SECOND EDITION

Chemistry

for the
IB DIPLOMA

Christopher Talbot,
Richard Harwood
and Christopher Coates

Dynamic Learning

Hodder Education

Chemistry

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Contents

Options

Available on the website accompanying this book: www.hoddereducation.com/IBextras

Answers and glossary

Answers and glossary appear on the website accompanying this book:
www.hoddereducation.com/IBextras

Introduction

Welcome to the second edition of *Chemistry for the IB Diploma*. The content and structure of this second edition has been completely revised to meet the demands of the 2014 *IB Diploma Programme Chemistry Guide*.

Within the IB Diploma Programme, the chemistry content is organized into compulsory core topics plus a number of options, from which all students select one. The organization of this resource exactly follows the *IB Chemistry Guide* sequence:

- **Core:** Chapters 1–11 cover the common core topics for Standard *and* Higher Level students.
- **Additional Higher Level (AHL):** Chapters 12–21 cover the additional topics for Higher Level students.
- **Options:** Chapters 22–25 cover Options A, B, C and D respectively. Each of these is available to both Standard and Higher Level students. (Higher Level students study more topics within the same option.) These are available on the Hodder website.

The syllabus is presented as topics, each of which (for the core and AHL topics) is the subject of a corresponding single chapter in the *Chemistry for the IB Diploma* printed book.

The Options (Chapters 22–25) are available on the website accompanying this book, as are a comprehensive Glossary and the answers to the end-of-chapter exam and exam-style questions: www.hoddereducation.com/IBextras.

Special features of the chapters of *Chemistry for the IB Diploma* are:

- Each chapter begins with **Essential Ideas** that summarize the concepts on which it is based.
- The text is written in **straightforward language**, without phrases or idioms that might confuse students for whom English is a second language. The text is also suitable for students of all abilities.
- The **depth of treatment** of topics has been carefully planned to accurately reflect the objectives of the IB syllabus and the requirements of the examinations.
- Photographs and full-colour illustrations support the relevant text, with annotations which elaborate on the context, function, language, history or applications of chemistry.

- The **Nature of Science** is an important new aspect of the IB Chemistry course, which aims to broaden students' interests and knowledge beyond the confines of its specific chemistry content. Throughout this book we hope that students will develop an appreciation of the processes and applications of chemistry and technology. Some aspects of the *Nature of Science* may be examined in IB Chemistry examinations and important discussion points are highlighted in the margins.
- The **Utilizations** and **Additional Perspectives** sections also reflect the Nature of Science, but they are designed to take students beyond the limits of the IB syllabus in a variety of ways. They may, for example, provide a historical context, extend theory or offer an interesting application. They are sometimes accompanied by more challenging, or research style, questions. They do not contain any knowledge which is essential for the IB examinations.
- Science and technology have developed over the centuries with contributions from scientists from all around the world. In the modern world science knows few boundaries and the flow of information is usually quick and easy. Some international applications of science have been indicated with the **International Mindedness** icon.
- **Worked examples** are provided in each chapter whenever new equations are introduced. A large number of **self-assessment questions** and some research questions are also placed throughout the chapters close to the relevant theory. They are phrased in order to assist comprehension and recall, and to help familiarize students with the assessment implications of the command terms.
- It is not an aim of this book to provide detailed information about experimental work or the use of computers. However, our **Applications and Skills** icon has been placed in the margin to indicate wherever such work may usefully aid understanding.
- A selection of **IB examination-style questions** are provided at the end of each chapter, as well as some past IB Chemistry examination questions. **Answers** to these are provided on the website accompanying this book.

- Extensive links to the interdisciplinary **Theory of Knowledge (ToK)** element of the IB Diploma course, including ethics, are made in all chapters.
- Comprehensive **glossaries** of words and terms, including IB command terms, for Core and AHL topics are included in the website which accompanies this book.
- This icon denotes links to material available on the website that accompanies this book: www.hoddereducation.com/IBextras.

■ Using this book

The sequence of chapters in *Chemistry for the IB Diploma* deliberately follows the sequence of the syllabus content. However, the *IB Diploma Chemistry Guide* is not designed as a teaching syllabus, so the order in which the syllabus content is presented is not necessarily the order in which it will be taught. Different schools and colleges should design course delivery based on their individual circumstances.

In addition to the study of the chemistry principles contained in this book, IB science students carry out experiments and investigations, as well as collaborating in a Group 4 Project. These are assessed within the school (Internal Assessment) based on well-established criteria.

■ Author profiles

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Chris teaches IB Chemistry and ToK at a leading IB World School in Singapore. He has also taught IB Biology, MYP Science and a variety of IGCSE Science courses. He has moderated IB Chemistry coursework and prepared students for the Singapore Chemistry Olympiad.

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■ Authors' acknowledgements

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Chris Talbot
Singapore, June 2015

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1 Stoichiometric relationships

1.1 Introduction to the particulate nature of matter and chemical change – *physical and chemical properties depend on the ways in which different atoms combine*

Chemistry is the study of chemical substances. The collective name for chemical substances is matter. Matter may be in the form of a solid, a liquid, or a gas. These are called the three states of matter and are convertible.

Matter may contain one chemical substance or a mixture of different chemical substances. Part of a chemist's work is to separate one substance from another and to identify single or pure substances.

■ **Figure 1.1** Hawaii National Park with volcano emitting steam (temperature above 100 °C), above which are clouds of water vapour (air temperature)

■ States of matter

There are three phases or **states of matter**: **solids**, **liquids** and **gases**. Any substance can exist in each of these three states depending on temperature and pressure.

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 (in all directions) when heated (thermal expansion) and usually decrease in size if cooled (thermal contraction).

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

A gas (Figure 1.1), at a given temperature, has neither a definite shape nor a definite volume. It will take up the shape of any container and will spread out evenly within it, by a process known as diffusion. The volumes of gases are greatly affected by changes in temperature.

Liquids and gases, unlike solids, are relatively compressible. This means that their volumes are decreased by applying pressure. Gases are much more compressible than liquids.

■ **Figure 1.2**
A sample of the element phosphorus (red allotropic form)

■ Elements

The chemical **elements** (Figure 1.2) are the simplest substances and are each composed of a single type of atom (see Chapter 2). (Many elements exist as a mixture of atoms of differing masses, known as isotopes – see Chapter 2). Elements cannot be split up or decomposed into simpler substances by a chemical reaction.

The elements can be classified into three groups based upon the state of matter they exist in at 25 °C. Most of the elements are solids, for example iron, but bromine and mercury are liquids at room temperature and the remainder of the elements are gases, for example oxygen and neon.

O=O

oxygen molecule O_2

N≡N

nitrogen molecule N_2

H—H

hydrogen molecule H_2

Cl—Cl

chlorine molecule Cl_2
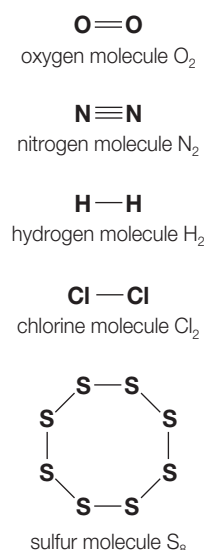
sulfur molecule S_8

■ **Figure 1.3**
Diagram of oxygen, nitrogen, hydrogen, chlorine and sulfur molecules

The elements can also be classified into two groups: **metals** and **non-metals** (see Chapter 4), based on their chemical and physical properties. For example, aluminium is a metal and chlorine is a non-metal.

Many elements exist as atoms, for example metals and the noble gases. However, many non-metals exist as atoms bonded together into **molecules** (Figure 1.3). Examples of non-metal molecules include oxygen, O_2, chlorine, Cl_2, nitrogen, N_2, phosphorus, P_4, and sulfur, S_8. Oxygen, nitrogen and chlorine exist as **diatomic** molecules.

Allotropy is the existence of two or more crystalline forms of an element. These different forms are called **allotropes**. Allotropes exist where there is more than one possible arrangement of bonded atoms. For example, solid carbon can exist in three allotropes: diamond, carbon-60 (C_{60}) or buckminsterfullerene, graphite (and graphene which is a single layer of graphite) (see Chapter 4); oxygen can exist in two allotropes: dioxygen (O_2) and trioxygen (ozone, O_3).

ToK Link

Priestley and Lavoisier's discovery of oxygen

Oxygen was first prepared in a reasonably pure state in the 18th century, and its preparation was followed by a theory of burning (combustion) which is still accepted. It completely replaced a theory called the phlogiston theory in a paradigm shift. This occurs when a scientific model or way of thinking is quickly and completely replaced by a very different scientific model or way of thinking.

Priestley strongly heated a red powder (mercury oxide) which he called calx of mercury. This substance decomposed into two substances: mercury and a gas (now known to be oxygen). He also discovered that flammable substances burned much more strongly in this gas (100% oxygen) than in normal air (20% oxygen). Priestley informed Lavoisier of his discovery, and Lavoisier carried out an experiment (Figure 1.4) in which he demonstrated that the gas which Priestley had made was identical to that 20% of the air which supports combustion (burning).

He kept the mercury in the retort, at a temperature just lower than its boiling point, for several days. He observed that the volume of gas had been reduced by 20%, this being shown by a rise in the level of the mercury in the bell jar. He also observed that a red powder (mercury oxide) had been formed on the surface of the hot mercury in the retort. The gas (now known to be nitrogen and noble gases) remaining in his apparatus would not support combustion.
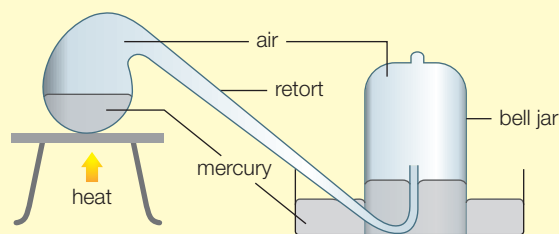
■ **Figure 1.4** Lavoisier's preparation of oxygen

On the basis of his and Priestley's observations, Lavoisier proposed the following explanation of combustion and the composition of air: 20% of air consists of oxygen; when substances burn they chemically combine with oxygen, forming oxides. When a substance burns completely, the mass of the oxide formed equals the combined mass of the original substance and the mass of the oxygen with which it has chemically combined.

■ Compounds

Many mixtures of elements undergo a chemical reaction when they are mixed together and heated. The formation of a **compound** (Figure 1.5) from its elements is termed **synthesis**. Heat energy is usually released during this reaction (see Chapter 5).

■ **Figure 1.5**
A model showing the structure of the compound calcium carbonate, $CaCO_3$ (black spheres represent carbon, red oxygen and clear calcium)
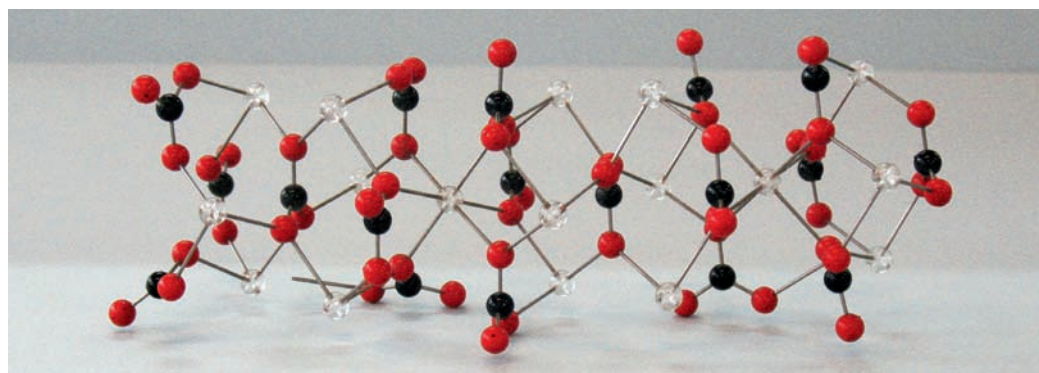
When a mixture of iron and sulfur is heated, large amounts of heat energy are released as the compound iron(II) sulfide, FeS, is formed (Figure 1.6). (Synthesis reactions like this are examples of redox reactions – see Chapter 9). Figure 1.6 describes this reaction in terms of atoms in iron and sulfur (Figure 1.7) reacting to form iron(II) sulfide.

■ **Figure 1.6**
A description of the formation of iron(II) sulfide in terms of atoms

iron atoms   sulfur atoms   iron(II) sulfide
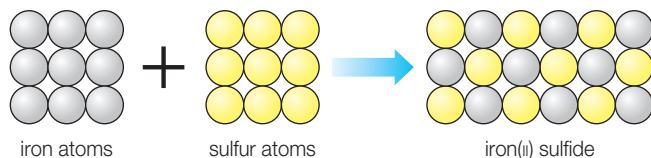
The word equation for this reaction is:

iron + sulfur → iron(II) sulfide
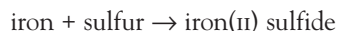
Mixtures of elements are easily separated by a physical method, since the atoms of the different elements are not bonded together. For example, iron can be separated from sulfur by the use of a magnet.

However, when a compound is formed the atoms it contains are chemically bonded together, so the compound will have different physical and chemical properties from the constituent elements (Table 1.1). For example, iron is magnetic, but the compound iron(II) sulfide is non-magnetic (Figure 1.8). A compound will contain either molecules or ions (Chapter 4).

■ **Table 1.1**
A summary of the different properties of iron, sulfur, an iron/sulfur mixture and iron(II) sulfide

Substance	Appearance	Effect of a magnet	Effect of dilute hydrochloric acid
Iron	Dark grey powder	Attracted to it	Very little reaction when cold. When warm, an odourless gas (hydrogen) is produced
Sulfur	Yellow powder	No effect	No reaction when hot or cold
Iron–sulfur mixture	Dirty yellow powder	Iron powder particles attracted to it	Iron powder reacts as described above
Iron(II) sulfide	Dark grey solid	No effect	A foul smelling gas (hydrogen sulfide) is produced

The splitting of a chemical compound into its constituent elements is termed **decomposition**. This process requires an input of energy, either heat (**thermal decomposition**) or electricity (electrolysis) (Chapter 9).
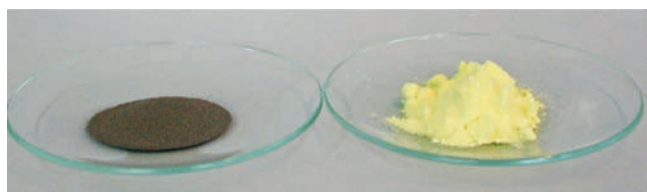
■ **Figure 1.7** The elements iron and sulfur
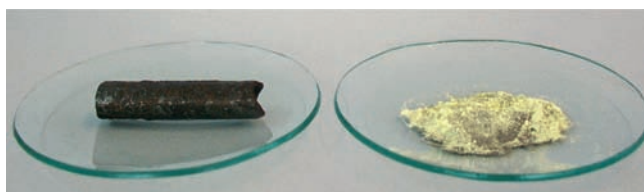
■ **Figure 1.8** A sample of iron(II) sulfide and a mixture of iron and sulfur

1 A mixture of magnesium and iodine was heated. A red glow spread through the mixture during the reaction. At the end of the experiment a white solid had been formed.
   a State one observation which shows that a chemical reaction has occurred.
   b Write a word equation to describe the reaction.
   c State two differences between compounds and elements.

■ Molecular kinetic theory

The simple diagram in Figure 1.9 shows the relationship between the states of matter and the arrangement (idealized, simplified and in two dimensions only) of their particles (ions, atoms or molecules). The arrows represent **physical changes** termed **changes of state**. In a physical change no new chemical substance is formed.

In a crystalline solid the particles (atoms, ions or molecules) are close together and packed in a regular pattern (known as a lattice). Studies using X-ray crystallography have confirmed how particles are arranged in crystal structures (see Chapter 22 on the accompanying website.)

■ **Figure 1.9**
The three states of matter and their interconversion
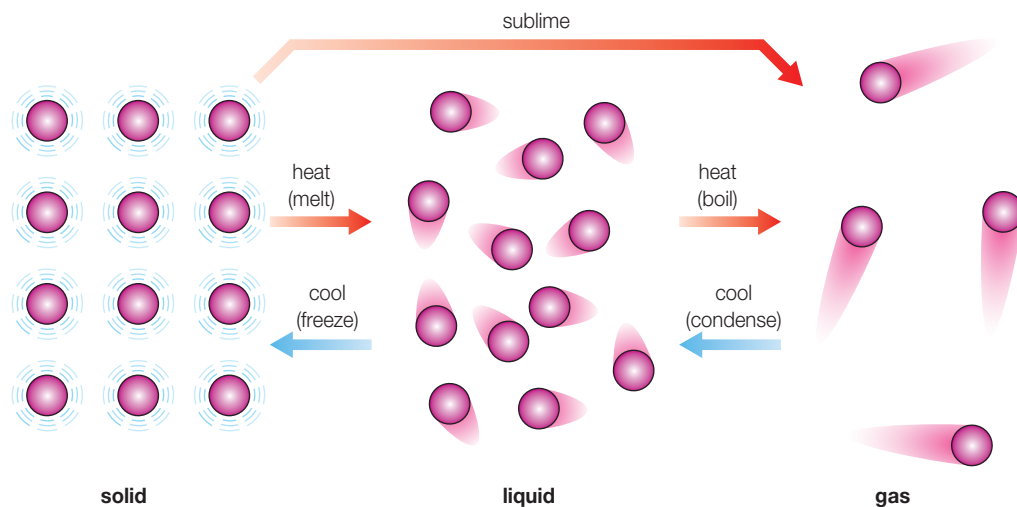
The particles vibrate around fixed positions and these vibrations become stronger as the temperature increases. The particles in a solid are strongly attracted to each other. In a liquid the particles are close together, but are free to move within the liquid. They are attracted to the other particles in the liquid. The particles move faster as the temperature increases. In a gas the particles are far apart and are free to move. The particles move so fast that there is little attraction between gas particles. The particles travel faster as the temperature increases.

This model about the way in which particles behave in the three states of matter is known as kinetic molecular theory. It describes all substances as being made up of particles in motion. It is a scientific model that explains how the arrangement of particles relates to the physical properties of the three states of matter.

■ Changes of state

The kinetic molecular model can be used to explain how a pure substance changes from one state of matter to another. If a crystalline solid is heated, the particles (atoms, ions or molecules) vibrate faster and with greater amplitude as they gain kinetic energy. This makes them 'push' their neighbouring particles further away from themselves. This causes an increase in the volume of the solid, which expands.

Eventually with a further increase in temperature the heat energy causes the forces of attraction between particles to weaken. The regular pattern of the particles in the lattice breaks down and the particles can now move around each other. The solid has melted and the temperature at which this occurs is the melting point. The temperature of a pure solid will remain constant until it has all melted. When the substance is a liquid there are still strong attractive forces operating between the particles.

There are energy changes which occur during changes of state. During **melting** and **boiling** heat is absorbed (from the surroundings). During condensing and freezing heat is released (to the surroundings). The heat supplied during melting and boiling is used to overcome or 'break' the attractive forces between particles by increasing their kinetic energy. The heat released during **condensing** and **freezing** is derived from the reduction in the average kinetic energy of the particles.

Certain solids, for example frozen carbon dioxide (dry ice), can change directly to a gas without passing through the liquid state. This is known as **sublimation** and the substance is said to sublime. This means molecules leave the solid with enough kinetic energy to exist as gas particles. If the temperature is lowered the gas particles slow down and re-form the solid without passing through the liquid state. This is known as vapour deposition or simply deposition.

Solids that have high melting points have stronger attractive forces (bonds or intermolecular forces) acting between their particles than those with low melting points. Table 1.2 shows a list of some substances (elements and compounds) with their corresponding melting and boiling points.

■ **Table 1.2** Selected melting and boiling points

Substance and formula	Melting point/°C	Boiling point/°C
Aluminium, Al	661	2467
Ethanol, C_2H_5OH	−117	79
Mercury, Hg	−30	357
Oxygen, O_2	−218	−183
Sodium chloride, NaCl	801	1413
Water, H_2O	0	100

If the liquid is heated, the particles (usually molecules) will move around even faster as their average kinetic energy increases. Their kinetic energy constantly changes due to collisions. Some particles at the surface of the liquid have enough kinetic energy to overcome the forces of attraction between themselves and the other particles in the liquid and they escape from the surface to form a gas. This process is known as **evaporation** and takes place at all temperatures below the boiling point. If the temperature is lowered the reverse process, known as **condensation**, occurs. The gas particles move more slowly and enter the surface of the liquid.

Eventually, a temperature is reached (the boiling point) at which the particles are trying to escape from the liquid so quickly that bubbles of gas form inside the bulk of the liquid. This is known as boiling. At the **boiling point** the pressure of the gas created above the liquid equals that in the air (atmospheric pressure) (Chapter 7).

Liquids with high boiling points have stronger forces (bonds or intermolecular forces) operating between their particles than liquids with low boiling points. Chemical bonding and intermolecular forces are discussed in Chapter 4.

When a gas is cooled, the average kinetic energy (speed) of the particles decreases and the particles (usually molecules) move closer and their average separation decreases. The forces of attraction become significant, and if the temperature is lowered to the condensation point the gas will condense to form a liquid. When a liquid is cooled to its freezing point (equal in value to the melting point) it freezes to a solid. During condensing and freezing heat energy is released.

The changes of state are physical changes: no new chemical substances are formed. Ice, water and steam all contain molecules with the formula H_2O. Whenever a change in state occurs the temperature remains constant during the change.

> **2** Identify the change of state which describes the following processes:
>   **a** Solid ethanol changing to liquid ethanol
>   **b** Molten metal solidifying in a mould
>   **c** Water changing to steam at 100 °C
>   **d** Bubbles of ethanol gas forming in liquid ethanol
>   **e** Ice forming from water vapour on the freezer compartment of a fridge
>   **f** Solid aluminium chloride forming a gas on gentle heating
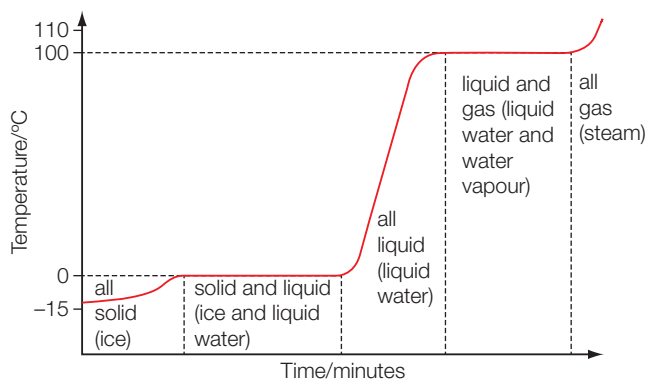
■ **Figure 1.10** Graph of temperature against time for the change from ice at −15 °C to water to steam at 120 °C

Heating and cooling curves

The graph shown in Figure 1.10 was constructed from data-logger measurements by plotting the temperature of water as it was heated steadily from −15 °C to 120 °C (at 1 atmosphere pressure). The heating curve shows that two changes of state have taken place. When the temperature was first measured only ice was present. After a short period of time the curve flattens, showing that even through heat energy is being absorbed, the temperature remains constant. This indicates that a change in state is occurring.

In ice the molecules of water are close together and are attracted to one another by intermolecular forces.

For ice to melt, the molecules must obtain sufficient kinetic energy to overcome the forces of attraction between the water particles to allow relative movement to take place. This is what the heat energy is doing. 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. The addition or presence of impurities lowers the melting point. Figure 1.11 shows a simple apparatus used to find the melting point of a solid. Commercial melting point apparatus uses a heating block to melt the sample.

The purity of substances is very important. Consumers must be certain that foods and medicines do not contain harmful substances. Very small amounts of some chemicals can cause death. The food and drug industries must check constantly to ensure that the substances they use are pure.

To boil a liquid such as water, it has be to given some extra heat energy. This can be seen on the graph (Figure 1.10) where the curve levels out at 100 °C, which is the boiling point of water (at 1 atmosphere pressure). The reverse processes of condensing and freezing occur on cooling. This time, however, heat energy is given out when the gas condenses to the liquid and the liquid freezes to give the solid. Both changes of state occur at constant temperature.

■ **Figure 1.11** Simple apparatus to find the melting point of a solid (in the melting point tube)

3 A solid molecular compound X was heated at constant power for 20 minutes. Its temperature varied as shown in the graph below.

  a Deduce the melting and boiling points of substance X.
  b State the physical state of X at 25, 50 and 100 °C.
  c Explain what is happening during the melting and boiling of X.
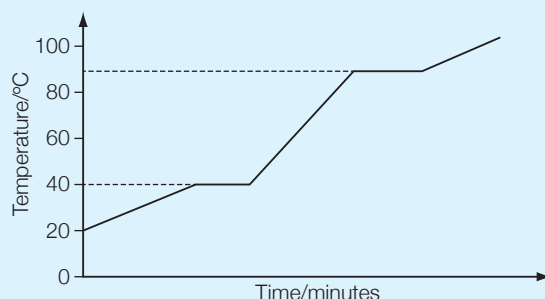
Utilization: Unusual states of matter

Liquid crystals

Liquid crystals (Figure 1.12) are a state of matter that look and flow like liquids (see Chapter 22 on the accompanying website). However, they have some order in the arrangement of their particles (molecules), and so in some ways behave like crystals. Liquid crystals are widely used in displays for digital watches, calculators, lap-top computer displays and in portable televisions. They are also useful in thermometers because certain liquid crystals change colour with temperature changes.

Plasma

A plasma is the superheated gaseous state consisting of a mixture of electrons and highly charged positive ions. It is found at extremely high temperatures in the interiors of stars or in intense electrical fields, such as low pressure discharge tubes (see Chapter 2). Astronomical studies have revealed that 99% of the matter in the Universe is present in the plasma state. Inductively coupled plasma spectroscopy is an important technique for detecting and quantifying small amounts of metals (see Chapter 22 on the accompanying website).

■ **Figure 1.12** A polarized light micrograph of liquid crystals

the coolant uses heat energy
from the air in the cabinet
to vaporize in the coils
around the ice box

pump
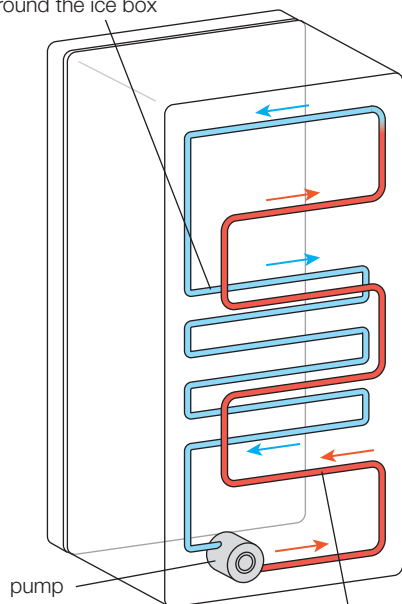
the coolant condenses in these pipes, giving
out thermal energy which heats the air

■ **Figure 1.13** The coolant system of a refrigerator

■ **Figure 1.14** A domestic air conditioner

■ **Figure 1.15** An astronaut eating a freeze-dried
meal on board Space Shuttle Discovery

Utilization of heat changes during changes of state: Refrigeration

It is difficult to over-estimate the importance of the invention of the modern refrigerator in the context of food transportation and storage. The invention of refrigerated transport for food led to a revolution in the globalization of markets and the availability of important commodities across, and between, continents.

A refrigerator takes advantage of the heat energy transfers when a volatile (low boiling point) liquid evaporates and condenses. The key stage of the system depends on the fact that evaporation absorbs heat from the surroundings. Within the body of the refrigerator (Figure 1.13) a pump circulates a low boiling point liquid around a circuit of pipes. This volatile liquid vaporizes in the pipes inside the refrigerator, taking in heat energy from the air inside the refrigerator and keeping the food and drinks inside cold.

Continuing around the circuit, the vapour (gas) is compressed by the pump as it flows out at the bottom of the refrigerator. The compressed vapour is hot. As it flows through the pipes at the back of the refrigerator the vapour cools and condenses back to a liquid, releasing heat energy and heating up the air around the back of the cabinet. Overall, heat energy is transferred from inside the refrigerator to the air in the room.

The use of the reversible evaporation–condensation cycle of volatile liquids in refrigeration and air conditioning (Figure 1.14) is one of the features of modern living. In the past, many air conditioners commonly used CFCs (chlorofluorocarbons) as their volatile liquid. However, in most countries the manufacture and use of CFCs is either banned or restricted. This is because when CFC molecules reach the upper atmosphere ultraviolet radiation from the Sun breaks the carbon–chlorine bond, yielding a chlorine atom. These chlorine atoms catalyse the breakdown of ozone (trioxygen) into dioxygen, depleting the ozone layer that protects the Earth's surface from strong ultraviolet radiation.

Utilization of removal of water at low pressure: Freeze-drying

The basic idea of freeze-drying is to completely remove water from food while leaving the basic structure and composition unchanged. Removing water keeps food from spoiling for a long period of time. Food spoils when bacteria digest the food and decompose it. Bacteria may release toxins that cause disease, or they may just release chemicals that make food taste bad. Additionally, naturally occurring enzymes in food can react with oxygen to cause spoiling and ripening. Bacteria need water to survive, so if water is removed from food it will not spoil. Enzymes also need to be hydrated to react with food, so dehydrating food will also stop spoiling.

Freeze-drying also significantly reduces the total mass of the food. Most food is largely made up of water and removing this water makes the food a lot lighter, which means it is easier to transport. The military and camping supply companies freeze-dry foods to make them easier for one person to carry. NASA has also freeze-dried foods for the cramped quarters on board spacecraft and the International Space Station (Figure 1.15).
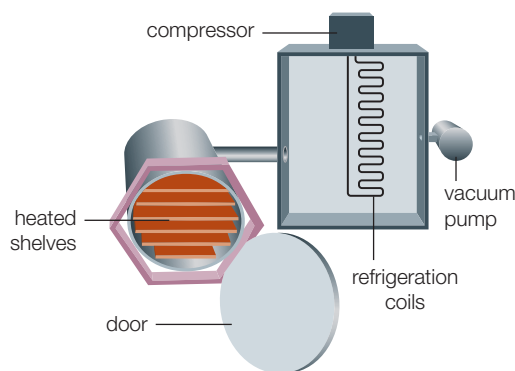
■ **Figure 1.16** Freeze-drying machine

Freeze-drying is not normally carried out by simple evaporation. It is difficult to remove water completely using evaporation because most of the water is not directly exposed to air. Unless all the water is removed then there will be some bacterial and enzyme activity. In addition, the heat involved in the evaporation process changes the shape, texture and composition of the food.

The fundamental principle in freeze-drying is sublimation, the phase change from a solid directly into a gas (at constant temperature). A lowering of the pressure (below 0.6 atmospheres) and an increase in temperature results in water being converted to a gas, rather than liquid water.

A typical freeze-drying machine (Figure 1.16) consists of a freeze-drying chamber with several shelves attached to heating units, a freezing coil connected to a refrigerator compressor, and a vacuum pump.

The machine runs the compressors to lower the temperature in the chamber. The food is frozen solid, which separates the water from everything around it, on a molecular level, even though the water is still present.

The heating units supply a small amount of heat to the shelves, causing the ice to change phase. Since the pressure is so low, the ice turns directly into water vapour. The water vapour flows out of the freeze-drying chamber, past the freezing coil. The water vapour condenses onto the freezing coil in solid ice form, in the same way as water condenses as frost on a cold day.

Utilization: Atom economy in chemical reactions

The atom economy examines the theoretical potential of a reaction, by considering the quantity of starting atoms in all the reactants that end up in the desired product.

$$\% \text{ atom economy} = \frac{\text{atomic mass of all utilized atoms}}{\text{atomic mass of all reactants}}$$

Taking the laboratory preparation of copper(II) sulfate from copper(II) oxide and sulfuric acid as an example, this is an acid–base reaction. This calculation is explained later in the chapter and involves calculating the mass of one mole of each substance from the formula and the relative atomic masses of the elements.

$$CuO(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l)$$
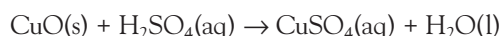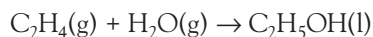
Mass of starting atoms is
$CuO = 63.5 + 16 = 79.5\,g$
$H_2SO_4 = 2 + 32 + 64 = 98\,g$
Total $= 177.5\,g$

Mass of desired product is
$CuSO_4 = 63.5 + 32 + 64 = 159.5\,g$

$$\% \text{ atom economy} = \frac{159.5}{177.5} \times 100 = 89.9\%$$

The atom economy for the production of ethanol from ethene and steam is shown below. This is known as an addition or hydration reaction.

$$C_2H_4(g) + H_2O(g) \rightarrow C_2H_5OH(l)$$

$$\% \text{ atom economy} = \frac{46}{46} \times 100 = 100\% \text{ (there are no unwanted products)}$$

A higher atom economy means that there is a higher utilization of the atoms of reactants into the final useful products. That is, there is a better use of materials and also less waste formation.

4 Calculate the atom economy for the reaction between carbon and steam to form carbon dioxide and hydrogen.

Green chemistry

Green chemistry consists of chemicals and chemical processes designed to reduce or eliminate impacts on the environment. The use and production of these chemicals may involve reduced waste products, non-toxic chemicals, and improved efficiency. Industrial chemists evaluate chemical pathways and their economic and environmental costs by calculating the relative efficiency of the chemical reactions involved. Percentage yield provides a means of comparison of the theoretical and actual quantity of product, and used to be the main way of evaluating reaction efficiency. However, calculation of 'atom economy' has become a more important means

of comparing the efficiency of chemical reactions. Atom economy is a measure of the proportion of reactant atoms that is incorporated into the desired product of a chemical reaction. Calculation of atom economy therefore also gives an indication of the proportion of reactant atoms forming waste products.

■ Mixtures

In a mixture of two elements there are two types of atoms present, but they are not chemically bonded to each other. Figure 1.17 shows a mixture of elements existing as atoms and a mixture of two elements existing as diatomic molecules.
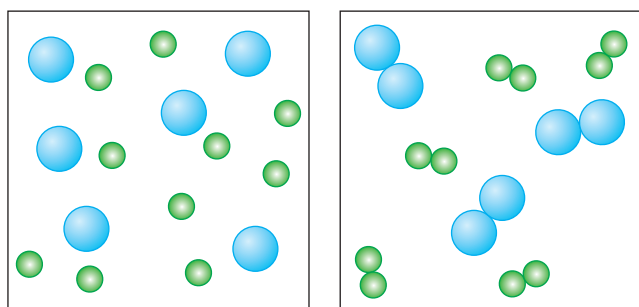
a mixture of two elements existing as atoms

b mixture of two elements existing as molecules

■ **Figure 1.17** Particle representations of a mixture of atoms and a mixture of molecules

A compound always contains the same proportion (by mass) of each element. For example, iron(II) sulfide has iron and sulfur in the ratio of 55.85 to 32.06, i.e. 1.742 to 1.000. However, a mixture can have any proportion of each element. For example, the percentage (by mass) of sulfur in an iron–sulfur mixture can range from close to 0% to almost 100%.

Alloys are mixtures of metals and other elements (often carbon) that have been melted together and then allowed to solidify. Common alloys include brass (a mixture of copper and zinc) and bronze (copper and tin).

The major differences between mixtures of elements and compounds are summarized below in Table 1.3.

■ **Table 1.3**
The major differences between mixtures of elements and compounds

Mixture	Compound
It contains two or more substances (elements or compounds)	It is a single pure substance
The composition is variable	The composition (by mass) is fixed
No chemical reaction takes place when a mixture is formed	A chemical reaction occurs when a compound is formed
The properties are those of the individual elements or compounds	The properties are very different to those of the component elements

5 State whether each of the boxes below contains an element, a compound or a mixture.

a

b

c

d

e

f
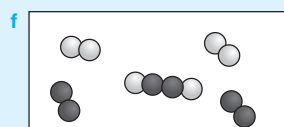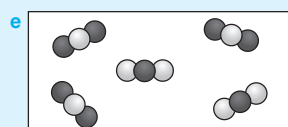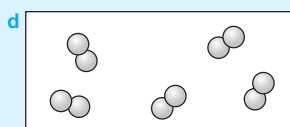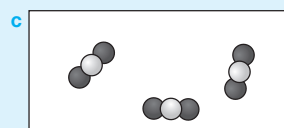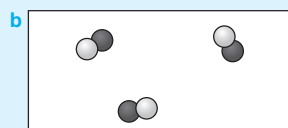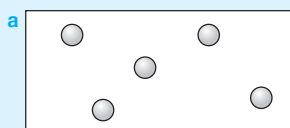
■ Types of mixtures

There are many different types of mixtures. One classification of mixtures is to classify them as **homogeneous** or **heterogeneous**. For example, if gaseous bromine is introduced into a gas jar filled with air (mainly nitrogen) it will diffuse and spread evenly through both gas jars (Figure 1.18). The concentrations of bromine and nitrogen will be the same throughout both gas jars. Mixtures of gases are described as being homogeneous since they have a uniform or constant composition. Figure 1.19 shows how kinetic molecular theory can be used to explain diffusion in gases. Gases diffuse quickly because the particles are moving rapidly and there are large spaces between the molecules.