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Third Edition

Bryan Earl
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International hazard warning symbols

You will need to be familiar with these symbols when undertaking practical experiments in the laboratory.

Corrosive
These substances attack or destroy living tissues, including eyes and skin.

Harmful
These substances are similar to toxic substances but less dangerous.

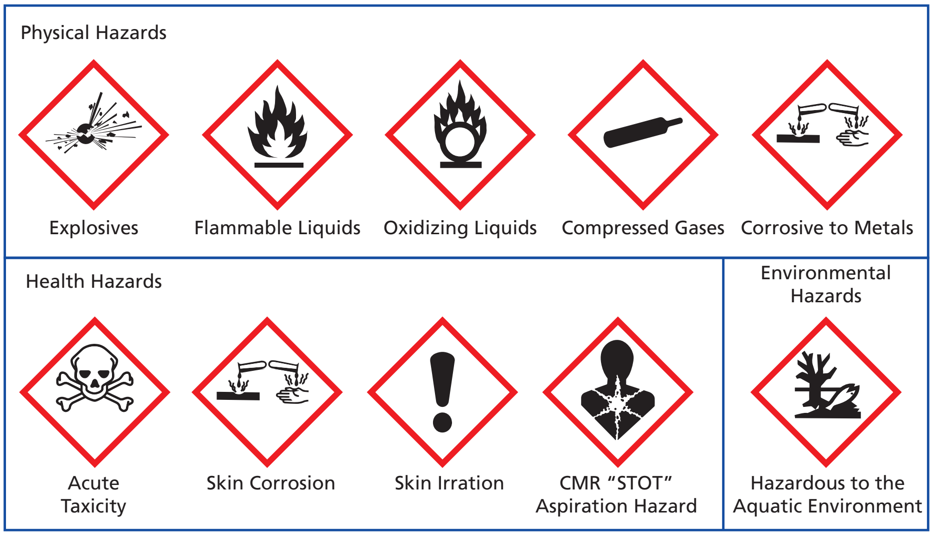
Irritant
These substances are not corrosive but can cause reddening or blistering of the skin.

Oxidising
These substances provide oxygen which allows other materials to burn more fiercely.

Toxic
These substances can cause death.

Highly flammable
These substances can easily catch fire.

Teachers and students should note that a new system for labelling hazards is being introduced between 2010 and 2015 and, in due course, you will need to become familiar with these new symbols:



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Preface to the reader

This textbook has been written to help you in your study of chemistry to Cambridge IGCSE. The different chapters in this book are split up into short topics. At the end of many of these topics are questions to test whether you have understood what you have read. At the end of each chapter there are larger study questions. Try to answer as many of the questions as you can as you come across them because asking and answering questions is at the heart of your study of chemistry.

Some questions in the style of Cambridge IGCSE examination papers are included at the end of the book. In many cases they are designed to test your ability to apply your chemical knowledge. The questions may provide certain facts and ask you to make an interpretation of them. In such cases, the factual information may not be covered in the text.

To help draw attention to the more important words, scientific terms are printed in **bold** the first time they are used. There are also checklists at the end of each chapter summarising the important points covered.

As you read through the book, you will notice three sorts of shaded area in the text.

Material highlighted in green is for the Cambridge IGCSE Extended curriculum.

Areas highlighted in yellow contain material that is not part of the Cambridge IGCSE syllabus. It is extension work and will not be examined.

Questions are highlighted by a box like this.

You will see from the box at the foot of this page that the book is divided into four different areas of chemistry: Starter, Physical, Inorganic and Organic chemistry. We feel, however, that some topics lead naturally on to other topics not in the same area. So you can, of course, read and study the chapters in your own preferred order and the colour coding will help you with this.

The accompanying **Revision CD-ROM** provides invaluable exam preparation and practice. We want to test your knowledge with interactive questions that cover both the Core and Extended curriculum. These are organised by syllabus topic.

Together, the textbook and CD-ROM will provide you with the information you need for the Cambridge IGCSE syllabus. We hope you enjoy using them.

Bryan Earl and Doug Wilford

We use different colours to define different areas of chemistry:

- 'starter' chapters – basic principles
- physical chemistry
- inorganic chemistry
- organic chemistry and the living world.

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1 The particulate nature of matter

Chemistry is about what **matter** is like and how it behaves, and our explanations and predictions of its behaviour. What is matter? This word is used to cover all the substances and materials from which the physical universe is composed. There are many millions of different substances known, and all of them can be categorised as solids, liquids or gases (Figure 1.1). These are what we call the **three states of matter**.

a solid

b liquid

c gas

Figure 1.1 Water in three different states.

Solids, liquids and gases

A **solid**, at a given temperature, has a definite volume and shape which may be affected by changes in temperature. Solids usually increase slightly in size when heated (**expansion**) (Figure 1.2) and usually decrease in size if cooled (**contraction**).

A **liquid**, at a given temperature, has a fixed volume and will take up the shape of any container into which it is poured. Like a solid, a liquid's volume is slightly affected by changes in temperature.

A **gas**, at a given temperature, has neither a definite shape nor a definite volume. It will take up the shape of any container into which it is placed and will spread out evenly within it. Unlike those of solids and liquids, the volumes of gases are affected quite markedly by changes in temperature.

Liquids and gases, unlike solids, are relatively **compressible**. This means that their volume can be reduced by the application of pressure. Gases are much more compressible than liquids.

Figure 1.2 Without expansion gaps between the rails, the track would buckle in hot weather.

The kinetic theory of matter

The **kinetic theory** helps to explain the way in which matter behaves. The evidence is consistent with the idea that all matter is made up of tiny **particles**. This theory explains the physical properties of matter in terms of the movement of its constituent particles.

The main points of the theory are:

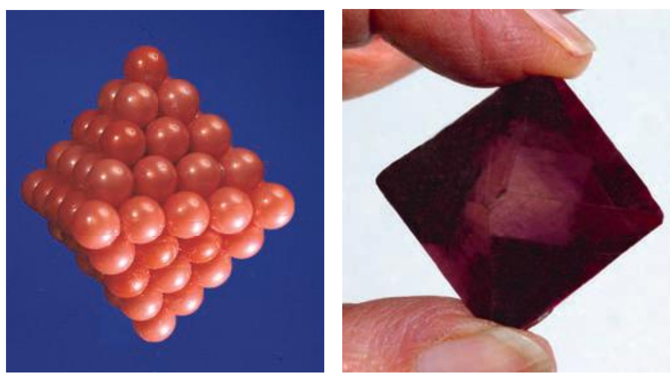
- All matter is made up of tiny, moving particles, invisible to the naked eye. Different substances have different types of particles (atoms, molecules or ions) which have different sizes.
- The particles move all the time. The higher the temperature, the faster they move on average.
- Heavier particles move more slowly than lighter ones at a given temperature.

The kinetic theory can be used as a scientific model to explain how the arrangement of particles relates to the properties of the three states of matter.

Explaining the states of matter

In a solid the particles attract one another. There are attractive forces between the particles which hold them close together. The particles have little freedom of movement and can only vibrate about a fixed position. They are arranged in a regular manner, which explains why many solids form crystals.

It is possible to model such crystals by using spheres to represent the particles (Figure 1.3a). If the spheres are built up in a regular way then the shape compares very closely with that of a part of a chrome alum crystal (Figure 1.3b).



a A model of a chrome alum crystal. **b** An actual chrome alum crystal.

Figure 1.3

Studies using X-ray crystallography (Figure 1.4) have confirmed how the particles are arranged in crystal structures. When crystals of a pure substance form under a given set of conditions, the particles present are always packed in the same way. However, the particles may be packed in different ways in crystals of different substances. For example, common salt (sodium chloride) has its particles arranged to give cubic crystals as shown in Figure 1.5.

Figure 1.4 A modern X-ray crystallography instrument, used for studying crystal structure.

Figure 1.5 Sodium chloride crystals.

In a liquid the particles are still close together but they move around in a random way and often collide with one another. The forces of attraction between the particles in a liquid are weaker than those in a solid. Particles in the liquid form of a substance have more energy on average than the particles in the solid form of the same substance.

In a gas the particles are relatively far apart. They are free to move anywhere within the container in which they are held. They move randomly at very high velocities, much more rapidly than those in a liquid. They collide with each other, but less often than in a liquid, and they also collide with the walls of the container. They exert virtually no forces of attraction on each other because they are relatively far apart. Such forces, however, are very significant. If they did not exist we could not have solids or liquids (see Changes of state, p. 4).

The arrangement of particles in solids, liquids and gases is shown in Figure 1.6.

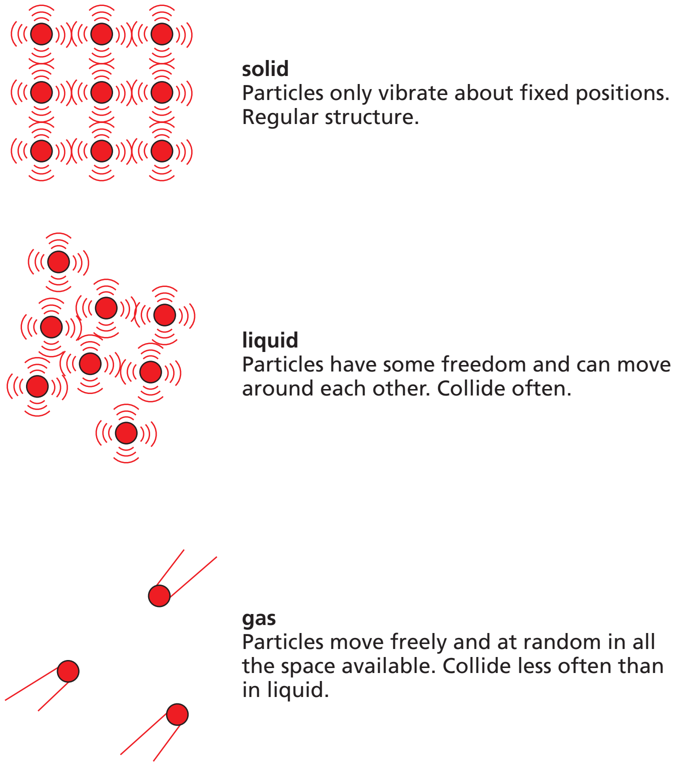


Figure 1.6 The arrangement of particles in solids, liquids and gases.

Questions

1. When a metal such as copper is heated it expands. Explain what happens to the metal particles as the solid metal expands.
2. Use your research skills on the Internet to find out about the technique of X-ray crystallography and how this technique can be used to determine the crystalline structure of solid substances such as sodium chloride.

Changes of state

The kinetic theory model can be used to explain how a substance changes from one state to another. If a solid is heated the particles vibrate faster as they gain energy. This makes them 'push' their neighbouring particles further away from themselves. This causes an increase in the volume of the solid, and the solid expands. Expansion has taken place.

Eventually, the heat energy causes the forces of attraction to weaken. The regular pattern of the structure breaks down. The particles can now move around each other. The solid has melted. The temperature at which this takes place is called the **melting point** of the substance. The temperature of a pure melting solid will not rise until it has all melted. When the substance has become a liquid there are still very significant forces of attraction between the particles, which is why it is a liquid and not a gas.

Solids which have high melting points have stronger forces of attraction between their particles than those which have low melting points. A list of some substances with their corresponding melting and boiling points is shown in Table 1.1.

Table 1.1

Substance	Melting point/°C	Boiling point/°C
Aluminium	661	2467
Ethanol	–117	79
Magnesium oxide	827	3627
Mercury	–30	357
Methane	–182	–164
Oxygen	–218	–183
Sodium chloride	801	1413
Sulfur	113	445
Water	0	100

If the liquid is heated the particles will move around even faster as their average energy increases. Some particles at the surface of the liquid have enough energy to overcome the forces of attraction between themselves and the other particles in the liquid and they escape to form a gas. The liquid begins to **evaporate** as a gas is formed.

Eventually, a temperature is reached at which the particles are trying to escape from the liquid so quickly that bubbles of gas actually start to form inside the bulk of the liquid. This temperature is called the **boiling point** of the substance. At the boiling point the pressure of the gas created above the liquid equals that in the air – **atmospheric pressure**.

Liquids with high boiling points have stronger forces between their particles than liquids with low boiling points.

When a gas is cooled the average energy of the particles decreases and the particles move closer together. The forces of attraction between the particles now become significant and cause the gas to **condense** into a liquid. When a liquid is cooled it **freezes** to form a solid. In each of these changes energy is given out.

Changes of state are examples of **physical changes**. Whenever a physical change of state occurs, the temperature remains constant during the change (see Heating and cooling curves, p. 5). During a physical change no new substance is formed.

An unusual state of matter

Liquid crystals are an unusual state of matter (Figure 1.7). These substances look like liquids and flow like liquids but have some order in the arrangement of the particles, and so in some ways they behave like crystals.

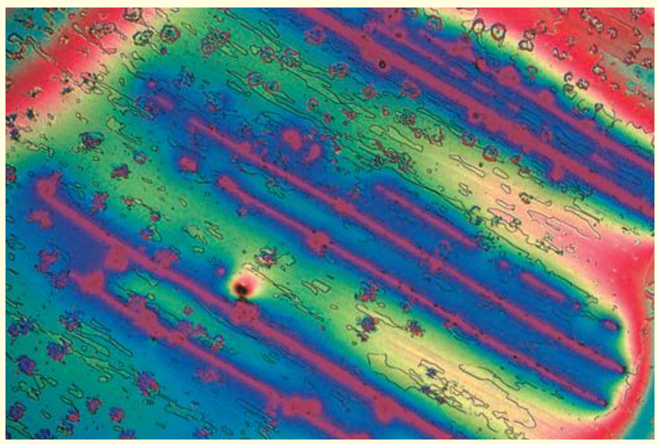


Figure 1.7 A polarised light micrograph of liquid crystals.

Liquid crystals are now part of our everyday life. They are widely used in displays for digital watches, calculators and lap-top computers, and in televisions (Figure 1.8). They are also useful in thermometers because liquid crystals change colour as the temperature rises and falls.

Figure 1.8 Liquid crystals are used in this TV screen.

An unusual change of state

There are a few substances that change directly from a solid to a gas when they are heated without ever becoming a liquid. This rapid spreading out of the particles is called **sublimation**. Cooling causes a change from a gas directly back to a solid. Examples of substances that behave in this way are carbon dioxide (Figure 1.9) and iodine.

$$\text{solid} \underset{\text{sublimation}}{\overset{\text{sublimation}}{\rightleftharpoons}} \text{gas}$$

Figure 1.9 Dry ice (solid carbon dioxide) sublimes on heating and can be used to create special effects on stage.

Carbon dioxide is a white solid called dry ice at temperatures below −78 °C. When heated to just above −78 °C it changes into carbon dioxide gas. The changes of state are summarised in Figure 1.10.

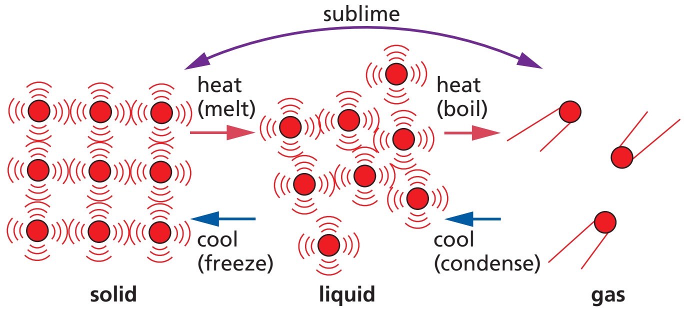


Figure 1.10 Summary of the changes of state.

Heating and cooling curves

The graph shown in Figure 1.11 was drawn by plotting the temperature of water as it was heated steadily from −15 °C to 110 °C. You can see from the curve that changes of state have taken place. When the temperature was first measured only ice was present. After a short time the curve flattens, showing that even though heat energy is being put in, the temperature remains constant.

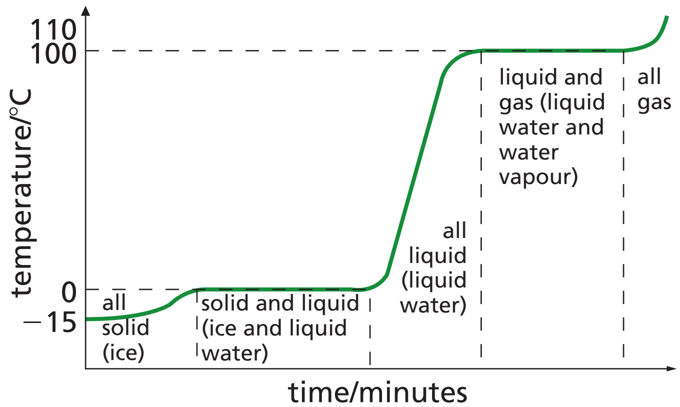


Figure 1.11 Graph of temperature against time for the change from ice at −15 °C to water to steam.

In ice the particles of water are close together and are attracted to one another. For ice to melt the particles must obtain sufficient energy to overcome the forces of attraction between the water particles to allow relative movement to take place. This is where the heat energy is going.

The temperature will begin to rise again only after all the ice has melted. Generally, the heating curve for a pure solid always stops rising at its melting point and gives rise to a sharp melting point. A sharp melting point indicates a pure sample. The addition or presence of impurities lowers the melting point. You can try to find the melting point of a substance using the apparatus shown in Figure 1.12.

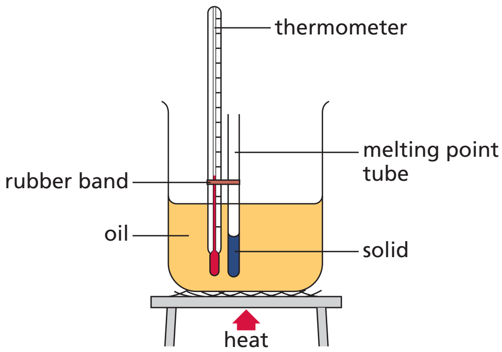


Figure 1.12 Apparatus shown here if heated slowly can be used to find the melting point of a substance such as the solid in the melting point tube.

In the same way, if you want to boil a liquid such as water you have to give it some extra energy. This can be seen on the graph (Figure 1.11) where the curve levels out at 100 °C – the boiling point of water.

Solids and liquids can be identified from their characteristic melting and boiling points.

The reverse processes of condensing and freezing occur on cooling. This time, however, energy is given out when the gas condenses to the liquid and the liquid freezes to give the solid.

Questions

1 Write down as many uses as you can for liquid crystals.
2 Why do gases expand more than solids for the same increase in temperature?
3 Ice on a car windscreen will disappear as you drive along, even without the heater on. Explain why this happens.
4 When salt is placed on ice the ice melts. Explain why.
5 Draw and label the graph you would expect to produce if water at 100 °C was allowed to cool to –5 °C.

Diffusion – evidence for moving particles

When you walk past a cosmetics counter in a department store you can usually smell the perfumes. For this to happen gas particles must be leaving open perfume bottles and be spreading out through the air in the store. This spreading out of a gas is called **diffusion** and it takes place in a haphazard and random way.

All gases diffuse to fill the space available. In Figure 1.13, after a day the brown–red fumes of gaseous bromine have spread evenly throughout both gas jars from the liquid present in the lower gas jar.

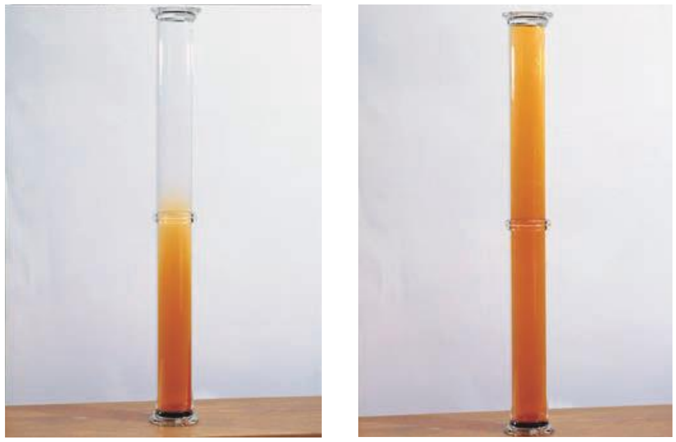


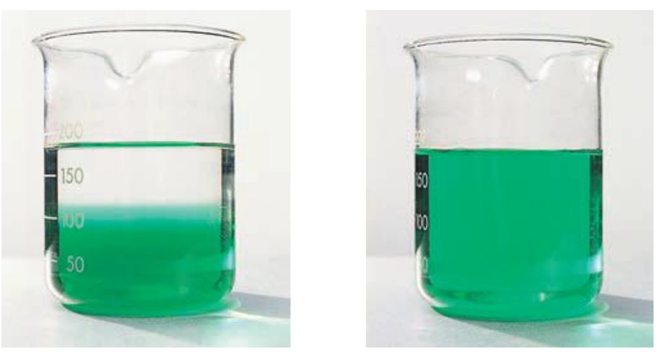
Figure 1.13 After 24 hours the bromine fumes have diffused throughout both gas jars.

Gases diffuse at different rates. If one piece of cotton wool is soaked in concentrated ammonia solution and another is soaked in concentrated hydrochloric acid and these are put at opposite ends of a dry glass tube, then after a few minutes a white cloud of ammonium chloride appears (Figure 1.14). This shows the position at which the two gases meet and react. The white cloud forms in the position shown because the ammonia particles are lighter and have a smaller relative molecular mass (Chapter 4, p. 62) than the hydrogen chloride particles (released from the hydrochloric acid) and so move faster.

Diffusion also takes place in liquids (Figure 1.15) but it is a much slower process than in gases. This is because the particles of a liquid move much more slowly.

When diffusion takes place between a liquid and a gas it is known as **intimate mixing**. The kinetic theory can be used to explain this process. It states that collisions are taking place randomly between particles in a liquid or a gas and that there is sufficient space between the particles of one substance for the particles of the other substance to move into.

Figure 1.14 Hydrochloric acid (left) and ammonia (right) diffuse at different rates.



a **b**

Figure 1.15 Diffusion within nickel(II) sulfate solution can take days to reach the stage shown on the right.

Questions

1 When a jar of coffee is opened, people in all parts of the room soon notice the smell. Use the kinetic theory to explain how this happens.
2 Describe, with the aid of diagrams, the diffusion of nickel(II) sulfate solution.
3 Explain why diffusion is faster in gases than in liquids.

Brownian motion

Evidence for the movement of particles in liquids came to light in 1827 when a botanist, Robert Brown, observed that fine pollen grains on the surface of water were not stationary. Through his microscope he noticed that the grains were moving about in a random way. It was 96 years later, in 1923, that another scientist called Norbert Wiener explained what Brown had observed. He said that the pollen grains were moving because the much smaller and faster-moving water particles were constantly colliding with them (Figure 1.16a).

This random motion of visible particles (pollen grains) caused by much smaller, invisible ones (water particles) is called **Brownian motion** (Figure 1.16b), after the scientist who first observed this phenomenon. It was used as evidence for the kinetic particle model of matter (p. 3).

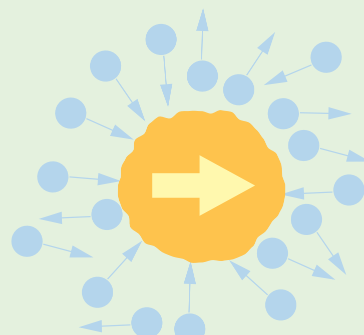


Figure 1.16a Pollen particle being bombarded by water molecules.

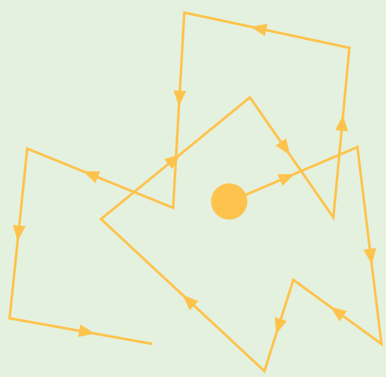


Figure 1.16b Brownian motion causes the random motion of the visible particle.

Checklist

After studying Chapter 1 you should know and understand the following terms.

- **Atmospheric pressure** The pressure exerted by the atmosphere on the surface of the Earth due to the weight of the air.
- **Boiling point** The temperature at which the pressure of the gas created above a liquid equals atmospheric pressure.
- **Condensation** The change of a vapour or a gas into a liquid. This process is accompanied by the evolution of heat.
- **Diffusion** The process by which different substances mix as a result of the random motions of their particles.
- **Evaporation** A process occurring at the surface of a liquid involving the change of state of a liquid into a vapour at a temperature below the boiling point.
- **Kinetic theory** A theory which accounts for the bulk properties of matter in terms of the constituent particles.
- **Matter** Anything which occupies space and has a mass.
- **Melting point** The temperature at which a solid begins to liquefy. Pure substances have a sharp melting point.
- **Solids, liquids and gases** The three states of matter to which all substances belong.
- **Sublimation** The direct change of state from solid to gas and the reverse process.

The particulate nature of matter

Additional questions

1 **a** Draw diagrams to show the arrangement of particles in:
- **(i)** solid lead
- **(ii)** molten lead
- **(iii)** gaseous lead.

b Explain how the particles move in these three states of matter.

c Explain, using the kinetic theory, what happens to the particles in oxygen as it is cooled down.

2 Explain the meaning of each of the following terms. In your answer include an example to help with your explanation.

a Expansion.
b Contraction.
c Physical change.
d Sublimation.
e Diffusion.
f Random motion.

3 **a** Why do solids not diffuse?

b Give two examples of diffusion of gases and liquids found in the house.

4 Use the kinetic theory to explain the following:

a When you take a block of butter out of the fridge, it is quite hard. However, after 15 minutes it is soft enough to spread.

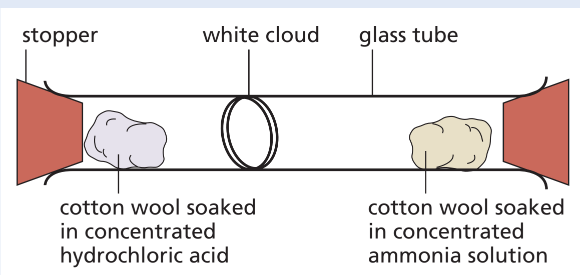
b When you come home from school and open the door you can smell your tea being cooked.

c A football is blown up until it is hard on a hot summer's day. In the evening the football feels softer.

d When a person wearing perfume enters a room it takes several minutes for the smell to reach the back of the room.

e A windy day is a good drying day.

5 The apparatus shown below was set up.



Give explanations for the following observations.

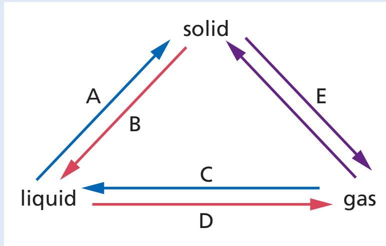
a The formation of a white cloud.

b It took a few minutes before the white cloud formed.

c The white cloud formed further from the cotton wool soaked in ammonia.

d Cooling the concentrated ammonia and hydrochloric acid before carrying out the experiment increased the time taken for the white cloud to form.

6 The following diagram shows the three states of matter and how they can be interchanged.



a Name the changes **A** to **E**.

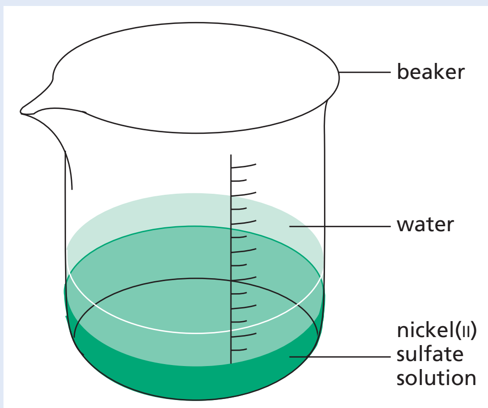
b Name a substance which will undergo change **E**.

c Name a substance which will undergo changes from solid to liquid to gas between 0 °C and 100 °C.

d Describe what happens to the particles of the solid during change **E**.

e Which of the changes **A** to **E** will involve:
- **(i)** an input of heat energy?
- **(ii)** an output of heat energy?

7 Some nickel(II) sulfate solution was carefully placed in the bottom of a beaker of water. The beaker was then covered and left for several days.



a Describe what you would see after:
- **(i)** a few hours
- **(ii)** several days.

b Explain your answer to **a** using your ideas of the kinetic theory of particles.

c What is the name of the physical process that takes place in this experiment?

2 Elements, compounds and experimental techniques

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

Figure 2.1 The planets in the universe are made of millions of substances. These are made up mainly from just 91 elements which occur naturally on the Earth.

In 1803, John Dalton (Figure 2.2) suggested that each element was composed of its own kind of particles, which he called **atoms**. Atoms are much too small to be seen. We now know that about 20×10^6 of them would stretch over a length of only 1 cm.

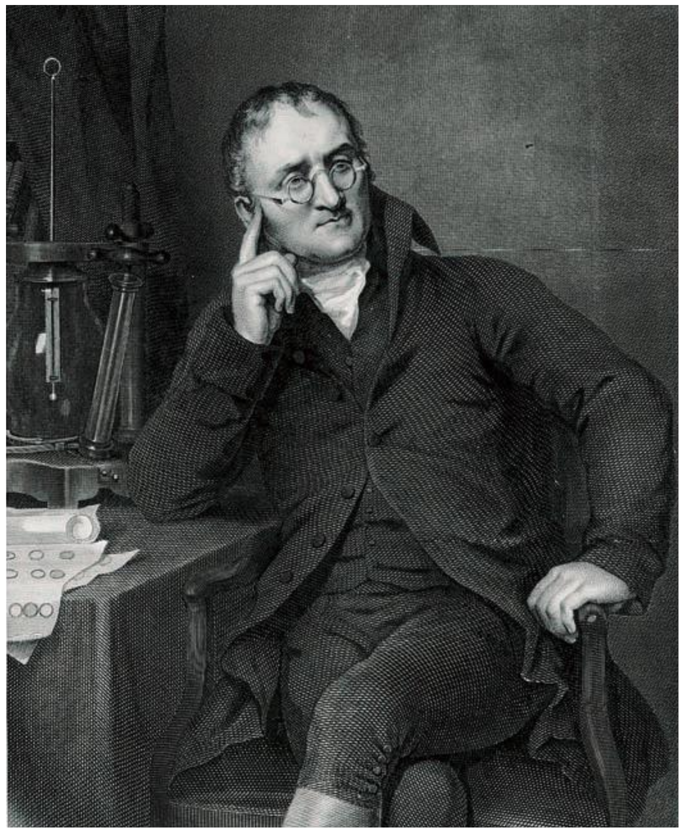


Figure 2.2 John Dalton (1766–1844).

Elements

Robert Boyle used the name element for any substance that cannot be broken down further, into a simpler substance. This definition can be extended to include the fact that each element is made up of only one kind of atom. The word atom comes from the Greek word *atomos* meaning ‘unsplittable’.

For example, aluminium is an element which is made up of only aluminium atoms. It is not possible to obtain a simpler substance chemically from the aluminium atoms. You can only make more complicated substances from it, such as aluminium oxide, aluminium nitrate or aluminium sulfate.

There are 118 elements which have now been identified. Twenty-seven of these do not occur in nature and have been made artificially by scientists. They include elements such as curium and unnilpentium. Ninety-one of the elements occur naturally and range from some very reactive gases, such as fluorine and chlorine, to gold and platinum, which are unreactive elements.

All elements can be classified according to their various properties. A simple way to do this is to classify them as **metals** or **non-metals** (Figures 2.3 and 2.4, p. 12). Table 2.1 shows the physical data for some common metallic and non-metallic elements.

You will notice that many metals have high densities, high melting points and high boiling points, and that most non-metals have low densities, low melting points and low boiling points. Table 2.2 summarises the different properties of metals and non-metals.

A discussion of the chemical properties of metals is given in Chapters 9 and 10. The chemical properties of certain non-metals are discussed in Chapters 9, 12 and 13.

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

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

Source: Earl B., Wilford L.D.R. Chemistry data book. Nelson Blackie, 1991 [a] At –254 °C [b] At –197 °C [c] At –184 °C.

a Gold is very decorative.

b Aluminium has many uses in the aerospace industry.

c These coins contain nickel.

Figure 2.3 Some metals.

Table 2.2 How the properties of metals and non-metals compare.

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

a A premature baby needs oxygen.

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

c Neon is used in advertising signs

Figure 2.4 Some non-metals.

Atoms – the smallest particles

Everything is made up of billions of atoms. The atoms of all elements are extremely small; in fact they are too small to be seen. The smallest atom known is hydrogen, with each atom being represented as a sphere having a diameter of 0.000 000 07 mm (or 7×10^{-8} mm) (Table 2.3). Atoms of different elements have different diameters as well as different masses. How many atoms of hydrogen would have to be placed side by side along the edge of your ruler to fill just one of the 1 mm divisions?

Table 2.3 Sizes of atoms.

Atom	Diameter of atom/mm
Hydrogen	7×10^{-8}
Oxygen	12×10^{-8}
Sulfur	20.8×10^{-8}

Chemists use shorthand symbols to label the elements and their atoms. The symbol consists of one, two or three letters, the first of which must be a capital. Where several elements have the same initial letter, a second letter of the name or subsequent letter is added. For example, **C** is used for **carbon**, **Ca** for **calcium** and **Cl** for **chlorine**. Some symbols seem to have no relationship to the name of the element, for example **Na** for **sodium** and **Pb** for **lead**. These symbols come from their Latin names, natrium for sodium and plumbum for lead. A list of some common elements and their symbols is given in Table 2.4.

Molecules

The atoms of some elements are joined together in small groups. These small groups of atoms are called **molecules**. For example, the atoms of the elements hydrogen, oxygen, nitrogen, fluorine, chlorine, bromine and iodine are each joined in pairs and they are known as **diatomic** molecules. In the case of phosphorus and sulfur the atoms are joined in larger numbers, four and eight respectively (P_4, S_8). In chemical shorthand the molecule of chlorine shown in Figure 2.5 is written as Cl_2.

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

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

The complete list of the elements with their corresponding symbols is shown in the Periodic Table on p. 294.

The gaseous elements helium, neon, argon, krypton, xenon and radon are composed of separate and individual atoms. When an element exists as separate atoms, then the molecules are said to be **monatomic**. In chemical shorthand these monatomic molecules are written as He, Ne, Ar, Kr, Xe and Rn respectively.

Cl—Cl

a As a letter-and-stick model.

b As a space-filling model.

Figure 2.5 A chlorine molecule.

Molecules are not always formed by atoms of the same type joining together. For example, water exists as molecules containing oxygen and hydrogen atoms.

Questions

1. How would you use a similar chemical shorthand to write a representation of the molecules of iodine and fluorine?
2. Using the Periodic Table on p. 294 write down the symbols for the following elements and give their physical states at room temperature:
   a chromium **b** krypton **c** osmium.

Compounds

Compounds are pure substances which are formed when two or more elements chemically combine together. Water is a simple compound formed from the elements hydrogen and oxygen (Figure 2.6). This combining of the elements can be represented by a word equation:

$$\text{hydrogen} + \text{oxygen} \rightarrow \text{water}$$

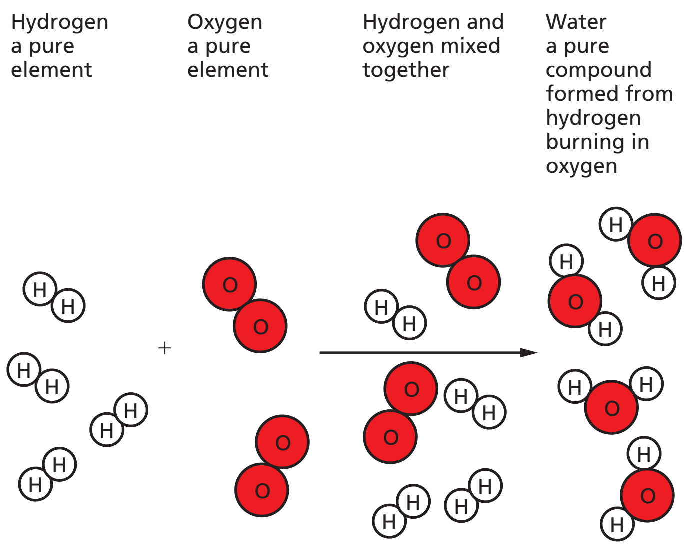


Figure 2.6 The element hydrogen reacts with the element oxygen to produce the compound water.

Water molecules contain two atoms of hydrogen and one atom of oxygen, and hence water has the **chemical formula** H_2O. Elements other than hydrogen will also react with oxygen to form compounds called oxides. For example, magnesium reacts violently with oxygen gas to form the white powder magnesium oxide (Figure 2.7). This reaction is accompanied by a release of energy as new chemical bonds are formed.

Figure 2.7 Magnesium burns brightly in oxygen to produce magnesium oxide.

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

magnesium + oxygen → magnesium oxide

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

Reduction is the opposite of oxidation. In this process oxygen is removed instead of being added.

A **redox** reaction is one which involves the two processes of reduction and oxidation. For example, the oxygen has to be removed in the extraction of iron from iron(III) oxide. This can be done in a blast furnace with carbon monoxide. The iron(III) oxide loses oxygen to the carbon monoxide and is reduced to iron. Carbon monoxide is the **reducing agent**. A reducing agent is a substance that reduces another substance during a redox reaction. Carbon monoxide is oxidised to carbon dioxide by the iron(III) oxide. The iron(III) oxide is the **oxidising agent**. An oxidising agent is a substance which oxidises another substance during a redox reaction.

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

For a further discussion of oxidation and reduction see Chapter 3 (p. 39) and Chapter 5 (p. 73).

Both **reduction** and **oxidation** have taken place in this chemical process, and so this is known as a **redox** reaction.

More about formulae

The formula of a compound is made up from the symbols of the elements present and numbers to show the ratio in which the different atoms are present. Carbon dioxide has the formula CO_2. This tells you that it contains one carbon atom for every two oxygen atoms. The 2 in the formula tells you that there are two oxygen atoms present in each molecule of carbon dioxide. For further discussion see p. 43.

Table 2.5 shows the names and formulae of some common compounds which you will meet in your study of chemistry.

Table 2.5 Names and formulae of some common compounds.

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

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

Balancing chemical equations

Word equations are a useful way of representing chemical reactions but a better and more useful method is to produce a **balanced chemical equation**. This type of equation gives the formulae of the reactants and the products as well as showing the relative numbers of each particle involved.

Balanced equations often include the physical state symbols:

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

The word equation to represent the reaction between iron and sulfur is:

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

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

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

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

In the case of magnesium reacting with oxygen, the word equation was:

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

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

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

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

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

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

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

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

Instrumental techniques

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

Many of the instrumental methods that have been developed are quite sophisticated. Some methods are suited to identifying elements. For example, atomic absorption spectroscopy allows the element to be identified and also allows the quantity of the element that is present to be found (Figure 2.8).

Figure 2.8 This instrument allows the quantity of a particular element to be found. It is used extensively throughout industry for this purpose. It will allow even tiny amounts of a particular element to be found.

Some methods are particularly suited to the identification of compounds. For example, infrared spectroscopy is used to identify compounds by showing the presence of particular groupings of atoms (Figure 2.9).

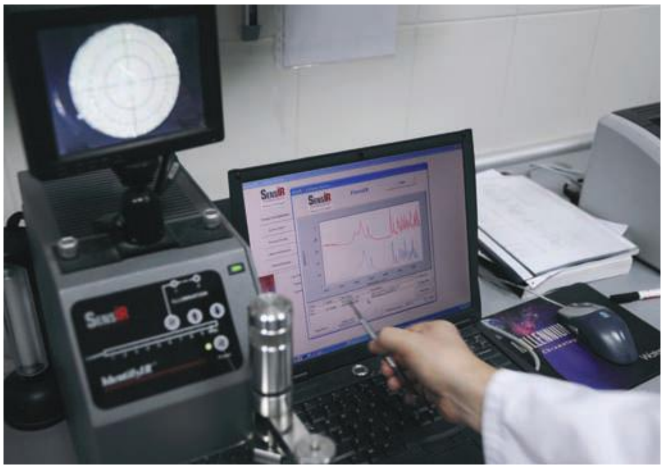


Figure 2.9 This is a modern infrared spectrometer. It is used in analysis to obtain a so-called fingerprint spectrum of a substance that will allow the substance to be identified.

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

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

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

Questions

1 Write the word and balanced chemical equations for the reactions which take place between:
 a calcium and oxygen b copper and oxygen.
2 Write down the ratio of the atoms present in the formula for each of the compounds shown in Table 2.5.
3 Iron is extracted from iron(III) oxide in a blast furnace by a redox reaction. What does the term 'redox reaction' mean?
4 Identify the oxidising and reducing agents in the following reactions:
 a copper(II) oxide + hydrogen → copper + water
 b tin(II) oxide + carbon → tin + carbon dioxide
 c $PbO(s) + H_2(g) \rightarrow Pb(s) + H_2O(l)$

Mixtures

Many everyday things are not pure substances, they are mixtures. A mixture contains more than one substance (elements and/or compounds). An example of a common mixture is sea water (Figure 2.10).

Figure 2.10 Sea water is a common mixture.

Other mixtures include the air, which is a mixture of elements such as oxygen, nitrogen and neon and compounds such as carbon dioxide (see Chapter 11, p. 173), and alloys such as brass, which is a mixture of copper and zinc (for a further discussion of alloys see Chapter 10, p. 165).

What is the difference between mixtures and compounds?

There are differences between compounds and mixtures. This can be shown by considering the reaction between iron filings and sulfur. A mixture of iron filings and sulfur looks different from the individual elements (Figure 2.11). This mixture has the properties of both iron and sulfur; for example, a magnet can be used to separate the iron filings from the sulfur (Figure 2.12).

Substances in a mixture have not undergone a chemical reaction and it is possible to separate them provided that there is a suitable difference in their physical properties. If the mixture of iron and sulfur is heated a chemical reaction occurs and a new substance is formed called iron(II) sulfide (Figure 2.11). The word equation for this reaction is:

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