. 28)
heterogeneous (p. 5)
homogeneous (p. 5)
hydrosphere (p. 10)
immiscible (p. 15)

impure substance (p. 5)
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solute (p. 16)
solution (p. 16)
solvent (p. 17)
states of matter (p. 20)
suspension (p. 16)
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vapourising (p. 12)

Test yourself

- Explain the meaning of each of the items in the ‘Important new terms’ section above.
- Name four substances which are homogeneous and four which are heterogeneous.
- Name four substances which are mixtures and four which are compounds.
- Name six elements and four compounds you are likely to find around your home.
- Sketch the portions of planet Earth near the surface and identify the lithosphere, hydrosphere, atmosphere and biosphere.
- What substances are present in **a** the atmosphere? **b** the hydrosphere? State which are present as uncombined elements and which as compounds.
- Describe some of the substances present in the lithosphere, indicating whether they are mixtures or pure substances.
- What mixtures can you separate by sieving? Upon what physical property does sieving depend?
- Explain how sedimentation and decanting can be used to separate a mixture; give a specific example.

- 10** Describe one specific example of using the procedure evaporating to dryness to separate one substance from a mixture.
- 11** How would you separate two liquids having boiling points of
a 88°C and 175°C respectively and b 110°C and 122°C respectively?
Draw the apparatus you would use for a.
- 12** Explain how you would separate two solids if one was soluble in water while the other was insoluble. Give full details and include a diagram.
- 13** How would you separate a mixture of cooking oil and water? Name and draw the apparatus you would use.
- 14** How would you separate oxygen, nitrogen and argon from a sample of air?
- 15** List five common physical properties used to characterise (identify) pure substances.
- 16** What names do we give to the following changes of state: solid to liquid, liquid to gas, liquid to solid, gas to liquid?
- 17** For any one substance, how are the melting point of the solid and the freezing point of the liquid related?
- 18** How can we use melting point to decide if a sample of a solid is pure or not?
- 19** How can we use the boiling point of a liquid to decide if the liquid is pure?
- 20** How would you measure the density of:
a a liquid b a solid of which you had an exactly cubic block?
- 21** Arrange the states of matter—liquid, gas, solid—for a particular substance in order of increasing density. Is this order true for all substances?
- 22** Suggest three situations in which a gravimetric analysis of a mixture would be useful.
- 23** List five elements that occur on Earth as uncombined elements.
- 24** Why do most elements occur as compounds?
- 25** How do you distinguish metals from non-metals?
- 26** Where do the non-metals appear in the Periodic Table? Where do the semi-metals appear?
- 27** What are the broad general uses for metals in our everyday life?
- 28** Give common uses of three non-metals that are used as elements. What properties make them suitable for these uses?
- 29** Name two elements you would expect to have similar properties to
a magnesium b chlorine c nitrogen d lithium.
- 30** Name two elements that are liquids at room temperature and eight that are gases.

Atoms, molecules and ions

IN THIS CHAPTER

Particle nature of matter
Atoms and molecules
Symbols and formulae
Nucleus and electron cloud
Structure of some simple atoms
Atomic number and mass number
Energy levels for electrons
Electron configurations
Stable electron configurations
The Periodic Table
Achieving noble gas configurations

Formation of ions
Ionic bonding and the Periodic Table
Covalent bonding
Covalency and the Periodic Table
Electron-dot structures for ions
Properties of ionic and covalent molecular substances
Covalent lattice solids
Metallic bonding
Properties of solids summarised

In Chapter 1 we took a macroscopic view of the materials of the Earth. We talked about substances we could see and hold, separations we could perform in a laboratory, and properties we could easily observe or measure. However, an important part of chemistry is trying to explain macroscopic observations in terms of the microscopic structure of the substances—the particles that make up materials and the ways in which they interact. In this chapter we shall focus on the particles that make up matter. We shall start with the simplest ‘particle’ view of matter, in which we make no attempt to identify the nature of the particles.

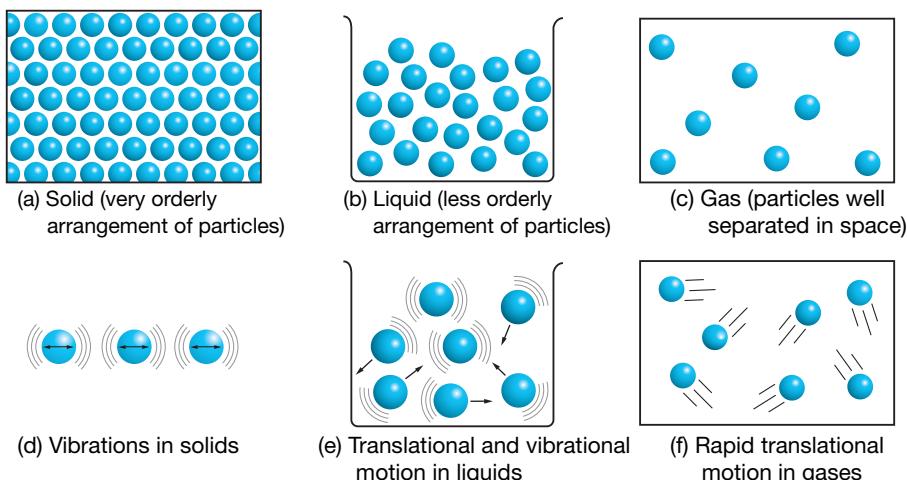
2.1 PARTICLE NATURE OF MATTER

All matter is made up of small particles which we often imagine as tiny spheres. In solids, these particles are packed closely together in an orderly array as shown in Figure 2.1(a). There are strong forces holding the particles to one another. These cause solids to have definite shapes and to be relatively hard. In solids the particles are not stationary as Figure 2.1(a) may suggest, but rather they are vibrating slightly about their rest positions as shown in Figure 2.1(d).

In liquids, the particles are arranged in a much less orderly fashion than in solids, as shown in Figure 2.1(b), and they move about much more freely. Particles in liquids possess more motion—more kinetic energy, we say—and the forces between neighbouring particles are comparatively weaker than in solids. Hence liquids do not have definite shapes but instead take the shape of

FIGURE 2.1

The particle nature of, and motions in, solids, liquids and gases



the container they are placed in. Liquids flow and deform easily. The motion of particles in liquids is not confined to vibration as in solids, but also involves random translation from one place to another in the bulk of the liquid as Figure 2.1(e) shows.

Because the particles are generally packed more closely together in solids than in liquids, solids have greater densities than liquids. (The exception is water and ice which we shall discuss in Section 7.7.)

In gases, the particles are much further apart than in solids or liquids, as shown in Figure 2.1(c), and they are in very rapid random motion as Figure 2.1(f) shows. The particles are so far apart and moving so rapidly that there are no significant forces between them. Because the particles are in such rapid random motion, gases quickly spread out to fill the whole volume available to them.

We imagine the particles making up matter as being hard and incompressible, so compressing materials means pushing particles closer together. In solids and liquids the particles are already quite close to one another. This means that solids and liquids cannot be compressed very much. In gases there are big spaces between particles. This means that it is relatively easy to push the particles closer together; hence gases are easily compressed.

So far we have been talking about the ‘particles’ that make up substances and explaining some properties in terms of how close the particles are to one another, what motions they have, and how strong are the forces of attraction between pairs of particles. This particle view of matter has easily explained the properties of solids, liquids and gases shown in Table 1.6. To explain other properties we need to consider the *actual* structure of the particles making up substances.

EXERCISES

- 1 How would you expect the motion of particles in:
 - a a solid
 - b a liquidto change as the temperature was increased?
- 2 A sample of gas is enclosed in a syringe with its tip sealed. If the plunger is pushed in (by exerting an increased pressure on it), how would you expect the number of particles per millilitre to change—to increase, decrease, or remain the same? If the syringe and the gas it contained were then heated, what would happen to the number

of particles per millilitre if the volume was kept constant? Why would it be necessary to apply a greater pressure to the syringe plunger to keep the volume constant as the gas was heated?

- 3 When two gases are brought into contact, they mix more quickly than two liquids do when brought together. Explain why, in terms of the particle nature of gases and liquids.

2.2 ATOMS AND MOLECULES

The particles that make up elements are called *atoms*.

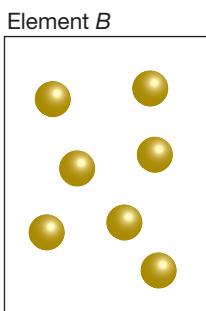
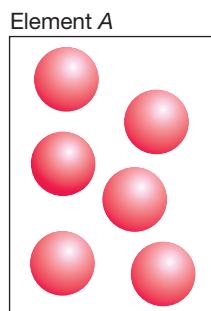
An **atom** is the smallest particle of an element which is still recognisable as that element.

This means that if we had a lump of the element copper and kept breaking it up into smaller and smaller pieces, eventually we would get to the stage where the individual pieces would be atoms of copper. It is possible to break these atoms into even smaller pieces but those pieces would no longer be copper (they would be sub-atomic particles we call electrons, protons, neutrons—Section 2.7).

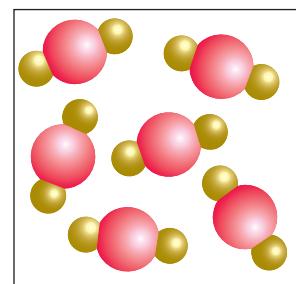
We envisage atoms as tiny spheres. All the atoms of one element are identical (except for some variation in mass for certain elements), but they are different from the atoms of all other elements.

Compounds consist of two or more elements combined together in definite proportions by mass. In *some* compounds the particles which make up the compound consist of several atoms joined tightly together. We call these particles *molecules*. The elements hydrogen and oxygen form the compound water. Water is made up of molecules. Each molecule of water has two atoms of hydrogen firmly joined to one atom of oxygen. All the molecules of water are identical. Figure 2.2 illustrates these ideas of atoms and molecules, while Figure 2.3 shows some common molecules. We define a molecule as follows:

A **molecule** is the smallest particle of a substance that is capable of separate existence.



The atoms of one element are all alike, but different from the atoms of all other elements.



Molecules are groups of atoms 'stuck' together.

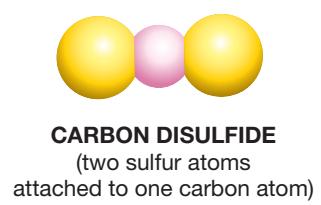
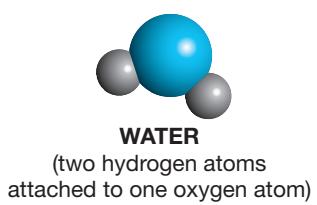


FIGURE 2.2
Atoms and molecules

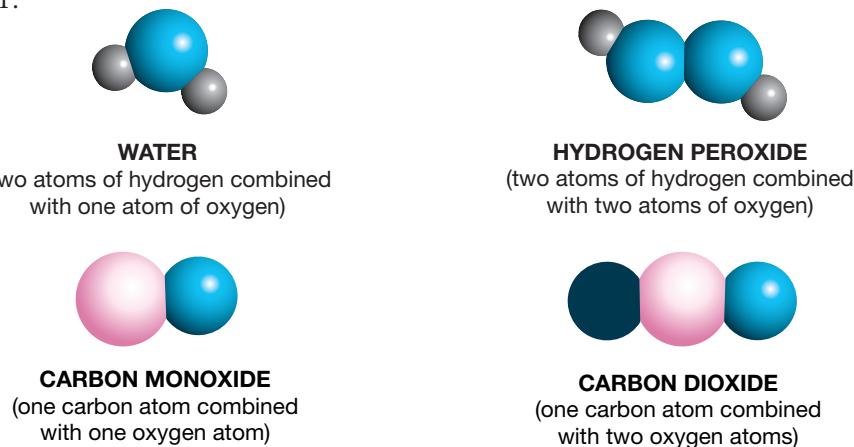
FIGURE 2.3
Some simple molecules

By this definition the entities shown in Figures 2.2 and 2.3 are clearly molecules. However, the definition has some consequences for elements as we shall see in Section 2.6.

Two elements may combine to form more than one compound as is shown in Figure 2.4. Even though two compounds are made from the same elements, they have quite different properties. Hydrogen peroxide bleaches hair; water does not. Carbon monoxide is poisonous; carbon dioxide is not.

Note that only *some* compounds exist as discrete molecules as described here. We shall look at the forms in which other compounds exist in Sections 2.15 and 2.21.

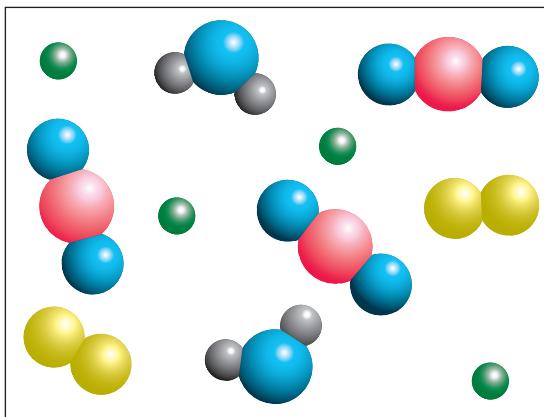
FIGURE 2.4
Pairs of elements forming
more than one compound



Particle views of elements, compounds and mixtures

Figure 2.2 gives a particle view of two elements and a compound. A particle view of a mixture is shown in Figure 2.5.

FIGURE 2.5
A particle view of a mixture



2.3 SIZES OF ATOMS AND MOLECULES

Atoms are typically 0.1 to 0.3 nanometres in diameter (1 nanometre or $1 \text{ nm} = 10^{-9}$ metre). This means that along the edge of a 30 cm steel ruler there are about 1 to 2 billion atoms side by side (1 billion = 10^9). Other indications of the small size of atoms are that a 5 cent coin has a mass of about 3 grams and contains about 3×10^{22} atoms, while a small grain of charcoal (carbon) with a mass of about 1 milligram contains 5×10^{19} atoms.

Molecules are of course a bit larger than atoms (since each molecule contains several atoms). A crystal of sugar (sucrose) with a mass of 10 milligrams

contains 2×10^{19} molecules. Since each molecule of sucrose contains 45 atoms (12 carbon atoms, 22 hydrogen atoms and 11 oxygen atoms), the crystal contains $45 \times 2 \times 10^{19} = 9 \times 10^{20}$ atoms.

If atoms were as big as marbles, oranges would be about the size of the Earth!

2.4 SYMBOLS FOR ELEMENTS

Chemists talk and write about elements so frequently that they have introduced a set of **symbols** to represent the elements. These symbols are a form of chemical shorthand. They fall into three categories:

- 1 Some symbols are simply the first letter of the name of the element (always a capital letter). These symbols are generally used for very common elements.

B for boron	C for carbon	F for fluorine	H for hydrogen
I for iodine	N for nitrogen	O for oxygen	P for phosphorus
S for sulfur	U for uranium		

- 2 Where there are several elements beginning with the same letter, a second letter is used, frequently though not always the second letter of the name of the element. The first letter is a capital (upper case) and the second a small (lower case) letter.

Ar for argon	Al for aluminium	As for arsenic
Ba for barium	Bi for bismuth	Br for bromine
Ca for calcium	Cl for chlorine	Co for cobalt
Ni for nickel	Ne for neon	Cr for chromium
Pt for platinum	Pu for plutonium	

- 3 For some elements we use a symbol derived from a non-English name of the element.

Na for sodium (natrium)	Ag for silver (argentum)
Cu for copper (cuprum)	K for potassium (kalium)
Au for gold (aurum)	Fe for iron (ferrum)
Hg for mercury (hydrargyrum)	Pb for lead (plumbum)

You must know the symbols of the common elements for they are part of the language of chemistry. In handwritten work the first letter of a symbol must be a *printed* capital (upper case) and the second letter a *written* lower case one as shown in Figure 2.6.

By hand we write Na, Fe, Hg, Cl, Br, As, Pt.
Co is cobalt; CO is the compound, carbon monoxide.

FIGURE 2.6
Symbols in handwritten text

2.5 FORMULAE

Symbols are used for elements. Compounds are made up of elements, so it is natural that combinations of symbols should be used to denote compounds. Such combinations of symbols are called **formulae**.

H_2O is the formula for water. It is however more than just shorthand. This formula tells us that:

- water is made of the elements hydrogen and oxygen
- the water molecule contains *two* atoms of hydrogen bound to *one* atom of oxygen.

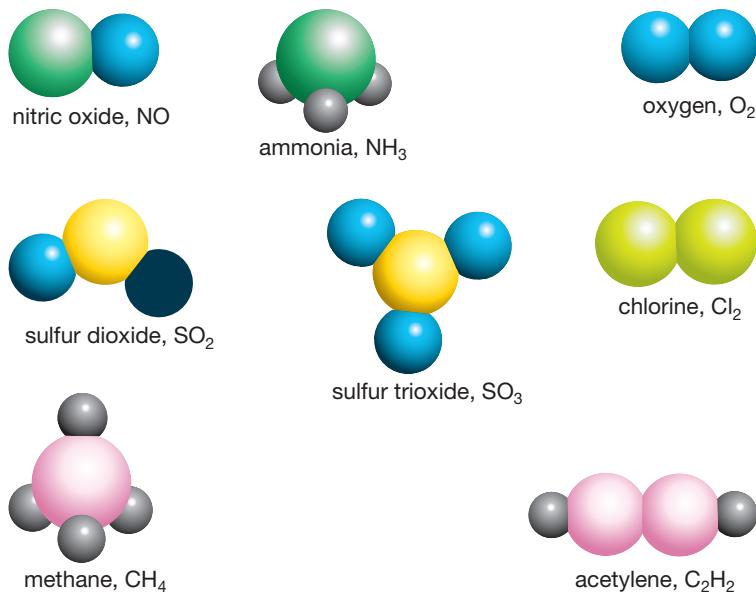
A subscript in a formula refers to the element immediately *preceding* it (as in water above). Another example is carbon dioxide, CO_2 : each molecule contains two atoms of oxygen and one atom of carbon. Other formulae are shown in Table 2.1.

TABLE 2.1 Some common formulae and their meaning

Name of compound	Formula	One molecule of the compound contains
ammonia	NH_3	3 atoms of hydrogen and 1 atom of nitrogen
sulfuric acid	H_2SO_4	2 atoms of hydrogen, 1 atom of sulfur and 4 atoms of oxygen
boric acid	$\text{B}(\text{OH})_3$	1 atom of boron, 3 atoms of oxygen and 3 atoms of hydrogen
		<i>Note carefully that the subscript 3 refers to all the atoms within the brackets. There is a similar situation in the next compound.</i>
dimethyl mercury	$(\text{CH}_3)_2\text{Hg}$	1 atom of mercury, 2 atoms of carbon and 6 atoms of hydrogen (2×3)

Figure 2.7 shows some common molecules with their formulae.

FIGURE 2.7
Some simple molecules with their formulae



2.6 MOLECULES OF ELEMENTS

Oxygen is an element. Its symbol is O. However in the atmosphere, oxygen gas is not present as separate independent atoms. Instead it exists as *pairs* of atoms permanently stuck together to form molecules. Or stated differently, in oxygen

gas the small particles moving about randomly are molecules, each molecule being a pair of atoms chemically bonded together. Hence for oxygen gas the formula is O_2 . It is a **diatomic molecule**.

This is true for all the common gaseous elements except the noble gases; the smallest particle that is able to exist on its own for long periods of time is a molecule containing two atoms. Nitrogen gas is N_2 , chlorine gas is Cl_2 , and hydrogen gas is H_2 . Two of these diatomic gaseous molecules are shown in Figure 2.7.

These molecules can be broken into separate atoms in chemical reactions as when oxygen gas reacts with hydrogen to form water, H_2O . The O_2 molecule splits into two separate O atoms for this to happen.

The noble gases, helium, neon, argon, krypton, xenon and radon, exist as independent atoms. By our definition of molecule in Section 2.2 these atoms are therefore molecules. We say that the noble gases consist of **monatomic molecules**.

EXERCISES

- 4** For each of the following compounds (for which the formula is given) state what kinds of atoms and how many of each kind are present in a molecule of the compound:
- a nitrous oxide, N_2O (used as an anaesthetic)
 - b sulfur dioxide, SO_2 (common air pollutant)
 - *c phosphoric acid, H_3PO_4 (used to acidify some foods)
 - *d ethanol, C_2H_6O ('ordinary' alcohol in beer and wine)
- 5** **a** Write the formulae of the compounds shown in Figures 2.3 and 2.4.
- b** Write the formulae of the following compounds. The molecules contain the atoms listed:
- i phosphine (1 phosphorus and 3 hydrogen atoms)
 - ii dinitrogen pentoxide (2 nitrogen and 5 oxygen atoms)
 - *iii urea (1 carbon, 1 oxygen, 2 nitrogen, 4 hydrogen atoms)
 - *iv sucrose (12 carbon, 22 hydrogen, 11 oxygen atoms)
- 6** State the total number of atoms in one molecule of:
- a chloric acid, $HClO_3$
 - b butane, C_4H_{10}
 - *c silicic acid, $Si(OH)_4$
 - *d dimethyldichlorosilane, $(CH_3)_2SiCl_2$
- 7** Give the names and formulae of:
- a three gases which are monatomic
 - b six gases that are diatomic (including two that are compounds)
 - c three gases that are triatomic
- *8** How many elements and how many compounds are represented in the mixture in Figure 2.5? How many atoms or molecules of each species are shown there?



The account of atoms and molecules just given is essentially an early nineteenth century one. While it is helpful in introducing molecules, symbols and formulae, it does not explain *why* atoms combine to form compounds or *how* they

do it or *why* some atoms form compounds much more readily than others. To answer these how and why questions we need to look first at the internal structure of atoms, then at the ways in which they bond together to form compounds.

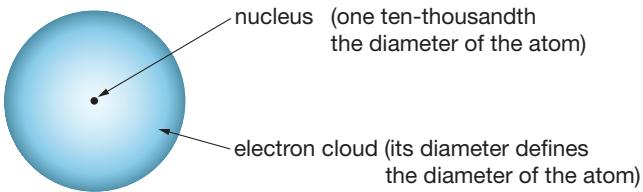
An understanding of chemical bonding and reactivity will lead us to a better appreciation of why there is such a diverse range of substances present in the Earth.

2.7 AN ATOM—A NUCLEUS AND AN ELECTRON CLOUD

An atom consists of an extremely small dense **nucleus** or core which contains the bulk of the mass of the atom and which carries positive electrical charges. This nucleus is surrounded by a ‘cloud’ of rapidly moving extremely light particles carrying negative electrical charges. These light negative particles are called *electrons*. This electron cloud constitutes the bulk of the volume of the atom, but contributes only very slightly to the mass of the atom. The amount of negative charge carried by these rapidly and randomly moving electrons is equal to the amount of positive charge on the nucleus, so that the atom is neutral overall. Figure 2.8 shows the nucleus and the surrounding electron cloud.

FIGURE 2.8

An atom (diameter 0.1 to 0.3 nanometres) consists of a nucleus and an electron cloud



The nucleus contains more than 99.95% of the mass of the atom, but its diameter is less than one ten-thousandth of the diameter of the whole atom.

An **electron** is an extremely small negatively charged particle with a mass of 1/2000 of the mass of a hydrogen atom.

Electrons in atoms move very rapidly and randomly around the nucleus. They do not (repeat, not) move in fixed orbits like planets around the sun, but rather they move so that their distance from the nucleus varies greatly; at one instant an electron may be close to the nucleus, the next instant far away from it. Because electrons are extremely small and because the volume available to them to wander around in is so huge, they are able to move about haphazardly without ever colliding. It is because we have many electrons moving randomly through the volume surrounding the nucleus that we use the term ‘**electron cloud**’ to describe it.

Here is an illustration of the relative sizes. If the nucleus of an atom were the size of a cherry stone placed at the centre of the Sydney Olympic stadium, electrons would be about the size of fruit flies and they would wander through a volume as big as the whole arena, spectator stands included! These fruit-fly electrons could be near the cherry stone, around the boundary of the field or even as far away as the back of the stands—anywhere in the whole volume. However electrons in atoms move much much faster than fruit flies in this arena—equivalent to moving from near the cherry stone to the back of the spectator stands in the shortest instant of time.

The nucleus is made up of two types of particles. One type carries a positive charge and is called a proton. The second type is neutral and is called a neutron.

A **proton** is a small positively charged particle having a mass approximately equal to the mass of a hydrogen atom and a charge equal in magnitude (but opposite in sign) to that of an electron.

A **neutron** is a small neutral particle which has the same mass as a proton.

The properties of electrons, protons and neutrons are summarised in Table 2.2.

As an atom is electrically neutral, *the number of electrons in the electron cloud is equal to the number of protons in the nucleus.*

Atoms of different elements contain different numbers of protons and neutrons in their nuclei and hence different numbers of electrons surrounding the nuclei.

TABLE 2.2 Properties of fundamental particles

	Symbol	Relative mass ^a	Relative charge ^b
electron	e	1/2000	-1
proton	p	1	+1
neutron	n	1	0

a on a scale where the proton is set arbitrarily at unity.

b The actual charge on an electron is 1.602×10^{-19} coulomb.

2.8 SOME SIMPLE ATOMS

The nucleus of a hydrogen atom consists of one proton only (and no neutrons), so there is one electron surrounding it. This is consistent with our earlier statements that the mass of an electron was 1/2000 of the mass of a hydrogen atom and that the mass of the proton was approximately that of the hydrogen atom.

A helium atom has two protons and two neutrons in its nucleus and two electrons surrounding the nucleus. Lithium has three protons and four neutrons in its nucleus and three electrons in its cloud.

Atoms of all the elements are obtained by successively increasing by one the number of protons in the nucleus. We cannot predict exactly how many neutrons are present, but for the simpler (lighter) elements, the number of neutrons is approximately equal to the number of protons. Table 2.3 shows the composition of the atoms of the first twenty elements. It shows, for example, that a sodium atom consists of a nucleus containing 11 protons and 12 neutrons and that there are 11 electrons in the electron cloud.

TABLE 2.3 Atomic structure of the first twenty elements

	Number of protons	Number of neutrons	Number of electrons	Mass number, A ^a
H	1	0	1	1
He	2	2	2	4
Li	3	4	3	7
Be	4	5	4	9
B	5	6	5	11
C	6	6	6	12

continued

TABLE 2.3 Atomic structure of the first twenty elements (continued)

	Number of protons	Number of neutrons	Number of electrons	Mass number, A^a
N	7	7	7	14
O	8	8	8	16
F	9	10	9	19
Ne	10	10	10	20
Na	11	12	11	23
Mg	12	12	12	24
Al	13	14	13	27
Si	14	14	14	28
P	15	16	15	31
S	16	16	16	32
Cl	17	18	17	35
Ar	18	22	18	40
K	19	20	19	39
Ca	20	20	20	40

^a Mass number, $A = \text{number of protons} + \text{number of neutrons}$.
Atomic number, $Z = \text{number of protons}$

2.9 ATOMIC NUMBER AND MASS NUMBER

Atomic number and mass number are two terms frequently used to describe atoms.

The **atomic number**, Z , of an element is the number of protons in the nucleus of an atom of that element.

From Table 2.3 we see that the atomic number of hydrogen is 1, of carbon, 6, of sodium, 11, and of calcium, 20.

The **mass number**, A , is the number of protons plus neutrons in the nucleus of an atom of the species concerned.

From Table 2.3 the mass number of helium is $2 + 2 = 4$; of carbon, $6 + 6 = 12$; of sodium, 23; and of calcium, 40.

EXERCISES

- 9 A sodium atom consists of 11 protons, 11 electrons and 12 neutrons. What is its atomic number and mass number?
- 10 Supply the missing numbers in the following table:

Element	Atomic number	Mass number	Number of protons	Number of neutrons	Number of electrons
helium	2	4			
oxygen	8			8	
boron		11			5
sulfur		32	16		
bromine				45	35
*iodine	53			74	
*argon	18	40			
*zinc		65	30		

- 11 Extend Table 2.3 to include the next six elements (mass number in brackets): Sc (45), Ti (48), V (51), Cr (52), Mn (55), Fe (56).

While the *numbers* of protons, neutrons and electrons in atoms is important, it is the way the electrons are *arranged* in the electron cloud that is of greater chemical significance.

2.10 ENERGY LEVELS FOR ELECTRONS

Since positive and negative charges attract each other, we might expect the electrons to be attracted to the nucleus and the atom to collapse into a tiny blob; that is, for the electron cloud to shrink into the nucleus. The fact that this does not happen indicates that the electrons possess energy—sufficient to resist the attraction towards the positive nucleus.

We find that the electrons in an atom exist in **discrete energy levels** which we call the first, second, third energy level and so on. Each electron in the first energy level has a certain constant amount of energy; each electron in the second energy level also has a fixed amount of energy but it is greater than that possessed by electrons in the first energy level. Similarly, the electrons in the third energy level have larger amounts of energy still.

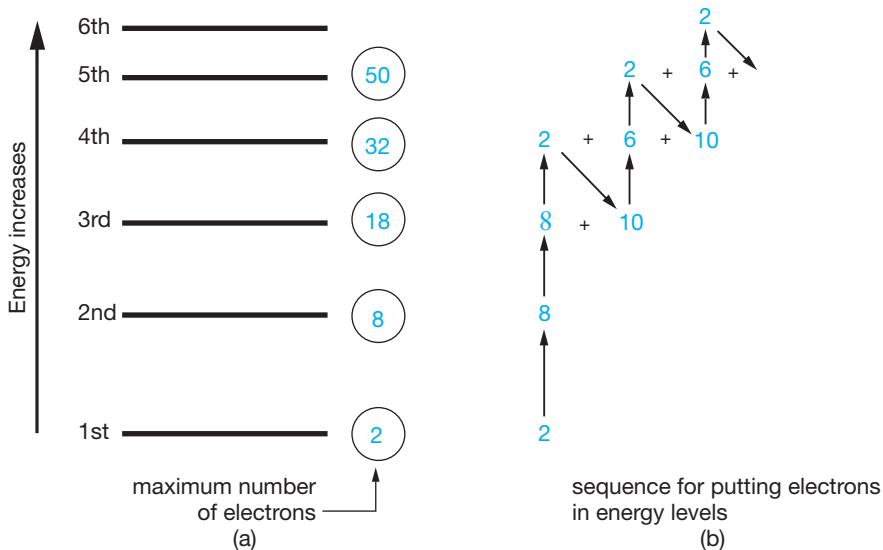
Each of these energy levels can accommodate only a certain maximum number of electrons as shown in Figure 2.9(a). The first energy level can accommodate only two electrons, the second can hold eight and the third eighteen; in general the n th energy level can accommodate $2n^2$ electrons. Electrons must be in one energy level or another—they cannot have energies that are intermediate between two levels.

The concept of electrons in energy levels is a bit like people in a multi-storey block of home units. The building consists of several floors (energy levels) and on each floor (level) we can accommodate a certain number of people (electrons). We can put fewer than this maximum number of people (electrons) on a particular floor (level) but we cannot exceed the maximum number. Just as people cannot live between floors (they must be on one floor or another), so too electrons cannot exist between energy levels; they must be in one or other of the levels. Actually it is an odd type of building: we can only accommodate 2 people on the first floor, but 8 on the second, 18 on the third, and 32 on the fourth! But like all analogies it should not be pushed too far.

In atoms, the electrons tend to be in the lowest energy levels possible consistent with not overfilling the levels.

FIGURE 2.9

- (a) Energy levels for electrons in atoms
(b) The sequence for filling them



2.11 ELECTRON CONFIGURATIONS

We work out the arrangement of electrons in any given atom by feeding the electrons into the energy levels starting at the lowest, and remembering that each level can hold only a certain maximum number.

For hydrogen (one electron only) we put the electron in the lowest energy level. For helium (two electrons) we put both electrons in the first level. Lithium has three electrons, two of which go into the first level with the third going into the second level. We write this as (2, 1). Beryllium with four electrons is (2, 2), and boron with five is (2, 3). Neon with ten electrons is (2, 8). The second energy level is now full (Figure 2.9(a)), so for sodium (eleven electrons) the last one goes into the third level to give (2, 8, 1).

The arrangement of electrons in energy levels is called the **electron configuration** of the atom. What is being described in this chapter is a simple form of electron configuration. It is perfectly adequate for our present purpose which is to understand chemical bonding. Table 2.4 presents the *simple form* electron configurations for the first 22 elements, plus a few more.

As we proceed from sodium (2, 8, 1) to argon which has 18 electrons, each additional electron goes into the third level, so that argon is (2, 8, 8).

As the third energy level can accommodate 18 electrons, we should expect potassium (19 electrons) to be (2, 8, 9), but it is not! *What we find is that after there are eight electrons in the third level, the next two go into the fourth level.* Potassium is (2, 8, 8, 1) and calcium (2, 8, 8, 2). This happens because these configurations are more stable than (2, 8, 9) and (2, 8, 10), although an explanation for them being more stable requires a more detailed picture of atomic structure than presented here.

Having put two electrons in the fourth level, we now go back and complete the filling of the third level. Scandium (21 electrons) is (2, 8, 9, 2), titanium (22 electrons) is (2, 8, 10, 2), and zinc (30 electrons) is (2, 8, 18, 2). The third level is now completely filled and further electrons go into the fourth level. Gallium (31 electrons) is (2, 8, 18, 3) and krypton (36 electrons) is (2, 8, 18, 8).

The sequence of feeding electrons into energy levels is shown in Figure 2.9(b).

Sometimes the term **energy shells** is used instead of energy levels. The two terms have exactly the same meaning.

TABLE 2.4 Electron configurations for simple elements

Element	Atomic number ^a	Electron configuration	Element	Atomic number	Electron configuration
hydrogen	1	1	chlorine	17	2, 8, 7
helium	2	2	argon	18	2, 8, 8
lithium	3	2, 1	potassium	19	2, 8, 8, 1
beryllium	4	2, 2	calcium	20	2, 8, 8, 2
boron	5	2, 3	scandium	21	2, 8, 9, 2
carbon	6	2, 4	titanium	22	2, 8, 10, 2
nitrogen	7	2, 5			
oxygen	8	2, 6			
fluorine	9	2, 7	zinc	30	2, 8, 18, 2
neon	10	2, 8	gallium	31	2, 8, 18, 3
sodium	11	2, 8, 1	germanium	32	2, 8, 18, 4
magnesium	12	2, 8, 2			
aluminium	13	2, 8, 3	bromine	35	2, 8, 18, 7
silicon	14	2, 8, 4	krypton	36	2, 8, 18, 8
phosphorus	15	2, 8, 5	rubidium	37	2, 8, 18, 8, 1
sulfur	16	2, 8, 6	strontium	38	2, 8, 18, 8, 2

^a number of electrons = number of protons = atomic number

EXERCISES

- 12 Without consulting any table or diagram, give the electron configuration of the following elements (atomic number in brackets): B (5), O (8), P (15), Ar (18), Ca (20), Mn (25), Se (34), Kr (36), Rb (37).

- 13 Atoms cannot have the following electron configurations:

- a** (1, 3) **c** (2, 9, 2) ***e** (2, 8, 14, 1) ***g** (3, 8, 8, 2)
b (2, 7, 4) **d** (2, 18, 8, 2) ***f** (2, 8, 20, 4) ***h** (2, 8, 8, 4)

For each, give the reason why not. Assuming that the total number of electrons is correct, suggest the correct electron configuration for each of these elements.



2.12 STABLE ELECTRON CONFIGURATIONS

When we set out the electron configurations of the atoms listed in order of increasing atomic number, as in Table 2.4, we can make some deductions about the relation between chemical reactivity and electron configurations.

First we notice that the so-called noble gases which (with a few minor exceptions) undergo no chemical reactions correspond to completely filled or semi-filled energy levels (semi-filled means filled to eight electrons). These gases are helium, neon, argon, krypton (and xenon and radon which are not shown): they are shaded pink in Table 2.4.

The electron configurations of these noble gases, then, must be extremely stable ones as the substances having them have virtually no tendency to undergo chemical reaction.

Our second major observation is that the so-called alkali metals, lithium, sodium, potassium, rubidium (and caesium which is not shown in Table 2.4), which all have very similar properties, all have electron configurations that have one more electron (in the next energy level) than the nearby noble gases. These elements are shaded yellow in Table 2.4.

The alkali metals all have similar chemical properties because they all tend to lose one electron to obtain the electron configuration of the nearby noble gas:

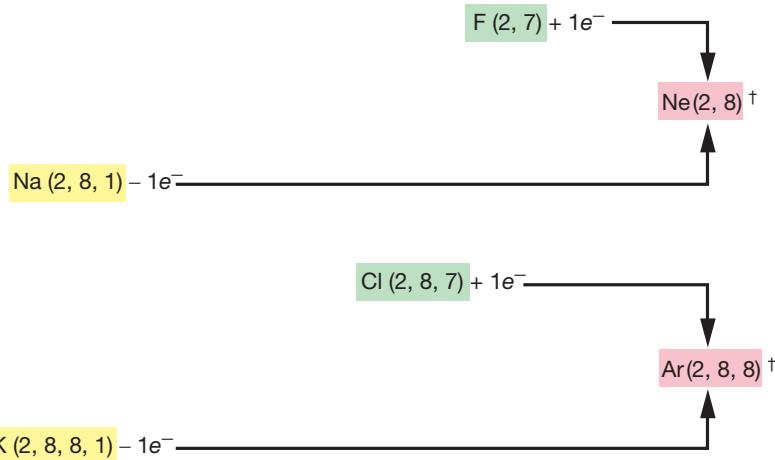
- *Lithium* tends to lose an electron and become like *helium*.
- *Sodium* tends to lose an electron and become like *neon*.
- *Potassium* tends to lose one electron and become like *argon* and so on.

A further observation is that the halogens, fluorine, chlorine, bromine (and iodine which is not shown in Table 2.4), which all have very similar properties, all have one electron less than the nearby noble gases. These elements are shaded green in Table 2.4.

The halogens have similar chemical properties because each tends to gain one electron to obtain the electron configuration of the nearby noble gas:

- *Fluorine* tends to gain an electron to become like *neon*.
- *Chlorine* tends to gain an electron to become like *argon* and so on.

Schematically:



From these and other similar observations we are able to make the following generalisation:

All elements, except the noble gases, tend to undergo chemical reactions and to form compounds in order to obtain the electron configurations of the nearby noble gases.

† These are just the electron configurations: the actual species will be positively or negatively charged, depending upon whether they were formed by the loss or gain of an electron.

This, then, is the ‘driving force’ behind chemical reactivity: *atoms tend to lose or gain electrons in order to become like the nearby noble gases*. Oxygen and sulfur tend to gain two electrons to become like neon and argon respectively (see Table 2.4). Magnesium and calcium each tend to lose two electrons to become like neon and argon respectively. Nitrogen and phosphorus tend to gain three electrons to become like neon and argon respectively.

A consequence of this general principle that atoms tend to become like the nearby noble gases is that a group of elements that differ from their nearby noble gases in the same way will have very similar chemical properties—they will form a *family of elements*. For example, beryllium, magnesium, calcium, strontium and barium all differ from the nearby noble gas by having two extra electrons. They all tend to lose these two electrons and therefore have very similar chemical properties.

Because of this tendency of atoms to gain or lose electrons to obtain noble gas configurations, it is the electrons in the highest energy level (outermost energy shell) that are involved in chemical bonding, as will be discussed in Sections 2.15 to 2.18.

Valence electrons

The electrons in the incompletely filled highest energy level (outermost energy shell) are called **valence electrons**. The outermost energy level is called the **valence shell**.

In magnesium (2, 8, 2) the valence electrons are the two in the third energy level. In fluorine (2, 7) the valence electrons are the seven in the second energy level; in silicon (2, 8, 4) they are the four in the third level. Since the noble gases have their highest energy levels filled or semi-filled to the stable configuration (8), we consider that they have no valence electrons.

The observation that there are families of elements with very similar properties leads to the concept of chemical periodicity and to the Periodic Table.

2.13 THE PERIODIC TABLE

In Section 1.25 the Periodic Table was introduced as a chart of the elements arranged so that elements with similar properties fell into the same vertical columns; these columns were called groups. We shall now consider how this table is constructed.

The Periodic Table is a chart in which the elements are listed in order of increasing atomic number, with elements having similar electron configurations (and therefore similar chemical properties) arranged to fall vertically underneath one another. Having ‘similar electron configurations’ means showing similar relationships to the nearby noble gases.

A typical Periodic Table is shown inside the front cover of this book. The noble gases, helium, neon, argon, krypton, xenon and radon, fall in a vertical column at the extreme right of the table. Those elements that have one electron fewer than these gases are the halogens—fluorine, chlorine, bromine and iodine; they fall in a column just to the left of the noble gases. The elements with one electron more than the noble gases are the alkali metals—lithium, sodium, potassium, rubidium and caesium; they fall in a column at the extreme left of the table.

Vertical columns of elements are called *groups*. The horizontal rows are called *periods*.

The **transition elements** or **transition metals** arise from converting a semi-filled level to a completely filled level (8 to 18 electrons). The sequence, scandium to zinc, is our first **transition series**.

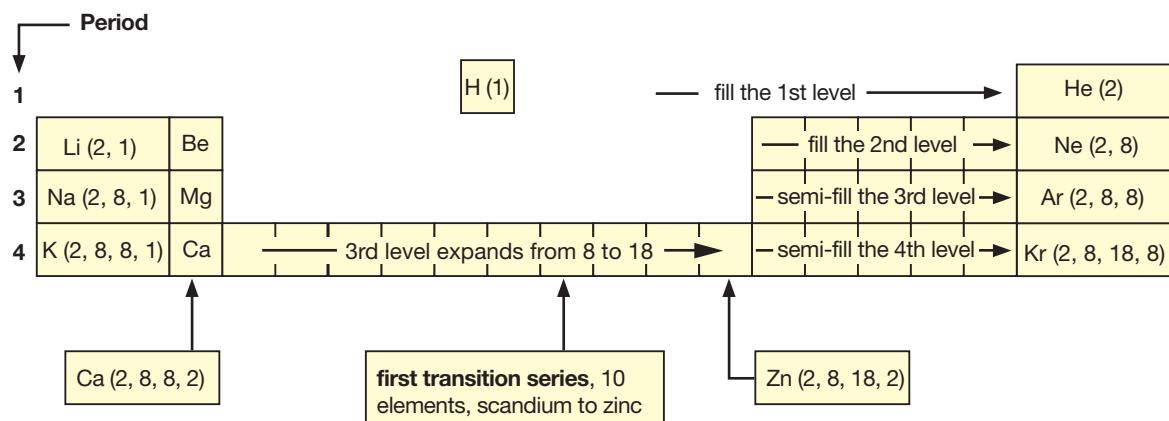
Hydrogen is often placed a little separate from the other elements in the Periodic Table because its chemical properties are quite different from those of members of any of the groups in the table. This arises because hydrogen can form stable electron configurations both by losing an electron to become a bare proton or hydrogen ion, H^+ , and by gaining an electron to become like helium. It is the only element with this option and therefore has unique properties.

As mentioned in Section 1.25, the value of the Periodic Table is that we no longer have 90 or so independent elements to study, but instead we have merely a small number of *groups* of elements with similar properties.

FIGURE 2.10
The relationship between the Periodic Table and the sequence of filling electron energy levels

The Periodic Table and electron configuration

The relationship between the layout of the Periodic Table and the filling of electron energy levels (shells) for the first four periods of the table is shown in Figure 2.10.



Valence electrons and position in the Periodic Table

Because of the way the Periodic Table has been constructed:

- elements in the same group have the same number of valence electrons
- the number of valence electrons is equal to the group number (if the groups are labelled as in this book; if the groups are un-numbered, as is usually the case in the HSC exam paper, it is the number of the column counting from the left, ignoring the transition metals and making the right-hand column (noble gases) zero).

Oxygen (2, 6) and sulfur (2, 8, 6) have six valence electrons; selenium and tellurium are in the same group, so they also have six valence electrons. Arsenic, antimony and bismuth are in the same group as nitrogen (2, 5) and phosphorus (2, 8, 5), so they also have five valence electrons.

These are very useful observations because they mean we do not have to work out the full electron configuration of an atom in order to determine how many valence electrons it has.

The Periodic Table is a basic tool of chemistry. Chemists always have it readily available and frequently refer to it because they can deduce much useful information about any particular element just by noting its position in the table. We shall take a closer look at the table and how it can be used in Chapter 6.

EXERCISES

14 Give the electron configuration of N (atomic number 7), Al (13), S (16), and Ca (20). For each of these elements decide whether the atoms tend to gain or lose electrons in reactions, and if so how many.

15 Underline the valence electrons in the configurations you wrote in Exercise 12.

Answer Exercises 16 and 17 without consulting a Periodic Table.

16 **a** Magnesium has two valence electrons; strontium and barium belong to the same group. How many valence electrons do strontium and barium have?
b Carbon (atomic number 6), germanium and lead are all in the same group. How many valence electrons do they each have?

17 Determine which element/s in Table 2.4 belong/s to the same group as:

- a** boron **b** carbon **c** nitrogen

Explain why.

18 Using the Periodic Table, how many valence electrons do the following elements have:

- a** tin (Sn) **b** iodine (I) **c** caesium (Cs) **d** indium (In)?

Explain your reasoning.

19 The electron configurations of some elements are A (2, 7), B (2, 8, 8, 2), C (2, 5), D (2, 8, 18, 8, 1), E (2, 8, 8). What configuration would you expect each of these elements to tend towards when they undergo chemical reaction? Which, if any, of these elements would you expect to be unreactive? Why?



Now that we have seen how electrons are arranged in atoms and have identified the ‘driving force’ behind formation of compounds—the drive towards noble gas configurations—we are ready to consider the nature of chemical bonding.

2.14 ACHIEVING NOBLE GAS CONFIGURATIONS

There are two ways in which atoms can obtain noble gas configurations:

- 1 by *outright transfer* of electrons from one atom to another to form what are called *ions*, positively or negatively charged particles; this results in what is called *ionic bonding*
- 2 by *sharing* electrons between pairs of atoms with each atom ‘considering’ that it ‘owns’ all the shared electrons; this is called *covalent bonding*.

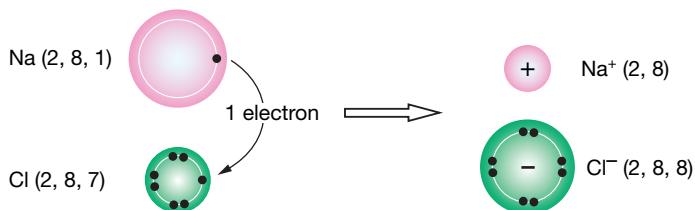
2.15 FORMATION OF IONS

The formation of ions and the ionic bonding it leads to are best explained by considering specific examples.

EXAMPLE 1

Sodium and chlorine combine to form the compound, sodium chloride.

The sodium atom, with electron configuration (2, 8, 1) (see Table 2.4), has a tendency to lose one electron to become like neon (2, 8); the chlorine atom, with electron configuration (2, 8, 7), tends to gain one electron to become like argon (2, 8, 8). Hence one electron is transferred from a sodium atom to a chlorine atom. When the neutral sodium atom loses one electron it becomes positively charged: we call it a *positive ion*. When the neutral chlorine atom gains an electron, it becomes negatively charged: it is a *negative ion*. We call it the *chloride ion*. Schematically:

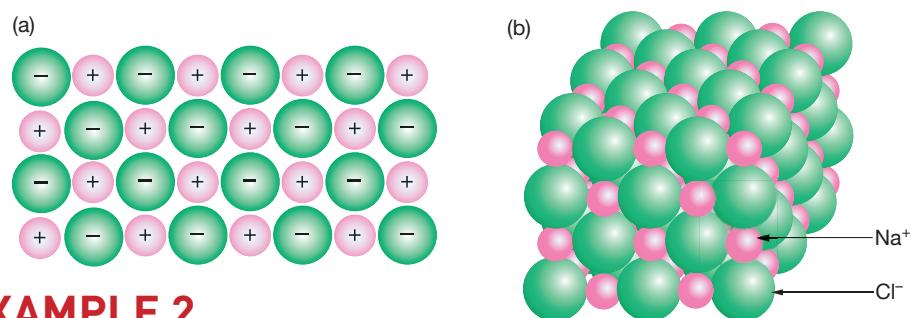


Ions are positively or negatively charged particles. The simplest ions form when atoms lose or gain electrons (as with sodium and chlorine here).

There is strong electrostatic attraction between positive and negative ions. This is what holds the ions together in ionic bonding. Sodium chloride is therefore NaCl (one sodium atom per chlorine atom, because that is the way to balance charges). We occasionally write it as Na^+Cl^- to stress the ionic nature of the bonding.

Crystals of sodium chloride consist of sodium ions and chloride ions packed in an orderly fashion. The electrostatic attraction between pairs of oppositely charged ions extends throughout the whole crystal. This is shown in Figure 2.11. There are no separate molecules of NaCl —just an extremely large array of positive and negative ions held tightly together.

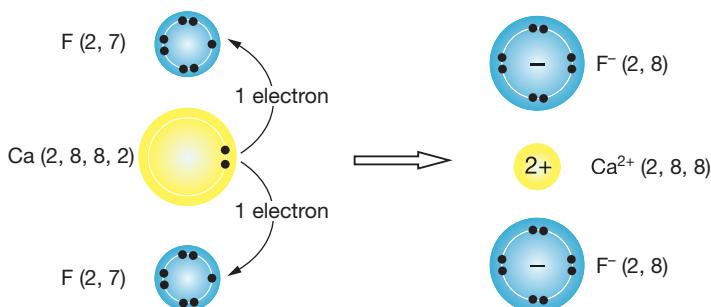
FIGURE 2.11
Ionic substances consist of orderly arrays of positive and negative ions:
(a) a two-dimensional array of ions,
(b) a portion of a sodium chloride crystal



EXAMPLE 2

Calcium and fluorine combine to form calcium fluoride.

Each calcium atom, with electron configuration (2, 8, 8, 2) (Table 2.4), loses two electrons to become like argon (2, 8, 8), while each fluorine atom (2, 7) gains one electron to become like neon (2, 8). Hence two fluorine atoms must combine with one calcium atom to form the ionic compound, CaF_2 :



To emphasise the ionic nature of the bonding we sometimes write it as $\text{Ca}^{2+}(\text{F}^-)_2$ or $\text{Ca}^{2+}2\text{F}^-$.

Ionic bonding is that type of chemical bonding which involves the outright transfer of electrons from one atom to another. The bonding consists of electrostatic attraction between the positive and negative ions formed by this transfer of electrons.

Note carefully that in ionic compounds there are no discrete molecules—just an infinite array of positive and negative ions.

Meaning of the formula of an ionic compound

For ionic compounds, the formulae (for example NaCl , CaF_2) specify the *ratios* in which the atoms (or ions) are present, *not* the composition of discrete molecules. Such formulae that give the ratio by atoms of elements in a compound rather than the actual numbers of atoms in a molecule are called **empirical formulae**. Formulae for ionic compounds are therefore always empirical formulae (because there are no molecules).

Cations and anions

Positive ions are called **cations**. Negative ions are called **anions**.

2.16 IONIC BONDING AND THE PERIODIC TABLE

Elements that can attain noble gas electron configurations by losing or gaining one or two electrons commonly form ions. This means that we can use the Periodic Table to predict which elements will form ions—generally those that are only one or two elements away from a noble gas. In particular:

- 1 *Group 1 metals* (Li , Na , K , Rb , Cs) all tend to lose one electron and therefore *form singly charged positive ions*: Li^+ , Na^+ , K^+ , Rb^+ , Cs^+ .
- 2 *Group 2 metals* (Be , Mg , Ca , Sr , Ba) tend to lose two electrons and therefore *form doubly charged positive ions*:[†] Be^{2+} , Mg^{2+} , Ca^{2+} , Sr^{2+} , Ba^{2+} .
- 3 *Group 7 elements* (non-metals, F , Cl , Br and I) all tend to gain one electron and therefore they *form singly charged negative ions*: F^- , Cl^- , Br^- , I^- .
- 4 *Group 6 elements* (non-metals, O , S , Se , Te) tend to gain two electrons and thus *form doubly charged negative ions*: O^{2-} , S^{2-} , Se^{2-} , Te^{2-} .

However while Groups 1 and 2 form *only* ionic compounds, as we shall see below, Groups 6 and 7 also form covalent compounds.

- 5 The *transition metals* all lose electrons to form positive ions (for example Cr^{3+} , Fe^{2+} , Cu^{2+} , Ag^+ , Zn^{2+}), but it is not possible from a simple look at the Periodic Table to predict just how many electrons any particular atom will lose.
- 6 Elements that are three places away from a noble gas *may* form ions: aluminium generally does (Al^{3+}), boron never does, nitrogen and phosphorus do so on rare occasions.

Another generalisation is that *metals generally form positive ions while non-metals (when they form ions) form negative ions*.

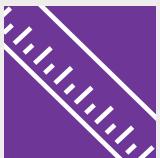
[†] Beryllium is exceptional in that it forms some covalent compounds.



WEBSITE

<http://www.chem4kids.com>

(select atoms; the bit on ions is particularly good, though the whole section is worth reading—a very simple introduction to atomic structure and ionic bonding)



EXERCISES

- 20** Show (in terms of electron configurations of the atoms involved and the ions formed) how the following pairs of elements form ionic compounds. Deduce the formulae of the compounds formed:

a potassium and fluorine ***c** sodium and oxygen
b calcium and bromine ***d** magnesium and oxygen

21 Explain why:

a potassium bromide is KBr and not K_2Br or KBr_2
***b** barium chloride is $BaCl_2$ and not Ba_2Cl or $BaCl_3$

22 Which of the following elements would you expect to form positive ions and which negative ions and which not form ions at all? What charge would you expect each ion to carry? Si, Rb, I, Mg, S, Ga, K, C, Br, Ar, Se, B, P, Ba, Cs, F

23 a Give the electron configuration of the following atoms and ions, taking atomic numbers from the Periodic Table if necessary:

i magnesium atom, magnesium ion
ii sulfur atom, sulfide ion
***iii** chlorine atom, chloride ion
***iv** potassium atom, potassium ion

b What do you expect for the formula of each of the following compounds?

i magnesium sulfide **iii** magnesium chloride
***ii** potassium chloride ***iv** potassium oxide

2.17 COVALENT BONDING

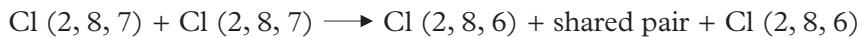
Covalent bonds are formed between pairs of atoms by the atoms sharing electrons.

Again the nature of this type of bonding is best explained by considering specific examples.

EXAMPLE 3

Two chlorine atoms combine to form a chlorine molecule.

A chlorine atom, with electron configuration (2, 8, 7), tends to gain one electron to acquire the configuration of argon (2, 8, 8). Two chlorine atoms can combine to form a chlorine molecule, Cl_2 , by sharing a pair of electrons with each atom contributing one electron to the shared pair. Each atom ‘considers’ that it ‘owns’ the shared pair and thus counts both members of the pair to determine that it has a noble gas configuration.



This shared pair of electrons occupies a volume of space that surrounds both atoms. By moving around both nuclei these electrons hold the atoms together and so form a **chemical bond**. This is illustrated in Figure 2.12. The rest of the electrons (that is, the unshared ones) remain in volumes of space that surround one nucleus only.

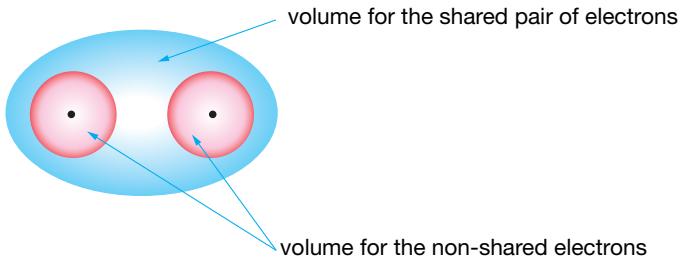


FIGURE 2.12
In a covalent bond the shared pair of electrons occupies a volume of space that surrounds both nuclei

Each chlorine atom in the chlorine molecule now ‘considers’ that it has the configuration (2, 8, 8) and is therefore ‘satisfied’. This bonding which results from sharing pairs of electrons is called covalent bonding. Each pair of shared electrons is called a covalent bond and is denoted by a dash. Molecular chlorine has the structure:

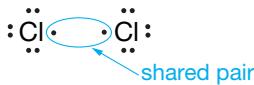


We call it a **covalent molecule**.

A common way of describing covalent bonding is to use what are called **electron-dot structures** (or **diagrams**). These are symbols or formulae of elements or compounds with the valence electrons (outer energy level electrons) shown as dots. The electron-dot formula for the chlorine atom is:



(The completely filled first and second main levels—the 2, 8—are ignored and just the seven valence electrons are shown.) We can then denote the pairing up of electrons by drawing:



which we write as Cl—Cl (above).

A dash equals one shared (bonding pair) of electrons or one covalent bond.

EXAMPLE 4

Hydrogen and chlorine combine to form the gaseous compound, hydrogen chloride. The hydrogen atom, with only one electron, needs to gain another electron to become like helium. Chlorine (2, 8, 7) also needs to gain one electron to become like argon (2, 8, 8). Hence a covalent bond is formed by each atom contributing one electron to a pair which is shared by both atoms. In this way both atoms ‘consider’ that they ‘own’ the pair and that they have achieved the desired noble gas configuration. The covalent molecule, HCl, has been formed. We can write it as H—Cl. Its electron-dot structure is:



EXAMPLE 5

Hydrogen and oxygen combine to form water.

Oxygen (2, 6) needs to gain two electrons to become like neon (2, 8). Hydrogen needs to gain one electron to become like helium. Hence covalent bonds are formed between one oxygen atom and two hydrogen atoms to form the covalent molecule water:



In this way each hydrogen atom and the oxygen atom ‘consider’ that they have achieved the noble gas configuration (because each atom ‘counts’ both electrons in the shared pair).

The actual appearances of the three covalent molecules in these examples is shown in Figure 2.13. While strong covalent bonds bind the atoms together within each molecule, the molecules are able to move about more or less independently of one another as described in Section 2.1.

FIGURE 2.13
Simple covalent molecules



Substances (elements or compounds) that are made up of simple covalent molecules as in Figure 2.13 are called **covalent molecular substances** or sometimes just *molecular substances* (to distinguish them from *covalent lattices* which will be treated in Section 2.23).

2.18 COVALENCY AND THE PERIODIC TABLE

Covalent bonding occurs when *both* of the elements forming the compound need to gain electrons to attain noble gas configurations—in the above three examples, chlorine and chlorine, hydrogen and chlorine, hydrogen and oxygen.

Elements in the centre and to the right of the Periodic Table tend to form covalent compounds—elements such as carbon, silicon, nitrogen, phosphorus, oxygen, sulfur, fluorine, chlorine (though the last four commonly form ionic compounds also).

The number of covalent bonds an atom forms is the number of electrons that atom needs to gain to acquire a noble gas configuration—one for hydrogen and chlorine, two for oxygen (as in the above examples), three for nitrogen and four for carbon.

The position of an element in the Periodic Table tells us how many electrons the atom needs to gain to attain a noble gas configuration and so it tells us how many covalent bonds the atom will form.

Nitrogen (Group 5) needs another three electrons so it forms three covalent bonds (for example in ammonia, NH_3). Carbon (Group 4) needs another four electrons so it forms four covalent bonds (for example in methane CH_4).

2.19 ELECTRON-DOT STRUCTURES FOR IONS

While electron-dot structures are particularly useful for describing the bonding in covalent compounds, they can be used for ions as well. For example the chlorine and sulfur atoms and the chloride and sulfide ions would be:



The electron-dot structure for a simple metal ion will not have any dots (for electrons) because these ions have no valence electrons:



The positive or negative charges must be included on electron-dot structures of ions.

2.20 DRAWING ELECTRON-DOT STRUCTURES

The steps for drawing an electron-dot structure for a compound containing two elements are:

	Example: Li and S	Example: O and Cl
1 Determine the number of valence electrons on the atom of each element (by using group number or by counting along the relevant period of the Periodic Table, ignoring transition elements) and draw this number of dots around the symbol for each element.	Li · ·S· ..	:Cl· ·O· ..
2 Decide how many electrons each atom needs to gain or lose to obtain a noble gas configuration.	Li needs to lose one e ⁻ (to become like He). S needs to gain two e ⁻ (to become like Ar).	O needs to gain two e ⁻ (to become like Ne). Cl needs to gain one e ⁻ (to become like Ar).
3 Decide whether the compound will be ionic or covalent: if one atom wants to gain electrons while the other wants to lose, the compound will be ionic; if both want to gain electrons then the compound will be covalent	Compound will be ionic.	Compound will be covalent.
4 For an ionic compound		
a Work out the number of atoms of each type that must combine so that the total number of electrons lost equals the total number gained.	Need two Lis for each S (so total number of e ⁻ s lost and gained is two).	
b Draw the atomic symbols with the electrons transferred.	Li :S: Li ..	
c Put the appropriate charges on the symbols; each electron lost produces one positive charge and each electron gained produces one negative charge. You now have the electron-dot structure for the compound.	Li ⁺ :S: ²⁻ Li ⁺ ..	
d If asked for it, write the formula for the compound.	Li ₂ S	
5 For a covalent compound		
a Work out the number of atoms of each type needed for all the atoms to obtain noble gas configuration. ^a		To gain the required two e ⁻ s O needs to share electrons with two Cls; but Cl only needs to share with one O so we need two Cls for each one O.
b Pair up electrons to give each atom a noble gas configuration. This is the required electron-dot structure.		:Cl· ·O· ·Cl:
c If a structural formula is required, draw it by using a dash for each shared pair of electrons (each covalent bond).		Cl—O—Cl
d If a molecular formula is required, write it.		Cl ₂ O

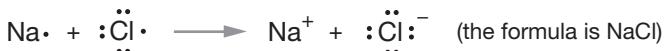
^a There are some atoms, such as B in BCl_3 , that form stable compounds with fewer electrons than the expected noble gas configuration, while there are others, such as P in PCl_5 and S in SF_6 , that can expand their valence shell beyond the eight electrons of the usual noble gas configuration. However, we can ignore these exceptions for HSC purposes.

2.21 IONIC EQUATIONS WITH ELECTRON-DOT STRUCTURES

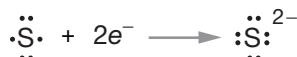
We can represent the formation of ions from metal and non-metal atoms by using equations involving electron-dot structures. For *sodium* and *chloride* ions:



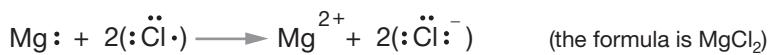
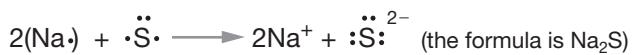
where e^- and \cdot each represent an electron. For the formation of the compound sodium chloride we can add these equations together to get:



Similar equations for the formation of sulfide and magnesium icons are:



To make equations for the formation of sodium sulfide and magnesium chloride, the appropriate ion-formation equations are added (after doubling one equation in each pair in order to balance out the electrons):



EXERCISES

24 Draw electron-dot diagrams and give molecular formulae for the covalent molecules formed between:

- a hydrogen and iodine
b two bromine atoms

- *c nitrogen and hydrogen
*d sulfur and fluorine

25 Explain why:

- a the compound formed between fluorine and oxygen is F_2O and not FO_2 or FO
*b ammonia is NH_3 and not NH_2 or NH_4

26 Decide the formula you would expect, and give your reasons, for the compound formed between:

- a sulfur and chlorine
b hydrogen and iodine

- *c silicon and hydrogen
*d phosphorus and fluorine

27 a Draw electron-dot structures for the bromide ion and the oxide ion.

- *b In a very small number of compounds nitrogen exists as the nitride ion. Draw an electron-dot structure for this ion.
c Draw electron-dot structures for barium and caesium atoms and the ions they form.

28 Draw electron-dot diagrams (in the form of equations) that show the formation (from atoms)

- a of the following ions:
i Ca^{2+} ii O^{2-} iii Al^{3+} iv F^-
b of the following compounds:
i calcium oxide ii aluminium fluoride

2.22 PROPERTIES OF COVALENT MOLECULAR AND IONIC SUBSTANCES

Table 2.5 contrasts the common properties of covalent molecular and ionic substances. Note that ionic substances are always compounds, whereas covalent molecular substances can be elements (e.g. iodine, phosphorus and sulfur) as well as compounds.

TABLE 2.5 Contrasting properties of covalent molecular and ionic substances

Ionic substances	Covalent molecular substances
<i>solids at room temperature</i>	at room temperature, generally gases (N_2 , SO_2 , NH_3) or <i>liquids</i> (H_2O , CCl_4 , methanol CH_3OH); a few are solids (I_2 , PCl_5 , CBr_4)
<i>high melting points</i> (typically above $400^\circ C$) and <i>high boiling points</i> (typically over $1000^\circ C$)	<i>low melting points</i> (generally below $200^\circ C$) and <i>low boiling points</i> (generally below $400^\circ C$)
<i>hard and brittle</i>	when solid they are <i>soft</i>
as solids <i>they do not conduct electricity</i>	pure covalent substances <i>do not conduct electricity</i> either as solids or as liquids
when molten or when in aqueous solution <i>they do conduct electricity</i> .	In aqueous solution <i>do not conduct electricity</i> (unless they actually react with water to form ions)

These properties can be explained by the nature of the two types of bonding.

Properties of ionic substances

Melting an ionic solid means breaking up the orderly arrangement of ions shown in Figure 2.11. As the electrostatic forces between ions are strong, much energy (and therefore a high temperature) is needed to do this. Boiling an ionic substance means producing a vapour that consists of well-separated ion pairs. This requires an even greater amount of energy and so an even higher temperature is needed.

The strong electrostatic attraction between pairs of ions makes ionic substances hard. If the orderly array of ions is disturbed by applying a strong force, then ions of the same charge come close together as in Figure 2.14. They then repel each other and this causes the crystal to shatter. This means that ionic crystals are brittle.

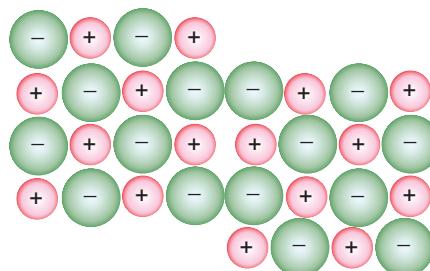


FIGURE 2.14 Ionic substances are brittle because distorting the crystal brings like charges together and they repel one another: this causes the crystal to break

Solid ionic compounds do not conduct electricity because in the solid the ions are tightly bound into an orderly array and so are unable to move

towards a charged electrode. However when ionic substances melt, the orderly arrangement of ions is broken up and the ions can move about relatively freely. They can then migrate towards a charged electrode; so molten ionic substances conduct electricity.

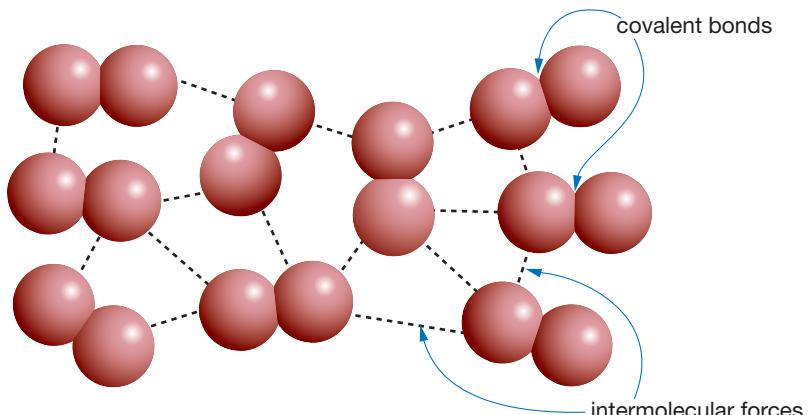
Similarly, when ionic substances are dissolved in water, the crystals are completely broken up and the solutions consist of individual ions moving randomly about through the water. These freely moving ions can migrate towards oppositely charged electrodes; so solutions of ionic substances conduct electricity.

Properties of covalent molecular substances

While the bonding forces holding atoms together *within* a covalent molecule are very strong, the forces between one molecule and its neighbours are quite weak. These weak forces between pairs of molecules are called **intermolecular forces**. Their nature will be discussed in Sections 7.5 and 7.6.

Figure 2.15 shows a sample of bromine liquid. There is a strong covalent bond holding the two bromine atoms together within each bromine molecule, Br_2 . There are only weak intermolecular forces (the dotted lines) between pairs of bromine molecules.

FIGURE 2.15
In liquid bromine there are weak intermolecular forces (dotted lines) between each bromine molecule and its neighbours. Boiling overcomes these relatively weak forces; it does not break the strong covalent bonds holding pairs of bromine atoms together as molecules



Boiling involves separating molecules from one another. This means that *boiling overcomes intermolecular forces: it does not break any covalent bonds*. Because the intermolecular forces holding the bromine molecules to one another in the liquid are weak, bromine has a relatively low boiling point of 58°C; that is, not much energy is required.

Similarly in solid bromine, the intermolecular forces holding the molecules to one another are relatively weak (compared with covalent bonds). Melting just involves disrupting the orderly arrangement of molecules; that is, melting, like boiling, only overcomes weak intermolecular forces. Therefore bromine has a relatively low melting point (-7°C).

Bromine is typical of all covalent molecular substances; there are only weak forces holding the molecules to one another and so they have low melting and boiling points.

However the strength of the intermolecular forces does vary from one molecular compound to another. We find that:

the stronger the intermolecular forces in molecular compounds, the higher are the melting and boiling points.

Because these weak intermolecular forces are easily overcome, it is easy to distort a solid covalent molecular substance. This means that such solids are soft.

Because covalent molecules are neutral species, they cannot conduct electricity either as pure substances or in solution (for example iodine, sucrose or urea in water). However, some covalent substances when mixed with water actually react with the water and form ions; these solutions do conduct electricity. Examples are hydrogen chloride and sulfuric and nitric acids which will be discussed in Section 7.12.

EXERCISES

- 29** Phosphorus trichloride is a liquid with a boiling point of 74°C; it does not conduct electricity. Calcium chloride is a solid with a melting point of 772°C; when molten it conducts electricity. Explain, in terms of bonding, why these two compounds have such different properties.
- 30** The boiling points of the fluorides of the elements of Period 2 of the Periodic Table are: LiF (1720°C), BeF₂ (1175°C), BF₃ (-101°C), CF₄ (-128°C), NF₃ (-120°C), OF₂ (-145°C) and F₂ (-188°C). Explain why the boiling points of lithium and beryllium fluorides are so much higher than the others.
- ***31** Calcium and arsenic both form hydrides. Calcium hydride is a solid with a melting point of 600°C, while the hydride of arsenic, called arsine, is a gas at room temperature.
- What do these facts imply about the bonding in the two compounds?
 - What do you expect the formulae of these compounds to be?



Other types of solids

In addition to ionic solids and covalent molecular solids there are two other types of solid, namely:

- metals, and
- covalent network solids (or covalent lattice solids or just covalent lattices).

2.23 COVALENT NETWORK SOLIDS

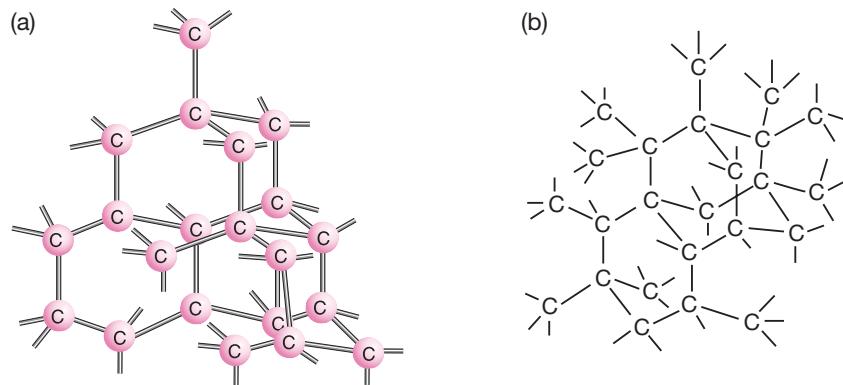
Covalent network solids are solids in which the covalent bonding extends indefinitely throughout the whole crystal.

Covalent network solids are sometimes called **covalent lattice solids** or just **covalent lattices**.

The word **lattice** is used to mean an infinite orderly array of particles. Ionic solids are sometimes called **ionic lattices**.

Carbon in the form of diamond is an example of a covalent network solid. As we expect from the Periodic Table, each carbon atom is covalently bonded to four other atoms, and since we are talking about an element it can only be to four other carbon atoms. This is shown in Figure 2.16: (a) is an exploded view of part of the crystal, showing atoms as spheres and using lines to represent the covalent bonds. The lines have no physical reality: in the actual diamond crystal the atoms are overlapping one another. (b) is the structure chemists usually draw for diamond: each dash or line represents a covalent bond. Literally billions of atoms are bonded together in this way to form what we call a covalent network solid.

FIGURE 2.16
The structure of diamond

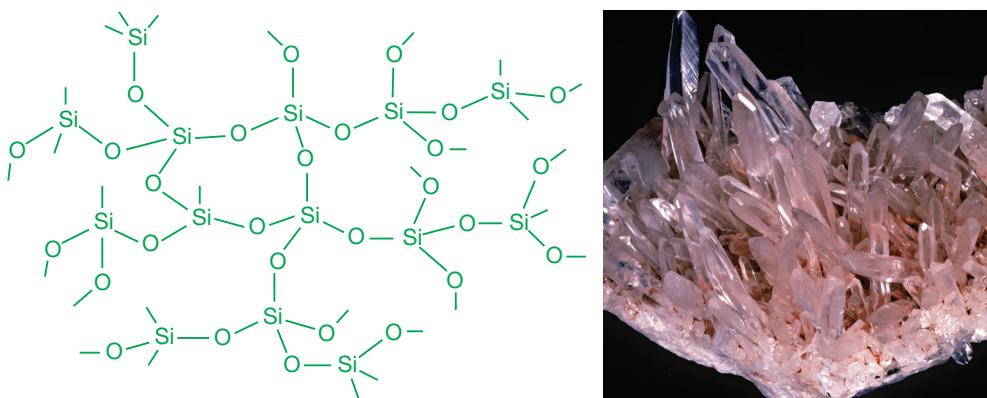


Silica (silicon dioxide or quartz, SiO_2) is another example. From the Periodic Table we see that silicon wants to form four covalent bonds while oxygen wants to form two. Therefore in silica each silicon atom is covalently bonded to *four* oxygen atoms, and each oxygen atom is bonded to *two* silicon atoms to form an infinite network of covalent bonds as shown in Figure 2.17.

As with ionic lattices (Section 2.15), the chemical formula for a covalent lattice compound represents the ratio in which the atoms are present in the compound. Hence SiO_2 for quartz. Because this formula represents a ratio, it is an empirical formula (page 51).

When covalent lattices melt, the one big covalent crystal breaks up into many smaller pieces, which are able to move about relative to one another. Hence melting covalent lattices involves breaking many covalent bonds that are very strong. This process requires a lot of energy and thus only occurs at high temperatures.

FIGURE 2.17
The structure of silica,
 SiO_2 . To its right, quartz



Covalent lattices have extremely high melting points, typically well above 1000°C (diamond $> 3550^\circ\text{C}$, silica 1700°C).

Except for graphite (discussed in Chapter 9) covalent network solids do not conduct electricity. This is because they do not contain any ions and all the electrons are tied up either being held by individual atoms or shared by pairs of atoms; none are free to roam around as in metals (next section).

Elements as covalent molecules or covalent lattices

Several elements exist as covalent molecules:

- H_2 , F_2 , Cl_2 , O_2 and N_2 are diatomic gases.
- Br_2 is a diatomic liquid while I_2 is a diatomic solid.

- Phosphorus and sulfur exist as covalent P_4 and S_8 molecules respectively.
- Some other elements exist as covalent lattices:
- Carbon exists as diamond which is a three-dimensional lattice and as graphite which is a two-dimensional lattice.
- The semi-metals B, Si, Ge, As, Sb and Te closely approximate to covalent lattices though their bonding electrons are not as firmly localised as in diamond.

We have already seen that the noble gases (He, Ne, Ar, Kr, Xe, Rn) exist as monatomic molecules; that is, without any chemical bonding at all.

The majority of the elements, however, display metallic bonding.

Covalent network solids in the Earth

Many substances in the lithosphere are covalent network solids:

- Sand and quartz are silicon dioxide.
- Some gemstones (amethyst and citrine) are silicon dioxide with traces of impurities which provide the colour, while others (emerald, aquamarine, topaz and garnet) are silicates or alumino-silicates which are covalent lattices with some ionic bits incorporated.
- Mica, talc and asbestos are also silicate lattices.
- Clays and zeolites are alumino-silicate lattices, again with some ionic portions.



Most gemstones are lattice compounds. Ruby and sapphire are ionic lattices, being corundum, a particular crystal form of aluminium oxide (with impurities to give the colours) while citrine is quartz (with iron impurity), completely covalent as is amethyst (not shown). Emerald and aquamarine are alumino-silicates (silicates with some Si atoms replaced by Al).

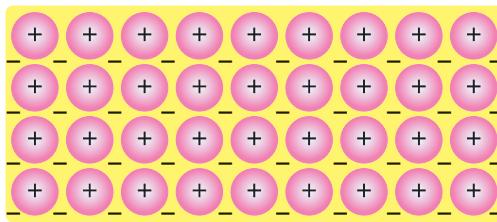
2.24 METALLIC BONDING

Metals, with the exception of mercury, are solids at room temperature. Most have relatively high melting points and are fairly hard. They are all good conductors of electricity. Properties of metals were described in Section 1.23.

These properties arise because a metal consists of an orderly three-dimensional array of positive ions held together by a mobile ‘sea’ of **delocalised electrons**. The valence electrons break away from their atoms, leaving behind positive ions. These free electrons, called *delocalised* because they no longer belong to particular atoms, move randomly through the lattice and, by being shared by numerous positive ions, provide the chemical bonding that holds the crystal together. This is shown in Figure 2.18.

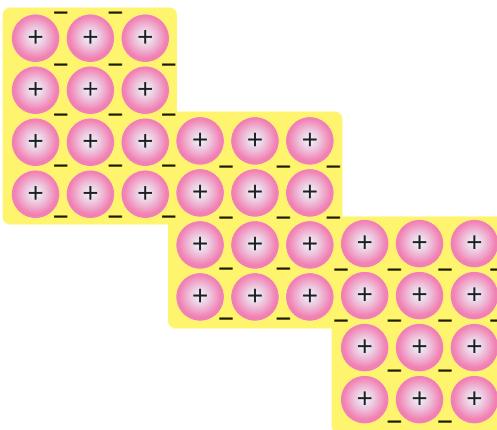
It is the ability of these delocalised electrons to move freely through the lattice that causes metals to be good conductors of electricity. Electric current through a metallic wire is a flow of electrons.

FIGURE 2.18
A metal is a three-dimensional lattice of positive ions immersed in a ‘sea’ of delocalised electrons. For clarity just a two-dimensional ‘slice’ is shown here



Metals can be bent, rolled into sheets, and drawn into rods and wires. These processes are possible because when the orderly array of positive ions is sheared, the mobile electrons are able to adjust to the new arrangement of positive ions and again provide the ‘glue’ to hold the whole assembly together as shown in Figure 2.19. This contrasts with ionic or covalent lattices; when these are sheared they shatter (Figure 2.14).

FIGURE 2.19
The mobile electrons are able to hold the positive ions of metals together even when the metal is distorted. This explains the malleability and ductility of metals



Ionic and covalent lattices are hard but brittle. Metals (generally) are also hard but by contrast they are malleable (able to be rolled into sheets) and ductile (able to be drawn into wires).

2.25 SOLIDS SUMMARISED

The common types of solids and the properties generally used to identify them are summarised in Table 2.6.

Note that the key physical properties for distinguishing between types of solids are melting points, electrical conductivity, hardness, malleability and ductility. Malleability and ductility are sometimes jointly referred to as *workability*.

TABLE 2.6 Types of solids and their properties

MOLECULAR SOLIDS		LATTICE SOLIDS		
		metallic	ionic	covalent network ^a
<i>melting and boiling points</i>	low	variable	high	high
<i>conduct electricity?</i>	no	yes	as solid, no molten, yes	no
<i>hardness and/or workability</i>	soft	variable hardness; malleable and ductile	hard and brittle	hard and brittle
<i>forces holding particles together in the solid</i>	intermolecular	delocalised electrons (metallic bonding)	electrostatic	covalent bonding throughout the crystal

^a sometimes called covalent lattice solids

EXERCISES

32 Silicon dioxide and carbon dioxide have similar formulae, SiO_2 and CO_2 , yet one is an extremely hard solid while the other is a gas at room temperature. Explain why they are so different.



***33** The melting points and electrical conductivities of seven elements are as follows:

Element	A	B	C	D	E	F	G
Melting point ($^{\circ}\text{C}$)	113	650	3550	1280	2030	1083	44
Conductivity ^a	10^{-20}	23	10^{-17}	29	10^{-10}	58	10^{-15}

^a of a 1 metre cube in megohm $^{-1}$.

- a** Which elements are metals?
- b** Which elements are made up of discrete small covalent molecules?
- c** Which elements are covalent lattices?
- d** Explain why each of these three classes of elements has the properties given in the table.

Important new terms

You should know the meaning of the following terms:

anions (p. 51)
atom (p. 35)
atomic number (p. 42)
cations (p. 51)
chemical bond (p. 52)
covalent bonding (p. 52)

covalent lattices (p. 59)
covalent lattice solids (p. 59)
covalent molecular substances (p. 54)
covalent molecule (p. 53)
covalent network solids (p. 59)
delocalised electrons (p. 62)
diatomic molecule (p. 39)
discrete energy levels (p. 43)
electron (p. 40)

CHAPTER 2

electron cloud (p. 40)	metallic bonding (p. 62)
electron configuration (p. 44)	molecule (p. 35)
electron-dot diagram (structure) (p. 53)	monatomic molecule (p. 39)
empirical formula (p. 51)	neutron (p. 41)
energy levels (shells) (p. 43, 45)	nucleus (p. 40)
formulae (p. 37)	proton (p. 41)
intermolecular forces (p. 58)	stable electron configurations (p. 45)
ionic bonding (p. 51)	symbols (for elements) (p. 37)
ionic lattice (p. 59)	transition elements (metals) (p. 48)
ions (p. 50)	transition series (p. 48)
lattice (p. 59)	valence electrons (p. 47)
mass number (p. 42)	valence shell (p. 47)

TEST YOURSELF

- 1 Explain the meaning of each of the items in the ‘Important new terms’ section above.
- 2 List ten elements for which the symbol is just the capital of the first letter of the name.
- 3 Give three elements and their symbols for which the symbol is two letters starting with:
a B **b** C **c** N
- 4 Give the symbols for six elements for which the symbol is based on a non-English name of the element.
- 5 What does the formula of a compound tell you about the compound?
- 6 Why do we say that oxygen gas is O₂ and not just O?
- 7 List four gaseous elements which exist as diatomic molecules and give their formulae.
- 8 Describe the structure of an atom. What are the main particles that make up an atom? Where is virtually all the mass located? What occupies most of the volume of the atom?
- 9 Why do we use the term ‘electron cloud’ in describing an atom?
- 10 Approximately what fraction of the radius of a *whole* atom is the radius of the *nucleus*?
- 11 Describe the relative mass and charge of protons, neutrons and electrons.
- 12 What particles are present in the nucleus of an atom?
- 13 How do atoms of one element differ from atoms of other elements?
- 14 In an atom what, if any, relation exists between:
 - a number of electrons and number of protons
 - b number of protons and number of neutrons?
- 15 Carbon has an atomic number of 6 and a mass number of 12. How many protons, neutrons and electrons are there in the atom?

- 16** A fluorine atom consists of nine protons, ten neutrons and nine electrons. What is its mass number and atomic number?
- 17** Do all electrons within the one atom have the same amount of energy? Explain.
- 18** What is the maximum number of electrons which can exist in each of the first, second, third and fourth energy levels?
- 19** The first twenty elements of the Periodic Table are H, He, Li, Be, B, C, N, O, F, Ne, Na, Mg, Al, Si, P, S, Cl, Ar, K, Ca. Give their electron configurations.
- 20** What common feature is there about the electron configurations of:
- a** the noble gases (He, Ne, Ar, Kr, Xe, Rn)
 - b** the alkali metals (Li, Na, K, Rb, Cs)
 - c** the halogens (F, Cl, Br, I)?
- 21** What is the ‘driving force’ behind chemical reactivity? (That is, when atoms react what stable states are they trying to achieve?)
- 22** Why do: **a** the alkali metals (Li, Na, K, Rb, Cs) **b** the halogens (F, Cl, Br, I) have similar properties?
- 23** Why do the metals Mg, Ca, Sr and Ba all undergo very similar reactions?
- 24** What principles are used in drawing up the Periodic Table of the elements?
- 25** In the Periodic Table, what name is given to:
- a** vertical columns of elements
 - b** horizontal rows of elements?
- 26** Sketch the Periodic Table in block form and show where the following occur:
- a** the noble gases **c** the alkali metals
 - b** the halogens **d** the alkaline earth metals (Be, Mg, Ca, Sr, Ba)
- 27** What name is used to describe the elements that correspond to building up an energy level from 8 to 18 electrons?
- 28** On your sketch for Question 26 show the location of the transition metals.
- 29** Why are we particularly interested in the valence electrons of atoms?
- 30** How can you determine the number of valence electrons in an atom without working out its full electron configuration?
- 31** When atoms combine to form compounds, what electron configurations do they ‘strive’ to achieve? In what two general ways do they do this?
- 32** Give two examples to illustrate formation of ions.
- 33** What holds the ions together in ionic compounds? Why do we say that there are no molecules in ionic compounds?
- 34** Give two examples of covalent bonding.
- 35** Draw electron-dot diagrams for hydrogen bromide and water.
- 36** What type of compound, ionic or covalent, does carbon always form?
- 37** What type of compound, ionic or covalent, is usually formed by:
- a** alkali metals (Group 1) **c** halogens
 - b** alkaline earth metals (Group 2) **d** oxygen and sulfur?

- 38** Give three common properties of ionic compounds.
- 39** How do melting points and boiling points of covalent compounds compare with those of ionic compounds? Explain why.
- 40** Do covalent compounds conduct electricity as pure substances? Explain why.
- 41** Based on structure we can divide pure solids into four classes. What are they?
- 42** Draw a structure of one covalent network solid; give its name and empirical formula.
- 43** Give three properties of covalent lattices that are consequences of the chemical bonding involved.
- 44** Draw a diagram to illustrate metallic bonding.
- 45** What properties of metals are a direct consequence of this type of bonding?
- 46** Complete the following table using the terms given below the table. If any property is inapplicable, put a dash.

Substance	Physical state ^a (at 25°C)	Melting point ^b	Electrical conductance ^c			Example
			Pure	In aq. soln	Hardness ^d	
ionic lattice						
covalent molecule						
metallic lattice						
covalent lattice						

^a solid, liquid or gas ^b low, high, very high ^c zero, low, high ^d of the solid if applicable: hard or soft.

- 47** Name four elements that exist as (exactly or approximately) covalent lattices.

Chemical reactions, names and formulae

IN THIS CHAPTER

Physical and chemical changes

Decomposition and direct combination reactions

Energy changes for these reactions

Chemical equations

Formulae and names for ionic compounds

Formulae and names for binary covalent compounds

In Chapter 1 we looked at the range of substances available to us from the Earth and saw how to separate out desired ones. Often, however, what we really want are substances that we extract or make from the separated Earth materials—aluminium from aluminium oxide, iron from haematite, calcium oxide from limestone, chlorine from sodium chloride. We use *chemical reactions* to obtain the desired substances. In this chapter, then, we shall explore what is meant by chemical reaction and look at some specific examples.

All of the diverse compounds we obtain from the Earth have names and chemical formulae. Ideally the name of a compound should allow us to write down its formula and vice versa. We shall therefore look at the matter of writing formulae and names for compounds.

Let us start with the distinction between physical and chemical change.

3.1 PHYSICAL AND CHEMICAL CHANGES

Many of the changes that we can make to materials involve:

- changing the state of a substance (melting lead, boiling water)
- changing the physical appearance (grinding big lumps of limestone into fine powder, rolling blocks of copper into sheets or drawing into wires)
- dissolving a solid in a liquid (iodine in hexane, sugar in water)
- separating mixtures (separating sea water into solid salt and liquid water, filtering sand from water).

In these changes no new substance has been formed.

A change in which no new substance is formed is called a **physical change**.



The sublimation and condensation of iodine
(solid → gas → solid)
are physical changes



FIGURE 3.1
Purple potassium permanganate solution undergoes a chemical change when mixed with colourless hydrogen peroxide solution

Colourless solutions of potassium iodide and lead nitrate react chemically to produce a yellow precipitate

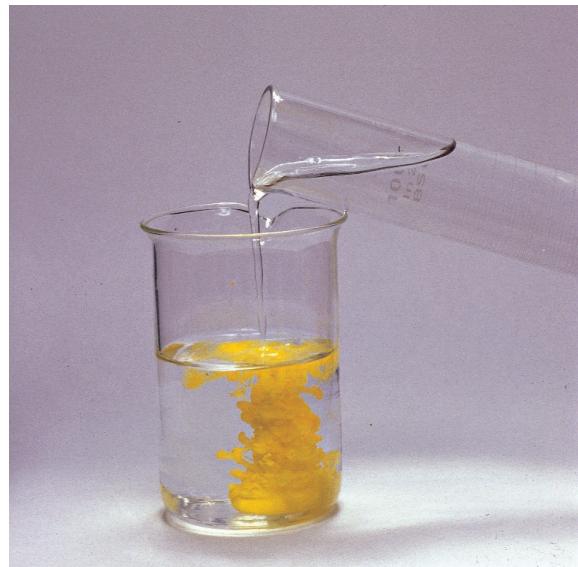
By contrast there are changes in which new substances are formed. Examples are:

- heating green copper carbonate to form a black solid and a colourless gas
- burning silvery magnesium ribbon to form a white powder
- heating a mixture of grey iron filings and yellow sulfur to form a homogeneous black solid.

A change in which at least one new substance is formed is called a **chemical change**.

Chemical changes are also called **chemical reactions**. Common ways of deciding that a chemical change (reaction) has occurred are:

- 1 if a *gas is evolved*, for example zinc granules dropped into hydrochloric acid solution
- 2 if a *solid (called a precipitate) is formed* when two solutions are mixed; for example when silver nitrate and sodium chloride solutions are mixed, a white solid (silver chloride) forms
- 3 if there is a *change in colour*; for example when purple potassium permanganate solution is added to hydrogen peroxide, the mixture becomes colourless (Fig. 3.1)
- 4 if there is a *significant change in temperature* of the mixture; for example when magnesium is burnt, the metal becomes very hot
- 5 if there is *disappearance of a solid* which is not merely physical dissolution of the solid in the solvent; for example white magnesium hydroxide powder added to hydrochloric acid produces a clear solution
- 6 if an *odour is produced*; for example when sodium hydroxide is added to a warm solution of ammonium chloride, the sharp smell of ammonia is detected.



In a chemical reaction the starting substances are called **reactants** and the substances that are formed are called **products**.

Table 3.1 summarises the differences between chemical reactions and physical changes.

TABLE 3.1 Differences between chemical and physical changes

Chemical change (reaction)	Physical change
at least one new substance is formed	no new substance is formed
difficult to reverse (hard to un-boil an egg)	easily reversed (melt a solid; freeze it again)
generally a large input or output of energy (burn natural gas)	relatively small energy changes involved (evaporate alcohol, dissolve sugar in water)

In all chemical reactions:

- 1 *Mass is conserved*; that is

$$\text{total mass of products} = \text{total mass of reactants}$$

- 2 *The number of atoms of each type is conserved*. For example in case 5 above, the number of chlorine atoms in the products is equal to the number in the reactants; similarly for magnesium atoms, oxygen atoms and hydrogen atoms.

These facts are often summarised in what is called the **law of conservation of mass** (or *conservation of matter*):

Matter can be neither created nor destroyed, but merely changed from one form to another.

That mass is conserved can be shown by:

- burning a candle in an enclosed jar of air and weighing the jar and its contents before and after the reaction (no change in mass), or
- in cases 2 and 5 above, weighing the reactants and then the mixture after reaction has occurred (but not for cases 1, 3, 4 and 6. Why?).

Although new substances are formed in chemical reactions, *no new or extra atoms are created*. New substances (elements or compounds) are produced by rearranging the atoms from the old ones (the reactants).

Physical and chemical properties

Properties of substances that relate to physical changes, such as melting and boiling points, are called **physical properties**. Other common physical properties are appearance, density, electrical conductivity, and hardness.

Properties that relate to the chemical reactions that substances undergo are called **chemical properties**. Common chemical properties are ease of decomposition by heat, effect of light, and reactivity with other substances such as oxygen, chlorine and sulfur.



EXERCISES

- 1 Classify each of the following as a *physical* or *chemical* change:

a tarnishing of copper	d melting wax
b dissolving sugar in water	e evaporating sea water to dryness
c burning a candle	f rusting of iron
- 2 When a silvery granule of zinc is dropped into a pale blue solution of copper sulfate, a reddish brown solid deposits on the zinc and the solution loses its blue colour. Is this a physical or chemical change? Give your reasons.
- 3 Decide whether the changes in **a** and **b** are physical or chemical. Explain your reasoning.
 - a** When 1.00 g of a certain white solid was mixed with 20 mL of water, a clear solution resulted. When this solution was evaporated to dryness, 1.41 g of white solid was recovered. The original white solid remained as a solid when heated to 400°C; the recovered white solid melted at 360°C.
 - b** When 1.00 g of a different white solid was mixed with 20 mL of water, a clear solution resulted. When this solution was heated to evaporate off the water, 1.00 g of white solid was recovered. The white solid had the same melting point as the starting solid.

3.2 DECOMPOSITION REACTIONS

In discussing elements and compounds (Section 1.2), we saw that some pure substances (compounds) can be decomposed into two or more other pure substances. This usually requires the input of energy in the form of heat, light or electricity.

Decomposing substances by heating them

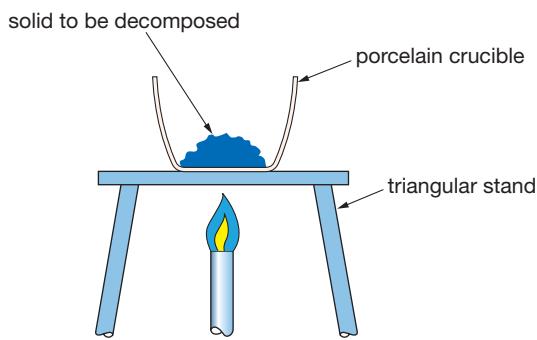
Some pure substances, such as copper nitrate, can be decomposed by heating.

Heating copper nitrate

Copper nitrate is a blue solid. It is soluble in water. It is homogeneous and its properties do not change after repeated purification procedures. Copper nitrate is therefore a pure substance. When copper nitrate is heated in a crucible as in Figure 3.2, a brown gas is evolved and a black solid remains. This black solid is insoluble in water. Hence it is a different substance from copper nitrate. This black solid always has a smaller mass than the sample of copper nitrate originally taken. Copper nitrate has therefore been decomposed into two other substances—a brown gas (nitrogen dioxide) and a black solid (copper oxide)—and a third substance, colourless oxygen gas, which went undetected. This is therefore a chemical change (or a chemical reaction).

FIGURE 3.2

Decomposing a solid by heating it in a crucible



Decomposing substances with electricity

Some substances do not decompose when heated but can be broken into two other substances by passing an electric current through them. Decomposing a substance with an electric current is called **electrolysis**. The conductors that introduce an electric current to a liquid or solution are called **electrodes**. Examples of electrolysis are discussed below.

1 Electrolysis of molten lead bromide

Lead bromide is a white crystalline solid. It is a pure substance which melts at 373°C without decomposing. If an electric current is passed through molten lead bromide at about 400°C as in Figure 3.3, a choking brown vapour is formed at the positive electrode while drops of silvery liquid form at the negative electrode. Lead bromide has been decomposed into molten lead (the silvery liquid) and brown bromine gas, again a chemical reaction.

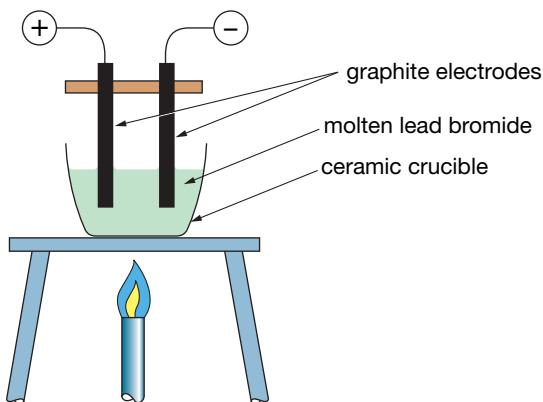


FIGURE 3.3
Electrolysis of molten lead bromide. This experiment is too dangerous to be performed in a school classroom

2 Electrolysis of water

Water can be electrolysed as follows. Two 10 mL measuring cylinders filled with water are inverted over a pair of inert electrodes in a beaker of water containing a little sulfuric acid, as in Figure 3.4. The acid is necessary because pure water is a very poor conductor of electricity. The electrodes are connected to a voltage source and a current allowed to flow.

After several hours we find that colourless gas has collected in each cylinder with the volume above the negative electrode being twice that above the positive electrode, as in Figure 3.4. Tests show that the larger volume of gas is hydrogen while the other gas is oxygen. An analysis would show that the amount of sulfuric acid present at the end of the experiment was the same as at the beginning. This means that it was the water, not the sulfuric acid, that had decomposed.

The experiment shows that the pure substance water can be decomposed into hydrogen and oxygen gases, so a chemical change has occurred.

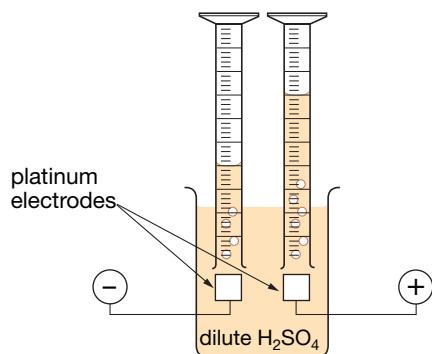


FIGURE 3.4
Electrolysis of water: initially both measuring cylinders were filled with water but after several hours gas has collected in both cylinders

Boiling and electrolysis of water

The two processes, boiling water (to convert liquid to vapour) and electrolysing water (to hydrogen and oxygen gases), clearly illustrate the differences between physical and chemical changes.

In terms of the differences listed in Table 3.1:

- electrolysis produces two new substances (hydrogen and oxygen gases), whereas boiling does not produce any new substance (just converts liquid water to gaseous water)
- electrolysis is difficult to reverse (need to mix the gases together and ignite them with a high temperature spark), whereas boiling is easily reversed (just cool the vapour and it goes back to liquid)
- electrolysis requires much more energy than boiling (electrolysis requires between 20 and 30 kilojoules of electrical energy per gram of water decomposed (depending on conditions used), whereas boiling requires only 2.3 kilojoules per gram of water vaporised).

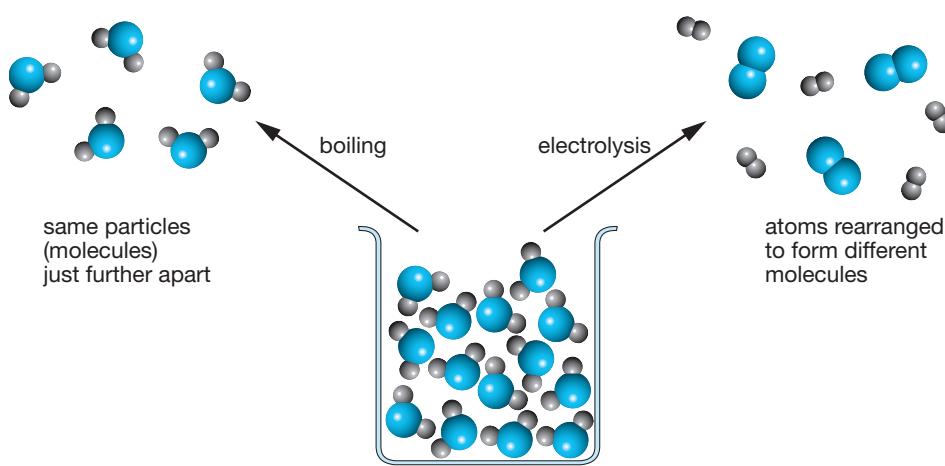
In terms of particles

Boiling water does not alter the actual particles (molecules): it just separates them from one another: the water vapour contains the same water molecules as the liquid did. Electrolysis actually breaks the particles up (water molecules are broken up and hydrogen and oxygen molecules are formed). These changes are shown in Figure 3.5.

This is generally true about changes: physical changes just rearrange the particles without changing their nature: chemical changes break up the particles (molecules) and rearrange the atoms into new substances. But in both types of change there is no alteration to the actual number of each type of atom present.

FIGURE 3.5

Boiling and electrolysis of water on the particle (molecule) level



Decomposing pure substances with light

Some pure substances can be decomposed by sunlight. Silver chloride is a white solid. When exposed to sunlight it turns first purple, then black. After several hours of exposure, the sample has a smaller mass than the starting material. Sunlight has decomposed silver chloride into a black solid (finely divided silver metal) and a gas (chlorine) which we did not see because it formed quite slowly in small amounts that were quickly dispersed by air currents.

Silver bromide and silver iodide behave similarly. In fact this decomposition of silver compounds by light is the basis of photography.

3.3 DIRECT COMBINATION REACTIONS

In addition to reactions in which one compound is decomposed into two or more substances (elements or compounds), there are many reactions in which a compound is formed from its elements. These are called **direct combination reactions**.

Some direct combination or synthesis reactions are:

- 1 *Silvery magnesium burns in air* (or oxygen) to form a white powder, magnesium oxide, MgO. This reaction liberates much light energy (the flame is extremely bright) and also releases a lot of heat.
- 2 *Reddish brown copper combines with yellow sulfur* when heated to form black copper(I) sulfide solid, Cu₂S. Much heat energy is released.
- 3 *Hydrogen and chlorine gases react vigorously* to form hydrogen chloride gas, HCl. Again much energy is released.
- 4 *Heated phosphorus reacts with chlorine* to form liquid phosphorus trichloride, PCl₃. Both heat and light energy are released.

Generally when elements combine directly to form compounds, energy is released. In the previous section we saw that to decompose compounds we generally need to supply energy.

3.4 EXPLANATION FOR ENERGY CHANGES

Decomposing a compound into atoms requires a large input of energy because it is necessary to overcome the strong chemical bonds holding the atoms together in compounds. As we saw in Sections 2.15 to 2.23 there are strong electrostatic attractions holding ions together in ionic compounds and strong covalent bonds holding atoms together in covalent molecules and in covalent lattices.

The stronger the chemical bonding in a compound, the more energy that is required to break the compound into atoms. Alternatively, the stronger the chemical bonding in a compound the more energy that is released when the compound is formed from its atoms.

3.5 EVERYDAY APPLICATIONS

Decomposition reactions are widely used in the everyday world, with examples being:

- in air bags in motor cars sodium azide is decomposed to sodium and nitrogen gas (by igniting it with a detonating cap)
- calcium carbonate (limestone) is decomposed (to calcium oxide and carbon dioxide) by heating it to make lime, cement and glass
- aluminium is extracted commercially by electrolysing molten aluminium oxide

Direct combination reactions are also common in:

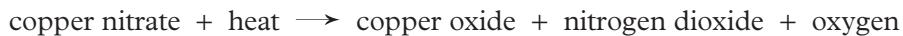
- the rusting of iron and steel (to form iron(III) oxide from the direct combination of iron with oxygen)
- the burning of coke (carbon) which releases much heat energy that we use in many different ways

- lightning which creates such a localised high temperature that nitrogen and oxygen gases combine to form gaseous nitric oxide, NO. This same combination reaction also occurs in the combustion chamber of cars and this nitric oxide along with nitrogen dioxide, NO₂, to which it is partly converted, is one of the major pollutants produced by cars.

3.6 EQUATIONS FOR CHEMICAL REACTIONS

Chemists use equations to describe what happens in chemical reactions. To start with we write **word equations**. Examples are (using information in Sections 3.2 and 3.3):

- Heating copper nitrate:*



Copper nitrate is the reactant; copper oxide, nitrogen dioxide and oxygen are the products.

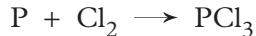
- Phosphorus reacting with chlorine*



Phosphorus and chlorine are the reactants; phosphorus trichloride is the product.

Before we can write a word equation we need to know *all* the reactants and *all* the products of the reaction.

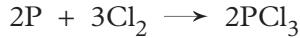
We then convert word equations into **symbolic equations**—equations that use symbols and formulae for the elements and compounds involved. Symbolic equations are often just called **chemical equations**. To do this we need to know, or be able to work out, the formulae of all the substances involved in the reaction. For the phosphorus plus chlorine reaction the first step is:



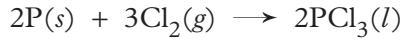
In Section 3.1 we saw that in chemical reactions mass is conserved and the number of each type of atom involved is conserved also. This means that:

a chemical equation must have the same number of atoms of each element on each side of the arrow.

We say that *chemical equations must be balanced*. To do this we put numbers in front of the formulae until the number of atoms of each type on both sides of the equation are equal:



The final thing we need to include in our equation is the physical state for each species present: in this case (*s*) for solid, (*l*) for liquid, and (*g*) for gas:



This equation now reads, *two* atoms of solid phosphorus react with *three* molecules of chlorine gas to form *two* molecules of liquid phosphorus trichloride. This is called a **balanced equation**.

Figure 3.6 shows a symbolic equation and what it means in terms of molecules. We shall return to the question of how to balance equations in Section 4.7.

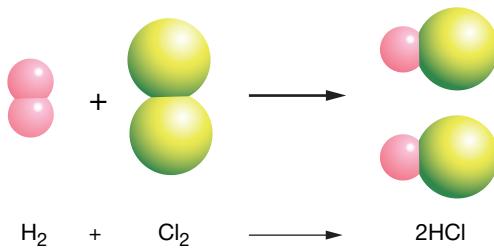


FIGURE 3.6
The equation for the reaction between hydrogen and chlorine to form hydrogen chloride

EXERCISES

- *4 Draw a diagram to illustrate the particle nature of **a** sublimation of solid iodine, I₂ (sublimation is the conversion of a solid directly to a gas) and **b** the chemical reaction of gaseous iodine with hydrogen to form gaseous hydrogen iodide, HI. Take care to differentiate clearly between the different types of particles involved.
- 5 **a** Aluminium, iron and copper each react with oxygen to form oxides. Heat is released in each case. The amounts of heat released are 14, 7 and 3 respectively (in units of 10^{-19} joules per atom of metal). What is the minimum amount of energy required (per atom of metal) to decompose the oxides of these metals? Explain your reasoning.
- b** What do the amounts of energy given in **a** tell you about the relative strengths of the chemical bonding in these three oxides? Explain why.
- 6 A 3.00 g sample of a silvery solid was heated in the apparatus shown in Figure 3.3 until it changed to a liquid. A voltage was applied to the electrodes and a current flowed. There was no apparent change to the liquid. The electrodes were removed, the liquid cooled and then re-weighed: it was still 3.00 g. The solid was then heated in air for several hours. It slowly changed to a white solid. When this solid was melted and electrolysed, silvery droplets formed at the negative electrode and colourless gas at the positive electrode.
- a** Is **i** the silvery solid **ii** the white solid an element or compound or is it not possible to decide? Explain.
- b** Would you expect the mass of white solid formed to be equal to, less than or greater than 3.00 g? After the white solid had been electrolysed, how would you expect its mass to compare with that before the electrolysis? Explain.
- c** Identify any chemical changes that occurred in this experiment.
- 7 Use information in Sections 3.2 and 3.3 to write word equations for:
- a** decomposition of silver chloride by light
- ***b** reaction of magnesium with oxygen
- 8 Write word equations and symbolic equations that convey the same information as the following statements:
- a** Two molecules of ethane, each of which contains two carbon and six hydrogen atoms, react with seven molecules of oxygen, each containing two oxygen atoms, to form four molecules of carbon dioxide, each containing one carbon atom and two oxygen atoms, and six molecules of water (H₂O).
- ***b** One molecule of dinitrogen pentoxide, which contains two nitrogen atoms and five oxygen atoms, reacts with one molecule of water to form two molecules of nitric acid, each containing one hydrogen, one nitrogen and three oxygen atoms.
- 9 Write in words all the information conveyed by each of the following equations (in a similar style to that used in Exercise 8):
- a** C₄H₈(g) + 6O₂(g) → 4CO₂(g) + 4H₂O(l) (C₄H₈ is called butene.)
- ***b** Fe₂O₃(s) + 3CO(g) → 2Fe(s) + 3CO₂(g)



(Fe_2O_3 , CO and CO_2 are called iron(III) oxide, carbon monoxide and carbon dioxide respectively.)

- 10 a** The equation for decomposition of copper nitrate (Section 3.2) is:



- b** The equation for combustion of butane, the fuel in handy-pack fuel cans, is:



For each of these equations, how many atoms of each type are there on each side of the arrow? Are these equations balanced? Explain why or why not.

We need to be able to write formulae from names of compounds, not just in order to write symbolic equations but for many other purposes in chemistry. Let's now turn to that matter.

3.7 FORMULAE AND NAMES FOR COMPOUNDS

As we have already seen, there is a huge number of compounds in existence. It would be impractical to have to memorise formulae for all of these. What we need is a system in which the name of a compound lets us work out its formula, and the formula of a compound allows us to write down its name. This avoids much confusion and minimises memory work.

For writing formulae and names different rules apply for ionic and covalent compounds, so we have to take each in turn.

3.8 FORMULAE FOR IONIC COMPOUNDS

We saw in Section 2.16 that for many elements we can deduce the charge on the ions they form from their position in the Periodic Table. This cannot be done for other elements such as lead and tin and the transition metals, particularly as most of these can form ions of more than one charge type (e.g. Fe^{2+} and Fe^{3+}). The ions formed by the common elements are therefore summarised in Table 3.2, opposite. We need to memorise these charges so that we can write formulae of ionic compounds.

Remember (from Section 2.15) that positive ions are called cations, and negative ions are called anions.

To write the formula for an ionic compound we use the basic principle that such compounds are overall neutral. This means that:

When ionic compounds are formed, the ions are present in ratios such that the total number of positive charges is equal to the total number of negative charges.

Sodium chloride is NaCl , one positive charge with one negative charge. In magnesium chloride we need *two* singly charged Cl^- ions to balance the double charge on magnesium, Mg^{2+} . Hence magnesium chloride is MgCl_2 . In general:

To determine the values of x and y in the formula A_xB_y of an ionic compound, we choose the smallest values of x and y that make

$$x \times (\text{magnitude of charge on A}) = y \times (\text{magnitude of charge on B}) \quad \dots (3.1)$$

TABLE 3.2 Ions formed by common elements

	Charge		Valency
one type of ion only	+1	Group 1 ^a (Li^+ , Na^+ , K^+ , Rb^+ , Cs^+), Ag^+	1
	+2	Group 2 ^a (Be^{2+} , Mg^{2+} , Ca^{2+} , Sr^{2+} , Ba^{2+}), Zn^{2+}	2
	+3	Al^{3+}	3
	-1	Group 7 ^a (F^- , Cl^- , Br^- , I^-)	1
	-2	Group 6 ^a (O^{2-} , S^{2-})	2
more than one type of ion	+1 and -1	H^+ and H^-	
	+1	Cu^+	1
	+2	Cu^{2+} , Fe^{2+} , Pb^{2+} , Sn^{2+}	2
	+3	Fe^{3+}	3
	+4	Pb^{4+} , Sn^{4+}	4

^a of the Periodic Table

Aluminium oxide contains the ions Al^{3+} and O^{2-} . If we write the formula as Al_xO_y then

$$x \times 3 = y \times 2$$

Therefore $x = 2$ and $y = 3$, giving the formula Al_2O_3 .

A term which is commonly used in these contexts is valence.

Valence or valency

The **valence** or **valency** of an element is a number that measures the combining power of the element when it forms compounds.

When an element forms ionic compounds, the valency of the element is the numerical value of the charge that the ion of the element carries.[†]

Na, K, Rb, Ag have valencies of 1

Mg, Ca, Ba, Zn have valencies of 2

F, Cl, Br, I have valencies of 1

O, S have valencies of 2

Valencies of the common elements forming ionic compounds are shown in Table 3.2.

3.9 NAMING SIMPLE IONIC COMPOUNDS

All the compounds in the previous section *consist of two elements only*. They are called **binary compounds**.

Ionic binary compounds are named by giving the positive ion first, then the negative ion; the positive ion has the same name as the element (sodium, calcium), while for the negative ion the ending of the element name is changed to *-ide* as shown in Table 3.3. Hence sodium chloride, magnesium oxide, and potassium sulfide.

[†] Some chemists define it as the actual charge and so attach signs to their valencies: +1, +2, -1, -2 etc. Don't worry about this difference—follow your teacher's advice.

TABLE 3.3 Names of negative ions formed from single atoms

Element		Negative ion formed		Example	
Cl	chlorine	Cl^-	chloride	ZnCl_2	zinc chloride
Br	bromine	Br^-	bromide	AgBr	silver bromide
I	iodine	I^-	iodide	MgI_2	magnesium iodide
O	oxygen	O^{2-}	oxide	CaO	calcium oxide
S	sulfur	S^{2-}	sulfide	K_2S	potassium sulfide
H	hydrogen	H^-	hydride	LiH	lithium hydride

For those metals which display variable valencies (Table 3.4), the valence of the metal in the compound being named is indicated either by:

- 1 putting the valence as a capital Roman numeral in brackets after the name of the metal, as in iron(II) chloride, FeCl_2 ; iron(III) chloride, FeCl_3 ; copper(I) oxide, Cu_2O , copper(II) oxide, CuO , or
- 2 using special endings: *-ous* for the lower valence state, and *-ic* for the higher valence state—sometimes with a Latinised name for the metal. Hence cuprous sulfide, Cu_2S ; cupric sulfide, CuS ; ferrous chloride, FeCl_2 ; ferric chloride, FeCl_3 . Such names are given in Table 3.4.

The first method (using Roman numerals) is preferred, but the ‘-ous’, ‘-ic’ names are still in such widespread use that we need to know them too.

TABLE 3.4 Names of positive ions formed from variable-valency metals

Element		Ion	Preferred name	Other common name
copper	Cu	Cu^+	copper(I)	cuprous
		Cu^{2+}	copper(II)	cupric
iron	Fe	Fe^{2+}	iron(II)	ferrous
		Fe^{3+}	iron(III)	ferric
lead	Pb	Pb^{2+}	lead(II)	plumbous
		Pb^{4+}	lead(IV)	plumbic
tin	Sn	Sn^{2+}	tin(II)	stannous
		Sn^{4+}	tin(IV)	stannic

Compounds of copper(I) and lead(IV) are fairly rare. Hence chemists frequently drop the (II) when writing compounds of copper and lead with a valency of 2. We frequently see copper sulfate instead of copper(II) sulfate and lead iodide instead of lead(II) iodide. If the valency is not given we always assume it is the more common valency. On the other hand, Cu_2S and PbO_2 must always be written as copper(I) sulfide and lead(IV) oxide.

When both valencies are equally common, as with iron, we must *always* include the valency in the name—iron(II) chloride, iron(III) oxide.

3.10 IONS THAT ARE NOT MONATOMIC

So far all the ions considered have been monatomic. We need to include some other ions which commonly occur in the Earth. These are shown in Table 3.5.

Although made up of two or more atoms, each of these ions acts as a separate entity when it forms compounds. Sodium sulfate Na_2SO_4 is two Na^+ ions and one SO_4^{2-} ion: the S and O atoms do not split up. Ions such as OH^- are called **diatomic ions** while ions such as SO_4^{2-} are called **polyatomic ions** (because they consist of three or more atoms).

TABLE 3.5 Some common polyatomic ions and their charges

Name of the ion	Formula	Valency	Example
ammonium	NH_4^+	1	ammonium sulfide, $(\text{NH}_4)_2\text{S}$
hydroxide	OH^-	1	calcium hydroxide, $\text{Ca}(\text{OH})_2$
nitrate	NO_3^-	1	potassium nitrate, KNO_3
sulfate	SO_4^{2-}	2	sodium sulfate, Na_2SO_4
carbonate	CO_3^{2-}	2	calcium carbonate, CaCO_3
phosphate	PO_4^{3-}	3	sodium phosphate, Na_3PO_4

Formulae for compounds involving these ions are written in the same way as for monatomic ions (using Equation 3.1). For example magnesium nitrate is $\text{Mg}(\text{NO}_3)_2$ because $1 \times 2 = 2 \times 1$. Calcium phosphate is $\text{Ca}_3(\text{PO}_4)_2$ because $3 \times 2 = 2 \times 3$.

EXERCISES

11 Write the formulae for the following ionic compounds:

- | | | | | | |
|---|-------------------|----|--------------------|----|-------------------|
| a | potassium bromide | e | lithium sulfide | *i | tin(II) bromide |
| b | zinc oxide | *f | barium chloride | *j | copper(I) sulfide |
| c | sodium oxide | g | aluminium fluoride | *k | lead(IV) chloride |
| d | calcium hydride | *h | lead fluoride | *l | iron(III) oxide |

12 a Name the compounds, CaH_2 , Li_2S , AlBr_3 , $^*\text{ZnO}$, $^*\text{CuCl}_2$, $^*\text{Ag}_2\text{S}$.

b Give two names for each of the following compounds and indicate which is preferred: Fe_2O_3 , SnS , PbO , CuCl , $^*\text{FeBr}_2$, $^*\text{SnCl}_4$.

13 Give the formulae of:

- | | | | | | |
|---|-------------------|----|---------------------|----|---------------------|
| a | sodium nitrate | c | potassium phosphate | *e | iron(III) hydroxide |
| b | lithium carbonate | *d | zinc nitrate | *f | ammonium carbonate |

14 Give the names of:

- | | | | | | |
|---|------------------------------|---|----------------------------|----|------------------------------|
| a | ZnSO_4 | c | $\text{Fe}(\text{OH})_2$ | *e | $\text{Sn}(\text{NO}_3)_2$ |
| b | $(\text{NH}_4)_2\text{CO}_3$ | d | $\text{Al}(\text{NO}_3)_3$ | *f | $\text{Fe}_2(\text{SO}_4)_3$ |

15 For each of the ionic compounds in the following lists:

- | | | | | | | | |
|----|--------------------------------------|-----|------------------------|-----|---------------|-------|----------------------|
| a | State the valency of the element, M. | | | | | | |
| i | MCl_2 | iii | M_2O_3 | v | MH_2 | *vii | M_2O |
| ii | M_2S | iv | MS_2 | *vi | MF_4 | *viii | MBr_3 |



b State the valency of the element, X.

- | | | | |
|-----------------------------|-----------------------------|-----------------------------|--------------------------------------------|
| i CaX | iii AlX ₃ | v Ag ₂ X | *vii Fe ₂ X ₃ |
| ii Na ₂ X | iv AlX | *vi ZnX ₂ | *viii SnX ₄ |

16 Give the name and formula of the ionic compound formed between:

- | | |
|---------------------------------|---------------------------------------------|
| a zinc and oxygen | f iron and chlorine (two compounds) |
| b potassium and iodine | *g silver and oxygen |
| c aluminium and fluorine | *h lead and oxygen (two compounds) |
| d magnesium and sulfur | *i barium and hydrogen |
| *e calcium and bromine | *j copper and sulfur (two compounds) |

3.11 FORMULAE FOR COVALENT COMPOUNDS

As with ionic compounds, the **valence** or **valency** of an element in a covalent compound is a measure of its combining power. It is the *number of covalent bonds that the element can form*. As explained in Sections 2.17 and 2.18, this can generally be deduced from position in the Periodic Table. However, several elements can display more than one valency in covalent compounds. Valencies shown by common elements in covalent compounds are therefore summarised in Table 3.6.

We need to memorise the usual valencies of the common elements so we can work out the formulae of the binary compounds they form.

To determine the values of x and y in the formula A_xB_y of a covalent compound, we choose the smallest values for x and y that make

$$x \times \text{valency of A} = y \times \text{valency of B} \quad \dots (3.2)$$

TABLE 3.6 Valencies of common elements in covalent compounds

Fixed valencies		Variable valencies	
1	H, F	1	Cl, Br, I ^c
2	O ^a	2	S
4	C ^b , Si	3	N ^d , P
		4	S
		5	N ^d , P
		6	S

^a except in H₂O₂ and a few other compounds where it is 1

^b except in carbon monoxide where it is 2

^c only in some compounds with O or with H and O do these three elements ever have valencies greater than 1

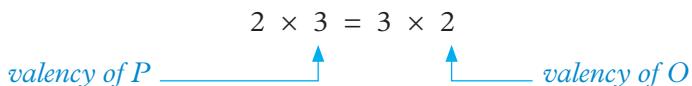
^d in compounds with O, N can display valencies of 1, 2 and 4 as well

When carbon forms a compound with chlorine, the formula is CCl₄ because:

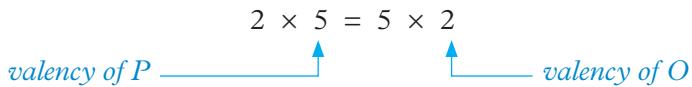
$$1 \times 4 = 4 \times 1$$

valency of C _____ *valency of Cl* _____

When phosphorus reacts with oxygen, two compounds are possible: one in which phosphorus has a valency of 3 and the other in which phosphorus has a valency of 5. For the first compound the formula is P₂O₃, because:



For the other compound the formula is P_2O_5 , because:



3.12 NAMING COVALENT BINARY COMPOUNDS

In covalent compounds many elements show variable valencies (Table 3.6), so an indication of valency is needed in the names of compounds. We therefore name covalent compounds differently from ionic ones.

To name *covalent binary compounds*:

- Use the normal element name for the first element and an *-ide* name for the second, for example hydrogen sulfide and phosphorus trichloride.
- The first element named is the one which occurs farther to the left in the Periodic Table (nitrogen dioxide, boron trifluoride).
- If both elements occur in the same group, the one lower down the group is named first (sulfur dioxide, bromine chloride).
- An exception to the above two rules is that oxygen is named last in compounds with Cl, Br and I (chlorine monoxide, Cl_2O , iodine pentoxide, I_2O_5)
- The number of atoms of each type is given by using the prefixes mono-, di-, tri-, tetra-, penta-, hexa-, ... in front of each part of the name (though mono- is usually omitted from the first-named element and occasionally other prefixes are omitted from it also).

Some examples will illustrate:

CO	carbon monoxide
CO_2	carbon dioxide
NO_2	nitrogen dioxide
N_2O_5	dinitrogen pentoxide
SCl_2	sulfur dichloride
SF_6	sulfur hexafluoride
P_2O_5	phosphorus pentoxide (more correctly diphosphorus pentoxide)
Cl_2O	chlorine monoxide (more correctly dichlorine monoxide)

Several non-metallic hydrides have special names: CH_4 , methane; NH_3 , ammonia; H_2O , water. Others are named in the systematic way: H_2S , hydrogen sulfide (more correctly dihydrogen sulfide); HF, hydrogen fluoride (and similarly HCl, HBr, HI).



EXERCISES

17 Write the formulae for the following compounds:

- | | |
|-------------------------------|-----------------------------------|
| a sulfur difluoride | e sulfur dioxide |
| b nitrogen trichloride | *f silicon tetrachloride |
| c carbon tetraiodide | *g phosphorus pentabromide |
| d boron trichloride | *h dinitrogen trioxide |

18 Name the following compounds:

- a** SiCl_4 , CS_2 , N_2O , Cl_3N , H_2S , * PBr_3 , * SiC
b P_2O_5 , N_2O_4 , Cl_2O , F_2O , * Cl_2O_3 , * SO_3 , * S_2F_2

19 Give the name and formula of the compound formed between:

- | | |
|--------------------------------|--------------------------------------------------|
| a nitrogen and hydrogen | *c hydrogen and iodine |
| b silicon and bromine | *d phosphorus and bromine (two compounds) |

Important new terms

You should know the meaning of the following terms:

balanced equation (p. 74)
binary compound (p. 77)
chemical change (reaction) (p. 68)
chemical equation (p. 74)
chemical properties (p. 69)
diatomic ions (p. 79)
direct combination reaction (p. 73)
electrodes (p. 71)

electrolysis (p. 71)
law of conservation of mass (p. 69)
polyatomic ions (p. 79)
physical change (p. 67)
physical properties (p. 69)
products (p. 69)
reactants (p. 69)
symbolic equation (p. 74)
valence (valency) (p. 77, 80)
word equation (p. 74)

CHAPTER 3

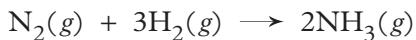
Test yourself

- 1 Explain the meaning of each of the items in the ‘Important new terms’ section above.
- 2 List five observations which can be used to decide whether a change is a chemical one.
- 3 Describe four physical changes and four chemical changes.
- 4 List three ways of decomposing compounds, giving an example for each.
- 5 Describe how you would electrolyse water in order to show the relative volumes of gases formed.
- 6 When a compound is decomposed into elements, what is generally true about the energy change involved?
- 7 Describe three direct combination (of elements) reactions. What is generally true about the energy change in such reactions?
- 8 When a compound is decomposed into atoms, what does the magnitude of the energy change tell you about the chemical bonding in the compound?

9 What do we mean by:

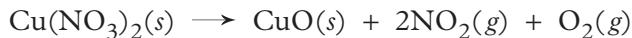
- a** conservation of mass **b** conservation of atoms
- in a chemical reaction?

10 What qualitative and quantitative information does the following chemical equation give you?



11 State in words the complete information conveyed by the chemical equation in Figure 3.6.

12 What is wrong with the following equation?



13 What is the basic rule for working out the formula of an ionic compound?

14 What is the formula of (including charge carried by) the ammonium, hydroxide, sulfate, carbonate, nitrate and phosphate ions?

15 Which elements generally have fixed valencies in covalent compounds and what are they? Name three important elements that show variable valencies and state what their common valencies are.

16 What is the basic rule for working out the formula of a binary covalent compound?

17 In naming covalent binary compounds, how do we decide which element is named first?

18 In naming covalent compounds, what prefixes do we use to denote 1, 2, 3, 4, 5, 6 atoms of a particular type?

EXTENDED RESPONSE EXAM-STYLE QUESTIONS FOR MODULE 1

Questions in this section are in a similar style to that currently being used in the extended response questions in the New South Wales HSC Chemistry examinations: they use the HSC list of verbs (such as analyse, assess, compare, contrast, describe, discuss, evaluate, explain, justify, outline, propose) and relate closely to items in the syllabus document. They place heavy emphasis upon recall of learnt information. Practice is the best way of developing skill in answering this type of exam question.

The actual HSC exam will be a mixture of this type of question and ones similar to the problem-solving exercises located within the chapters of Conquering Chemistry. The Revision Tests that follow these sets of questions at the end of the modules are similar mixes.

Marks (shown at the right-hand end of the question) are assigned to each question in order to indicate how much detail is required in your answer (that is how much time you should spend on each question). The marks are on the HSC exam basis of 100 marks for three hours work (or 1.8 minutes per mark).

MARKS

- 1 **Identify** the characteristics that are commonly used to distinguish between a mixture and a compound. **Evaluate** the usefulness of each for this purpose. **6**
- 2
 - a You are given a mixture of sand and seawater. **Describe** with diagrams how you would obtain samples of clean dry sand, dry salt and pure water from this mixture.
 - b **Explain** how you would determine the percentage composition of this mixture, that is, the percentage of each of the components, sand, salt and water. **6**
- 3 **Outline** two techniques for separating mixtures of solids and two for separating mixtures of liquids. What properties must the components of the mixtures have for these techniques to work? **6**
- 4 **Explain** the relationship between reactivity of an element and the likelihood of it existing on Earth as an uncombined element. Illustrate with examples. **3**
- 5 **Describe** an industrial process that involves the separation of a mixture obtained from the atmosphere, hydrosphere or lithosphere. What properties are used to bring about the separation? In what ways (if any) does this process affect the environment? **7**
- 6 To determine the percentage of dissolved salt in seawater a pair of students weighed a clean dry beaker (285.3 g) then poured into it an accurately measured volume of seawater (250 mL, assumed to have a mass of 250 g) then heated it gently on a hotplate until all the water had evaporated and there remained only a dry solid that was almost white but with a slight tinge of grey; after allowing the beaker and contents to cool, they determined the mass of beaker plus solid to be 294.1 g.
 - a **Calculate** the percentage salt in this seawater.
 - b **Identify** likely sources of error in this experiment and suggest ways of eliminating or minimising them. **5**

- 7** Compare and contrast the physical properties of metals and non-metals. **4**
- 8** You performed an experiment to examine some physical properties of common elements in order to classify them as metals, semi-metals or non-metals. Name the properties that you used and describe how you determined (measured) them. Prepare a table showing the values for these properties for one element you examined from each of the classes above. ‘Values’ can mean ‘high’ or ‘low’, ‘some’ or ‘none’ or other qualitative terms. **7**
- 9** Assess the importance of physical properties in deciding which metal to use for a particular purpose. Illustrate with at least three examples. **5**
- 10** What is meant by ‘valence electrons’? Evaluate the usefulness of this concept (as an alternative to using total electron configurations). **4**
- 11** Describe the chemical bonding in sodium chloride. Explain how it leads to the compound having a high melting point, being hard and brittle, and being a non-conductor of electricity as a solid but a conductor as a liquid. **5**
- 12** Discuss the role of the Periodic Table in predicting the ions formed by atoms of metals and non-metals. **4**
- 13** Describe with an example the nature of the bonding in small covalent molecules. Explain how this type of chemical bonding leads to the physical properties that are common to that class of substance. **4**
- 14** Draw electron-dot structures for the molecules formed between hydrogen and bromine and between hydrogen and sulfur. Assess the usefulness of such structures in predicting the formulae of such compounds. **4**
- 15** Compare and contrast the properties of ionic, covalent molecular and covalent network solids. **5**
- 16** Assess the usefulness of melting point, hardness and electrical conductivity either alone or in combination in classifying substances as ionic, covalent molecular or covalent network solids. **6**
- 17** Describe an experiment you performed to illustrate that the properties of a compound are very different from those of its constituent elements. Present your results in a suitable table. What extra observations or measurements do you wish you could have made but were unable to do because of lack of equipment or facilities? **6**
- 18** Explain why boiling water is a physical change whereas electrolysing water is a chemical one. Draw a diagram on the particle (molecular) level to illustrate your answer. **4**
- 19** Describe a decomposition reaction that you have performed in the laboratory. Mention any tests you performed to identify a product of the reaction. Write a chemical equation for the reaction involved. What safety precautions did you take? **5**
- 20** When atoms combine to form compounds, energy is released. Explain why this is so. However when nitrogen and oxygen gases combine to form nitric oxide, NO, energy is absorbed. Propose an explanation for this apparent anomaly. **4**

REVISION TEST FOR MODULE 1

Total marks: 50 Suggested time: 90 minutes

MULTIPLE CHOICE QUESTIONS

Select the alternative **a**, **b**, **c**, or **d** which best answer the question.

- 1 The list which contains only metallic elements is:

a iron, copper, graphite, manganese	c sodium, barium, silver, mercury
b aluminium, gold, brass, zinc	d magnesium, gold, boron, lead
- 2 The correct name for P_2O_5 is:

a phosphorus(V) oxide	c phosphoric oxide
b diphosphorus pentoxide	d phosphorous oxide
- 3 The bromide and sulfide of element M have the formulae MBr_2 and MS respectively. The formulae for the oxide and fluoride of M are:

a MO and MF_2	c MO and MF
b M_2O and MF	d M_2O and MF_2
- 4 The region of the Earth in which we are most likely to find uncombined elements is:

a atmosphere	c lithosphere
b hydrosphere	d biosphere
- 5 1.00 g of a yellow solid, Z, was heated slowly in an inert atmosphere of nitrogen. It remained solid until a temperature of 165°C was reached, then it melted sharply to a clear liquid. The liquid was heated further. When a temperature of 230°C was reached, the liquid started to give off brown fumes. When fuming ceased at about 260°C , a white solid remained. This white solid had a mass of 0.73 g. From this experiment we can conclude that:

a Z was an element that decomposed when heated
b Z was a mixture with one component vaporising when the temperature reached 230°C
c Z was a mixture in which the components combined chemically at high temperatures
d Z was a pure compound that decomposed at high temperature
- 6 The list that contains only elements is:

a bromine, cobalt, phosphorus, silica
b boron, graphite, methane, sodium
c beryllium, copper, krypton, magnesium
d alumina, fluorine, nitrogen, sulfur
- 7 Two liquids, L and M, and a solid, P, were mixed. The liquids are immiscible. L has a boiling point of 96°C and M of 108°C . P is soluble in M but not in L. The simplest way of obtaining pure samples of these three substances from the mixture would be to:

a distil the mixture: collect pure L first, then M, leaving P behind
b use a separating funnel to separate L from the solution of P in M, then distil M from this solution, leaving P behind

- c filter off P, then use a separating funnel to separate L and M
d use a separating funnel to separate off L, then evaporate the solution of P in M to dryness

8 The potassium ion has the same electron configuration as:

- a the sodium ion c the magnesium atom
b the neon atom d the sulfide ion

9 The list which contains only ionic substances is:

- a sodium fluoride, calcium oxide, hydrogen chloride, potassium sulfide
b potassium oxide, magnesium chloride, strontium sulfide, sodium hydride
c copper oxide, phosphorus trichloride, sodium iodide, calcium chloride
d zinc bromide, magnesium hydride, hydrogen sulfide, lithium oxide

10 The only one of the following statements that is true is:

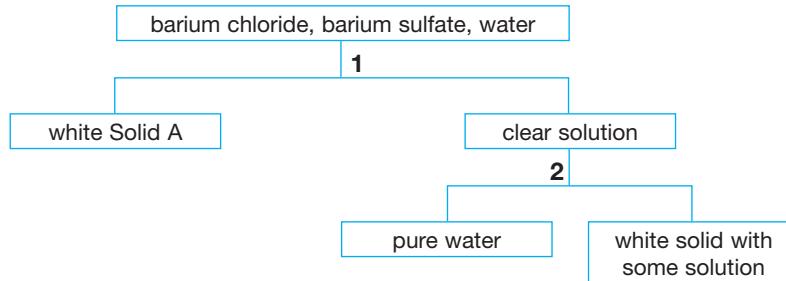
- a The elements on the right-hand side of the Periodic Table form mainly ionic compounds while those on the left form mainly covalent compounds.
b The elements on the left-hand side of the Periodic Table form mainly ionic compounds while those on the right form mainly covalent compounds.
c The elements that commonly form ionic compounds are all located on the right-hand side of the Periodic Table.
d The elements that commonly form covalent compounds are located in the centre and on the right side of the Periodic Table.

SHORT ANSWER AND EXTENDED RESPONSE QUESTIONS

The number of marks for each part of each question is shown at the right-hand end of the question. Marks allocated are based on the HSC scale of 1 mark per 1.8 minute.

MARKS

11 Barium chloride and barium sulfate are white solids. Barium chloride is soluble in water; barium sulfate is not. A mixture of these two compounds was shaken with a quantity of water. In order to separate this mixture into its components the procedure in the flow chart below was followed.



- a Name the processes 1 and 2 in the chart. 1
b Draw simple diagrams showing how these processes would be carried out in the laboratory. 1
c What is the identity of Solid A? 1
- 12 a Which of the following species contain the same number of electrons? 1
O²⁻, F⁻, Ne, Al³⁺
b Which of the species in (a) contain the same number of neutrons? 1

- 13** Discuss the importance of electrolysis in determining that water is a compound and not an element. **3**
- 14 a** Give two examples of decomposition reactions, one brought about by heat and the other by electricity. Write word equations for these decompositions. **2**
- b** Give an example of a direct combination (of elements) reaction and write a word equation for it. **2**
- c** Explain why we need to provide energy for the reactions in (a) and why energy is released by the reaction in (b). **1**
- 15 a** Two miscible liquids A and B have boiling points of 87°C and 169°C respectively. How would you separate a mixture of these liquids? Draw the apparatus you would use. **2**
- b** How would you test whether the sample of A that you obtained from your separation was pure or not? **1**
- c** Two other miscible liquids, C and D, have boiling points of 113°C and 122°C respectively. What would you do differently to separate a mixture of these liquids? **1**
- 16 a** What are the electron configurations of the hydrogen, sodium and sulfur atoms (atomic numbers 1, 11, 16 respectively)? **1**
- b** Draw electron-dot structures for the ions that would be present in an ionic compound formed between two of these elements. Give the formula for the compound. **2**
- c** Draw the electron-dot structure and give the formula for a covalent compound that can be formed between two of these elements. **2**
- 17** Plaster of Paris (used for setting broken limbs and for making plaster board for internal walls in buildings) is made by heating the mineral gypsum; it can be stored as a dry powder indefinitely. To make a cast, plaster of Paris is mixed with water to form a creamy paste, moulded to shape, then left to set. Setting takes from 20 minutes to several hours depending upon the size of the cast. To explore what was happening some students performed the following experiment.
- 5.64 g gypsum was heated for several hours in an evaporating dish over a Bunsen burner. After cooling the powder had a mass of 4.76 g. At the same time 6.82 g plaster of Paris was mixed into a paste with some water in another evaporating dish and allowed to stand overnight to set. Then it was warmed in an oven at 100°C for an hour (to evaporate any excess water present), cooled and weighed. The mass of the solid was 8.09 g.
- a** Calculate the percentage loss in mass of the original gypsum. **1**
- b** Does this percentage loss bear any relation to the gain in mass by the plaster of Paris? What do you conclude about the setting process for plaster of Paris? **2**
- 18** Five solids have the properties listed opposite. The relevant properties of sodium chloride and copper are also given.

Relative electrical conductivity

	Melting point (°C)	Of solids	Of liquid	'Hammer' test^a
A	327	5	2	flattens
B	2030	0	0	shatters
C	91	0	0	forms powder
D	734	0	0.2	forms powder
E	2870	0	0	shatters
NaCl	801	0	0.2	forms powder
Cu	1083	60	4	flattens

^a hit a sample of the substance repeatedly with a hammer

- a** Classify each of the solids, A, B, C, D and E as ionic, covalent molecular, covalent lattice or metallic. **2**
 - b** Explain why sodium chloride and copper have the conductivity properties listed in the table. **2**
 - c** Explain why covalent molecular compounds and covalent lattice compounds have the melting points and conductivities shown in the table. **2**
- 19 a** Describe the separation of a mixture (of at least three substances) that you have performed in the laboratory. Draw a labelled flow chart for the procedure. **3**
- b** Describe the difference in the property or properties that this separation is based upon. **1**
- 20** Evaluate the usefulness of appearance, melting point and electrical conductivity in deciding whether an element is a metal or non-metal. Include examples. **6**

MODULE 1 AND THE NEW SOUTH WALES HSC SYLLABUS

This section allows you to check that *Conquering Chemistry* has covered all necessary material for Module 1 of the New South Wales HSC Preliminary Course.

Syllabus content

The following table lists (for Module 1) the items from the *students learn to* column of the HSC syllabus and shows where they are treated in *Conquering Chemistry Preliminary Course (CCPC)*

Location of HSC syllabus Preliminary Course material in *Conquering Chemistry* for Module 1

Syllabus reference	Sections where found in CCPC
Students learn to	
8.2.1 <i>The living and non-living components of the Earth contain mixtures</i>	
■ construct word and balanced formulae equations of chemical reactions as they are encountered	
■ identify the difference between elements, compounds and mixtures in terms of particle theory	1.1, 1.2, 2.1, 2.2
■ identify that the biosphere, lithosphere, hydrosphere and atmosphere contain examples of mixtures of elements and compounds	1.3, 1.4
■ identify and describe procedures that can be used to separate naturally occurring mixtures of: – solids of different sizes – solids and liquids – dissolved solids in liquids – liquids – gases	1.5 to 1.14 (properties for identifying substances 1.15 to 1.18)
■ assess separation techniques for their suitability in separating examples of earth materials, identifying the differences in properties which enable these separations	1.13
■ describe situations in which gravimetric analysis supplies useful data for chemists and other scientists	1.20
■ apply systematic naming of inorganic compounds as they are introduced in the laboratory	3.7 to 3.12
■ identify IUPAC names for carbon compounds as they are encountered	Module 3
8.2.2 <i>Although most elements are found in combinations on Earth, some elements are found uncombined</i>	
■ explain the relationship between the reactivity of an element and the likelihood of its existing as an uncombined element	1.21, 1.22

Syllabus reference	Sections where found in CCPC
Students learn to	
■ classify elements as metals, non-metals and semi-metals according to their physical properties	1.23
■ account for the uses of metals and non-metals in terms of their physical properties	1.24
8.2.3 Elements in Earth materials are present mostly as compounds because of interactions at the atomic level	
■ identify that matter is made of particles that are continuously moving and interacting	2.1
(Revision of atoms, molecules, symbols, formulae and structure of the atom Sections 2.2 to 2.8)	
■ describe qualitatively the energy levels of electrons in atoms	2.10 to 2.13
■ describe atoms in terms of mass number and atomic number	2.9
■ describe the formation of ions in terms of atoms gaining or losing electrons	2.14, 2.15
■ apply the Periodic Table to predict the ions formed by atoms of metals and non-metals	2.16
■ apply Lewis electron-dot structures to – the formation of ions – the electron sharing in some simple molecules	2.19, 2.20, 2.21 2.17, 2.20
■ describe the formation of ionic compounds in terms of the attraction of ions of opposite charge	2.15
■ describe molecules as particles which can move independently of each other	2.2
■ distinguish between molecules containing one atom (the noble gases) and molecules with more than one atom	2.6
■ describe the formation of covalent molecules in terms of sharing of electrons	2.17, 2.18
■ construct formulae for compounds formed from – ions – atoms sharing electrons	2.15 to 2.19
8.2.4 Energy is required to extract elements from their naturally occurring sources	
■ identify the differences between physical and chemical change in terms of rearrangement of particles	3.1, 3.2
■ summarise the differences between the boiling and electrolysis of water as an example of the difference between physical and chemical change	3.2
■ identify light, heat and electricity as the common forms of energy that may be released or absorbed during the decomposition or synthesis of substances and identify examples of these changes occurring in everyday life	3.2, 3.3, 3.5

Syllabus reference	Sections where found in CCPC
Students learn to	
■ explain that the amount of energy needed to separate atoms in a compound is an indication of the strength of the attraction, or bond, between them	3.4
8.2.5 <i>The properties of elements and compounds are determined by their bonding and structure</i>	
■ identify differences between physical and chemical properties of elements, compounds and mixtures	3.1 (or 1.1, 1.2)
■ describe the physical properties used to classify compounds as ionic or covalent molecular or covalent network	2.22, 2.23, 2.25
■ distinguish between metallic, ionic and covalent bonds	2.15, 2.17, 2.24
■ describe metals as three-dimensional lattices of ions in a sea of electrons	2.24
■ describe ionic compounds in terms of repeating three-dimensional lattices of ions	2.15, 2.22
■ explain why the formula for an ionic compound is an empirical formula	2.15
■ identify common elements that exist as molecules or as covalent lattices	2.23
■ explain the relationship between the properties of conductivity and hardness and the structure of ionic, covalent molecular and covalent network structures	2.22, 2.23, 2.25

Compulsory experiments

The table below lists the compulsory experiments (from the right-hand column of the syllabus), along with places where you can find procedures for these experiments (in *CCPC BLM*) and the location of relevant information in *CCPC*.

CCPC is an abbreviation for this book *Conquering Chemistry Preliminary Course*.

CCPC BLM is *Conquering Chemistry Preliminary Course Blackline Masters* by Debra Smith, McGraw-Hill, Sydney, 2003.

The experiments in italics are suggested in the syllabus but are not strictly compulsory.

Information about compulsory experiments for Module 1

Experiment	Location in CCPC BLM	Related material in CCPC
1 Separate the components of a naturally occurring mixture such as sand, salt and water.	Module 1 Worksheet 3 (page 6)	Chapter 1 Exercises 8, 9, 13, 14, 15 Exam-style question 2 (a) Revision test question 19
2 Carry out a gravimetric analysis of a mixture to estimate its percentage composition.	Module 1 Worksheet 3 (page 6)	Chapter 1 Example on page 24 and Exercises 28, 29, 30 Exam-style questions 2 (b), 6

3	Examine some physical properties of a range of common elements and use them to classify these elements as metals, semi-metals or non-metals	Module 1 Worksheet 6 (page 12)	Exam-style question 8
4	Thermal decomposition of a carbonate	Module 1 Worksheet 11 (page 25)	Exam-style question 19
<i>E1 The effect of light on silver salts</i>		Module 1 Worksheet 12 (page 28)	Section 3.2
<i>E2 Electrolysis of water and its use in identifying water as a compound</i>			Section 3.2
5	Comparison of the properties of a compound with those of its constituent elements	Module 1 Worksheet 15 (page 37)	Exam-style question 17
6	Examine the physical properties of a range of common Earth resources in order to classify them as metallic, ionic, covalent molecular or covalent network substances	Module 1 Worksheet 18 (page 49)	Chapter 2 Exercises 41, 42 Exam-style question 16 Revision test question 18

Other items in the student activity (right-hand) column

The table below lists (in abbreviated form) the non-experimental items in the student activity (right-hand) column of the syllabus and indicates where in *Conquering Chemistry* relevant information is located.

Location of material for other Module 1 activities

Syllabus item	Relevant material in CCPC
8.2.1	<ul style="list-style-type: none"> ■ gather and present information ... to write equations ... all chemical reactions encountered ...
	An ongoing task beginning in Section 3.6, continuing in 4.5 and in later chapters
	<ul style="list-style-type: none"> ■ Experiment 1 (see above)
	<ul style="list-style-type: none"> ■ Experiment 2 (see above)
	<ul style="list-style-type: none"> ■ identify data ... analyse information ... to identify the industrial separation processes used ...
	For Investigation, page 19 Exam-style question 7
8.2.2	<ul style="list-style-type: none"> ■ Experiment 3 (see above)
	<ul style="list-style-type: none"> ■ analyse information ... to distinguish the physical properties of metals and non-metals
	Section 1.22 Chapter 1 Exercise 31 Exam-style question 11 Revision test question 20
	<ul style="list-style-type: none"> ■ process information ... use a Periodic Table to present ... classification of elements ...
	Section 1.24 Chapter 1 Exercises 33 to 35
8.2.3	<ul style="list-style-type: none"> ■ analyse information by constructing or using models ... of metals, ionic compounds and covalent compounds
	Depends on the models available in your school; Figs 2.17, 2.18, 2.10, 2.13, 2.12, 2.14 may help
	<ul style="list-style-type: none"> ■ construct ionic equations showing metal and non-metal atoms forming ions
	Section 2.19, 2.21 Chapter 2 Exercise 28
8.2.4	<ul style="list-style-type: none"> ■ Experiment 4 (see above)
	<ul style="list-style-type: none"> ■ Experiments E1 and E2 (see above)

Syllabus item	Relevant material in CCPC
■ analyse and present information to model ...boiling ... and electrolysis of water ...	Section 3.2 and Fig 3.5 Revision test question 13
8.2.5	
■ Experiment 5 (see above)	
■ choose resources ... to construct and discuss ... models of ionic lattices, covalent molecules and covalent and metallic lattices	Depends on the models available in your school; Figs 2.17, 2.18, 2.10, 2.13, 2.12, 2.14 may help
■ Experiment 6 (see above)	

Prescribed focus areas

The five prescribed focus areas of the syllabus are

- 1 the history of chemistry
- 2 the nature and practice of chemistry
- 3 applications and uses of chemistry
- 4 implications for society and the environment
- 5 current issues, research and development.

The material of Module 1 (in this book and in the syllabus) relates almost exclusively to areas 2 and 3; there is little scope in this module for discussing history, implications for society ... or current issues The best examples of the applications and uses of chemistry (area 3) in this module are separations of mixtures (such as in Sections 1.6 to 1.14 and Exercises 15 to 20 in Chapter 1) and everyday applications of decomposition and combination reactions (Section 3.5).