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Chemistry

for the IB Diploma

MULTI-COMPONENT SAMPLE

Third edition

Digital Access

Brighter Thinking
Better Learning

Dear Teacher,

Welcome to the new edition of our *Chemistry for the IB Diploma* series, providing full support for the new course for examination from 2025. This new series has been designed to flexibly meet all of your teaching needs, including extra support for the new assessment. This preview will help you understand how the coursebook, the workbook and the teacher's resource work together to best meet the needs of your classroom, timetable and students.

This Executive Preview contains sample content from the series, including:

- A guide explaining how to use the series
- A guide explaining how to use each resource

In developing this new edition, we carried out extensive global research with IB Chemistry teachers – through lesson observations, interviews and work on the Cambridge Panel, our online teacher research community. Teachers just like you have helped our experienced authors shape these new resources, ensuring that they meet the real teaching needs of the IB Chemistry classroom.

The coursebook has been specifically written to support English as a second language learners with key subject words, glossary definitions in context and accessible language throughout. We have also provided new features that help with active learning, assessment for learning and student reflection. Numerous exam-style questions with answers in the digital coursebook, which accompanies the print coursebook, ensure your students feel confident approaching the assessment and have all the tools they need to succeed in their examination.

Core to the series is the brand-new digital teacher's resource. It will help you support your learners and confidently teach to the new IB Chemistry guide, whether you are new to teaching the subject or more experienced. For each topic there are lesson ideas and activities, common misconceptions to look out for, worksheets, PowerPoint presentations, answers to the coursebook, extra wrap-up activities and more. Also included is a practical guide to help your students develop their academic writing.

Please take five minutes to find out how our resources will support you and your learners. To view the full series, you can visit our website or speak to your local sales representative. You can find their contact details here:

cambridge.org/gb/education/find-your-sales-consultant

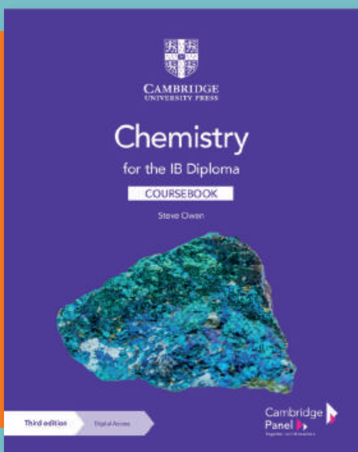
Best wishes,

Micaela Inderst
Senior Commissioning Editor for the IB Diploma

Cambridge University Press

How to use this series

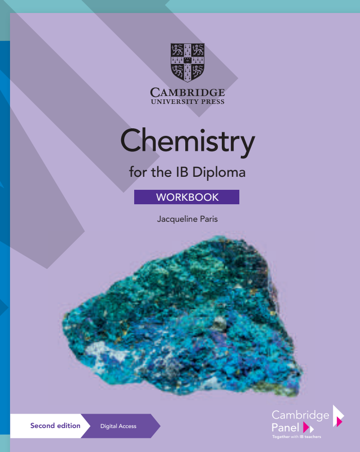
This suite of resources supports students and teachers of the Chemistry course for the IB Diploma programme. All of the books in the series work together to help students develop the necessary knowledge and scientific skills required for this subject.

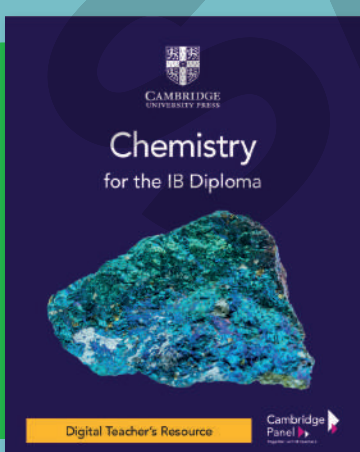


The coursebook with digital access provides full coverage of the latest IB Chemistry guide.

It clearly explains facts, concepts and practical techniques, and uses real world examples of scientific principles. A wealth of formative questions within each chapter help students develop their understanding, and own their learning. A dedicated chapter in the digital coursebook helps teachers and students unpack the new assessment, while exam-style questions provide essential practice and self-assessment. Answers are provided on Cambridge GO so support self-study and home-schooling.

The workbook builds upon the coursebook with digital access with further exercises and exam-style questions, carefully constructed to help learners develop the skills that they need as they progress through their IB Chemistry Diploma course. The exercises also help students develop understanding of the meaning of various command words used in questions, and provide practice in responding appropriately to these.





The teacher's resource supports and enhances the coursebook with digital access and the workbook. This resource includes teaching plans, overview of required background knowledge, learning objectives and success criteria, common misconceptions, and a wealth of ideas to support lesson planning, assessment and differentiation. detailed lesson ideas. It also includes editable worksheets for vocabulary support and exam practice (with answers) and exemplar PowerPoint presentations, to help you plan and deliver your best teaching.

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Chemistry

for the IB Diploma

COURSEBOOK

Steve Owen

Third edition

Digital Access

Cambridge Panel
Together with IB teachers

Contents

How to use this book

Throughout this book, you will find lots of different features that will help your learning. These are explained below.

UNIT INTRODUCTION

A unit is made up of a number of chapters. The key concepts for all the chapters covered in a unit are summarised in the Unit opening chapter as the introduction.

LEARNING OBJECTIVES

Each chapter in the book begins with a list of learning objectives. These set the scene for each chapter, help with navigation through the coursebook and indicate the important concepts in each topic. A bulleted list at the beginning of each section clearly shows the learning objectives for the section.

GUIDING QUESTIONS

These are questions on subject knowledge you will need before starting each chapter.

Link

These are a mix of questions and explanation that refer to other Chapters or sections of the book.

The content in this book is divided into Standard and Higher Level material. Either a chevron or a vertical line running down the margin of all Higher Level material, allows you to easily identify Higher Level from Standard material.

Key terms are highlighted in **orange bold** font at their first appearance in the book so you can immediately recognise them. At the end of the book, there is a glossary that defines all the key terms.

KEY POINTS

This feature contains important key learning points (facts) and/or equations to reinforce your understanding and engagement.

EXAM TIPS

These short hints provide useful information that will help tackle the tasks in the exam.

SCIENCE IN CONTEXT

This feature presents real-world examples and applications of the content in a chapter, encouraging you to look further into topics. You will note that some of these features end with questions intended to stimulate further thinking, prompting you to look at some of the benefits and problems of these applications.

NATURE OF SCIENCE

Nature of Science is an overarching theme of the IB Chemistry Diploma course The theme examines the processes and concepts that are central to scientific endeavour, and how science serves and connects with the wider community. Throughout the book, there are 'Nature of Science' paragraphs that discuss particular concepts or discoveries from the point of view of one or more aspects of Nature of Science.

THEORY OF KNOWLEDGE

This section stimulates thought about critical thinking and how we can say we know what we claim to know. You will note that some of these features end with questions intended to get you thinking and discussing these important Theory of Knowledge issues.

INTERNATIONAL MINDEDNESS

Throughout this Chemistry for the IB Diploma course, the international mindedness feature highlights international concerns. Chemistry is a truly international endeavour, being practised across all continents, frequently in international or even global partnerships. Many problems that chemistry aims to solve are international and will require globally implemented solutions.

EXTENSION

This feature highlights information in the book that is extension content and is not part of the syllabus.

TEST YOUR UNDERSTANDING

These questions appear within each chapter, to help you develop your understanding. The questions can be used as the basis for class discussions or homework assignments. If you can answer these questions, it means you have understood the important points of a section.

WORKED EXAMPLE

Many worked examples appear throughout the text to help you understand how to tackle different types of questions.

REFLECTION

The questions appear at the end of each chapter. The purpose is for you as a learner to reflect on the development of your skills proficiency and your progress against the objectives. The reflection questions are intended to encourage your critical thinking and inquiry-based learning.

EXAM-STYLE QUESTIONS

Exam-style questions at the end of each topic provide essential practice and self-assessment. These are signposted in the print coursebook and can be found in the digital version of the coursebook.

SELF-ASSESSMENT CHECKLIST

These appear at the end of each Chapter as a series of statements. You might find it helpful to rate how confident you are for each of these statements when you are revising. You should re-visit any topics that you rated 'Needs more work' or 'Almost there'.

I can	Section	Needs more work	Almost there	Confident to move on

Free online material

Additional material to support the Chemistry for the IB Diploma course is available online.

This includes Assessment guidance – a dedicated chapter in the digital coursebook helps teachers and students unpack the new assessment and model exam specimen papers. Additionally, answers to the Test your understanding and Exam-style questions are also available.

Visit Cambridge GO and register to access these resources at www.cambridge.org/GO.

Unit 1

The nature of matter

INTRODUCTION

We have all heard of atoms: the particles from which everything is made. Democritus and his teacher, Leucippus, fifth-century BCE Greek philosophers, are usually credited with first suggesting the idea of the atom as the smallest indivisible particle from which all matter is made, but the modern understanding of science in terms of atoms only really began in the 19th century with the work of John Dalton (1766–1844). An understanding of atoms and atomic structure is now regarded as fundamental to chemistry, but we usually talk about atomic *theory*, so does that mean that atoms may not really exist? We will not look specifically at the evidence for the existence of atoms, but does the fact that everything in this book and other scientific literature is explained by assuming the existence of atoms provide that evidence?

So, assuming that atomic theory is the best way of explaining the world around us, what do we know about atoms? Atoms are most definitely small – there are many more atoms in a drop of water than there are stars in the Milky Way, and there are probably more atoms in a glass of water than there are stars in the universe (although no one is sure how many stars there are in the universe). We know that there are different types of atoms, but how many are there? A simple answer would be as many as there are elements, but there are also isotopes, and which isotope we are talking about can make a big difference to the properties of the element and to the world – a country with a storage vault full of uranium-235 (which can be used for making nuclear weapons) will be viewed very differently by other governments from one with uranium-238!

Chapter 1

The particulate nature of matter

LEARNING OBJECTIVES

In this chapter you will:

- understand the terms element, compound and mixture
- understand the differences between heterogeneous and homogeneous mixtures
- understand how to separate the components of a mixture
- use kinetic molecular theory to understand the properties of solids, liquids and gases
- understand that temperature in K is proportional to the average kinetic energy of particles
- understand how to convert temperatures between K and °C
- use state symbols in chemical equations
- use kinetic molecular theory to explain changes of state.

GUIDING QUESTIONS

- What are the differences between elements, compounds and mixtures?
- How can the components of a mixture be separated?
- How can kinetic molecular theory be used to explain the properties of solids, liquids and gases?

Introduction

The song 'Woodstock', released in 1970, includes the words 'we are stardust' and, strangely enough, this is pretty much true. The lightest **elements** (mostly hydrogen and helium with some lithium) were formed in the immediate aftermath of the Big Bang, but the other elements that we, and everything around us, are made of were formed in stars. In this chapter, we will look at the distinction between elements, **compounds** and **mixtures**, explain their properties in terms of kinetic molecular theory and look at how to separate the components of mixtures. The distinction between elements, compounds and mixtures is fundamental to an understanding of chemistry and, although in subsequent chapters we will mention very little about mixtures, it is important to remember that most substances in everyday life are actually mixtures.

1.1 Elements, compounds and mixtures

Elements

Elements are the primary constituents of matter. There are 118 elements that have been discovered so far, and these are shown in the periodic table. Of these, about 90 occur naturally in reasonable amounts, and the rest are present in only trace amounts or are artificially made. By far the most abundant element in the universe is hydrogen, followed by helium, but in the Earth's crust oxygen is the most abundant and astatine is the least abundant. Astatine has no stable isotopes and scientists estimate that, at any one time, there is probably less than 30 g present in the whole of the Earth's crust.

An element can be defined in different ways and, for the moment, we will define it in terms of its properties:

KEY POINT

An **element** is a chemical substance that cannot be broken down into a simpler substance by chemical means.

Gold only contains gold **atoms** and sulfur only contains sulfur atoms, and because of this, these cannot be broken down into anything simpler than gold atoms or sulfur atoms using chemical reactions.

In Chapter 2, we will look at the structure of atoms, and this will allow us to define an element in terms of the particles that make up the atom:

KEY POINT

An element is a pure substance in which each atom has the same number of protons in the nucleus (see Chapter 2).

So, for example, gold is an element and all samples of pure gold contain only atoms that have 79 protons in the nucleus.

The symbols for elements are shown in the periodic table (Chapter 10, section 10.1). In a sample of an element, the atoms may be present as individual atoms (e.g. helium, He), be chemically bonded as individual **molecules** (e.g. oxygen, O_2, or ozone, O_3) or be chemically bonded as part of a giant structure (e.g. gold, Au, or carbon, C). Some representations of elements are shown in Figure 1.1. The key thing to notice is that, in each part of Figure 1.1, all the atoms are the same.

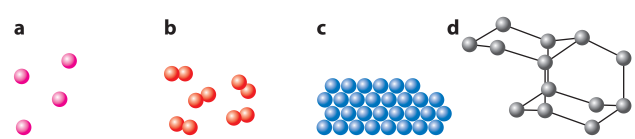


Figure 1.1: Some elements. **a** This could be a noble gas, such as helium, which consists of just individual atoms. **b** This could be gaseous oxygen, consisting of O_2 molecules, in which the oxygen atoms are chemically bonded to each other. **c** This could be a metal, such as gold. **d** This could be carbon – the lines represent bonds between atoms.

THEORY OF KNOWLEDGE

What is an element?

The concept of an element is fundamental to the study of chemistry, but, strangely enough, chemists do not necessarily agree on the definition of an element. If we say that oxygen is an element, that is fine, but do we mean an O atom, a sample of oxygen gas, which contains O_2 molecules, or even ozone, which contains O_3 molecules?

Chemistry is partly a study of how chemical elements combine to make the world and the universe around us. When different elements combine chemically, they form compounds.

Compounds

In water (H_2O), a compound, there are always exactly twice as many hydrogen atoms as oxygen atoms – this ratio never varies for a particular compound. If the ratio is different, it is a different compound, for example, if the ratio is 1:1, the compound is hydrogen peroxide (H_2O_2) and not water.

KEY POINT

A **compound** is a pure substance formed when two or more elements combine chemically in a fixed ratio.

The atoms (or ions) in a compound are chemically bonded to each other – this may be covalent bonding (see Chapter 7, section 7.1) or ionic bonding (see Chapter 6, section 6.1). Some representations of the structures of compounds are shown in Figure 1.2. Sometimes the chemical bonds (lines) will be shown (two of the structures of water) and sometimes not.

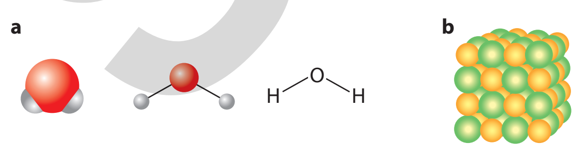


Figure 1.2: Some representations of compounds. The key thing to notice here is that more than one type of atom is present in each structure. **a** Three different ways of showing the structure of water. **b** An ionic compound, such as sodium chloride.

KEY POINT

The elements in a compound are chemically combined, and therefore, compounds can only be converted into their elements again by chemical reactions.

For example, hydrogen can be obtained from water by reacting it with sodium, or hydrogen and oxygen could both be produced by electrolysis (passing electricity through water).

The **physical properties** and **chemical properties** of a compound are different from those of the elements from which it is formed.

KEY POINT

Chemical properties how a substance behaves in a chemical reaction.

Physical properties all the other properties of a substance, such as melting point, density, hardness and electrical conductivity.

For example, hydrogen an explosive gas, combines with oxygen, a highly reactive gas, to form water, which is a liquid at room **temperature**. Water reacts in very different ways to hydrogen and oxygen (it has different chemical properties) – you would not try to put a fire out with hydrogen or oxygen!

Similarly, when sodium (a highly reactive metal) is heated with chlorine (a toxic gas), a white, crystalline substance, sodium chloride (common salt), is formed, which reacts in very different ways to sodium and chlorine.

Mixtures

Elements and compounds are pure substances, but most things around us are not pure, they are mixtures. We breathe in air, which is a mixture; all the foods we eat are mixtures; oxygen is carried around our body by blood, another mixture.

The components of a mixture can be elements or compounds – or even mixtures! Air is a mixture of mostly elements (nitrogen, oxygen, argon) with smaller amounts of compounds (carbon dioxide, water vapour etc.). Representations of mixtures are shown in Figure 1.3.

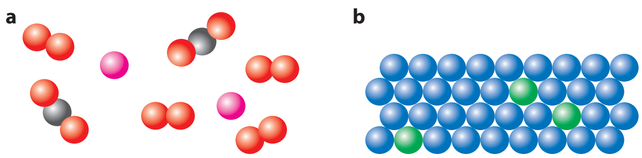


Figure 1.3: Some mixtures. **a** A mixture of gases. **b** An alloy.

> **KEY POINT**
>
> The components of a mixture are not chemically bonded together, so they retain their individual properties.

In a mixture of iron and sulfur, the two elements retain their individual chemical and physical properties, so iron is magnetic and will react with dilute acids to form hydrogen gas, and sulfur is yellow and burns in air to form sulfur dioxide. When the mixture is heated and forms the compound iron sulfide, this has a different appearance, is not magnetic (Figure 1.4) and, for example, reacts with acids to form the extremely smelly and toxic gas hydrogen sulfide – the compound has different properties to its elements.

Figure 1.4: The iron in a mixture of iron and sulfur (left) retains its magnetic properties, but iron sulfide (right) is not magnetic.

> **KEY POINT**
>
> The components of a mixture can be mixed together in any proportion.

When atoms combine to form compounds, they do so in fixed ratios (according to the formula of the compound), but there are no such limitations on making a mixture, and iron and sulfur can be mixed together in absolutely any proportions. **Solutions** are mixtures, and a solution of sodium chloride could be made by dissolving 1 g of sodium chloride in 100 cm^3 of water, 2 g of sodium chloride in 100 cm^3 water, 10 g of sodium chloride in 100 cm^3 of water etc., up to the limit of solubility (how much dissolves at a certain temperature).

Link

Alloys are mixtures of metals with other metals (or non-metals). In alloys, there is metallic bonding throughout the structure (see Chapter 8, section 8.1), so the components of the mixture are actually chemically bonded to each other. An alloy is, however, still regarded as a mixture because it will not have a fixed composition – the metals can be mixed together in various proportions.

Mixtures can be homogeneous or heterogeneous

> **KEY POINT**
>
> A **homogeneous mixture** has the same (uniform) composition throughout the mixture – it consists of only one phase.
>
> Solutions and mixtures of gases are homogeneous mixtures.
>
> A **heterogeneous mixture** does not have uniform composition – it consists of separate phases.

The term *phase* can be used in different ways in chemistry; here, it refers to a region that is the same throughout, in terms of chemical composition and physical properties. In a **heterogeneous mixture**, there will be distinct boundaries between different phases.

An example of a **homogeneous mixture** is a solution. The concentration is the same throughout: if several 1cm^3 samples of a solution of sodium chloride are taken from a beaker and evaporated separately to dryness, the same mass of solid sodium chloride will be recovered from each sample.

One example of a heterogeneous mixture is sand in a sample of water. Sand and water can be distinguished from each other – they are separate phases. Other examples include milk and orange juice. Orange juice is a complex mixture, containing an aqueous phase with various substances dissolved or suspended in it. Suspended material in orange juice includes cellulose, proteins, lipids and pectins. If you leave some freshly squeezed orange juice to stand, some parts will settle out, but will it become completely clear? Milk, as a colloid, is discussed in detail in the Science in Context section below.

Mixtures of solids are always heterogeneous mixtures. For example, a mixture of iron and sulfur is a heterogeneous mixture. Even though the mixture may have been made very carefully, so that there are the same masses of iron and sulfur in each cubic centimetre, the composition is not uniform because there are distinct particles of iron and sulfur (you may need to use a magnifying glass to see them), and each particle of iron and sulfur represents a different phase.

When looking at mixtures that are liquids or gases, if the mixture is clear, so that you can see through it, it is a homogeneous mixture; if it is cloudy/opaque, so that some/all of the light is scattered as it passes through it, then the mixture is heterogeneous.

SCIENCE IN CONTEXT

Solutions and mixtures

Tea or coffee without milk are solutions. This is usually easier to see with tea, but, if you dilute your black coffee in a glass cup with some water, you will be able to see that, although it is coloured, it is clear, so that light passes through it without being scattered, and therefore, it is a solution and a homogeneous mixture (although, if you used a cafetiere or a not-very-good filter, you may still have a few coffee grounds in it, which would make it a heterogeneous mixture!). If you add sugar and stir it well, the sugar dissolves, and so, you still have a homogeneous mixture; however, if you add milk (Figure 1.5), your coffee goes cloudy – this is now a heterogeneous mixture. Milk is a type of mixture called a colloid (or colloidal system) and contains very small droplets of fat and solid protein particles dispersed throughout an aqueous phase. These particles scatter light (the Tyndall effect) and, therefore, white coffee is not clear but opaque.

Figure 1.5: White coffee and doughnuts are heterogeneous mixtures.

To consider:

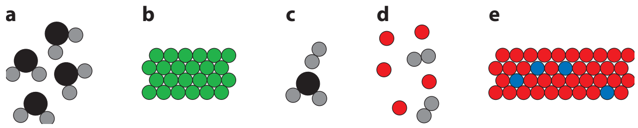
1. Other heterogeneous mixtures you will come across in a coffee shop include whipped cream, hot chocolate and muffins… can you think of any more?
2. How do the methods for separating homogeneous mixtures differ from methods for separating heterogeneous mixtures?
3. Is it possible to separate all the components from white coffee or doughnuts?

TEST YOUR UNDERSTANDING

1. Classify each of the following as an element, a compound or a mixture:
   - **a** water
   - **b** oxygen
   - **c** potassium iodide
   - **d** vanadium
   - **e** ammonia
   - **f** air
   - **g** hydrogen chloride
   - **h** magnesium oxide.

CONTINUED

2 Classify each of the diagrams shown as an element, compound or mixture:



3 Classify each of the following as a heterogeneous or a homogeneous mixture:

a a mixture of carbon dioxide gas and helium gas

b a mixture of solid copper(II) oxide and solid calcium carbonate

c potassium hydroxide solution

d mayonnaise.

Separating the components of a mixture

The components of a mixture can be separated from each other by physical processes – physical processes are things like filtration and distillation, which do not involve chemical reactions.

Filtration

In a chemistry laboratory, filtration is usually used to separate an insoluble solid from a liquid. It can also be used to separate a solid from a gas. The apparatus most often used for filtration is shown in Figure 1.6. The solid left in the filter paper is called the residue, and the liquid that passes through the filter paper is called the filtrate. The filter paper acts as a physical barrier to the pieces of solid but allows the liquid to pass through gaps between fibres.

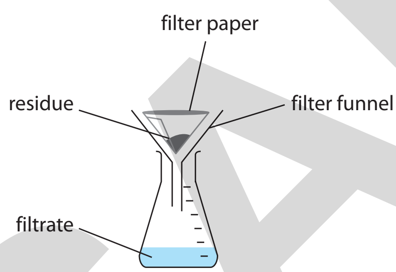


Figure 1.6: Filtration can be used to separate a solid from a liquid.

Filtration is used as part of the process in the preparation of copper(II) sulfate. Copper(II) sulfate solution can be made by the reaction between copper(II) oxide (a black solid that is insoluble in water) and dilute sulfuric acid. Excess (more than enough to react with all the sulfuric acid) copper(II) oxide is added to hot sulfuric acid. The excess copper(II) oxide is then filtered off. In this case, copper(II) oxide is the residue and copper(II) sulfate solution is the filtrate.

The equation for the reaction is

$$CuO(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l)$$

Link

Copper(II) oxide is a base and reacts with sulfuric acid in a neutralisation reaction (Chapter 19).

INTERNATIONAL MINDEDNESS

Diesel engines

Diesel engines are used extensively in heavy-duty commercial vehicles, such as lorries and buses, as well as cars. One of the major environmental concerns with the use of vehicles with diesel engines is that they emit significantly more particulate matter (soot) than gasoline (petrol) engines, and this can be damaging to health. Diesel engines are, therefore, fitted with particulate filters, to filter out as much of the particulate matter as possible. Different countries have different regulations on emissions from diesel vehicles.

Evaporation

Evaporation can be used to remove a **solvent** from a solution to leave the **solute**. If a solution of sodium chloride is heated, water will evaporate/boil off to leave solid sodium chloride (Figure 1.7).

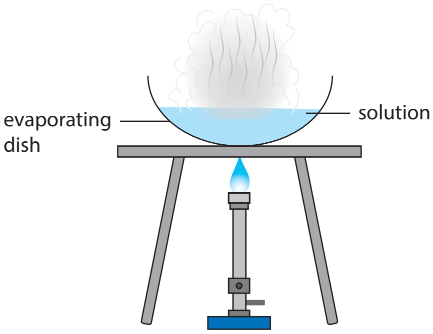


Figure 1.7: Evaporation of the solvent can be used to obtain a solute from a solution. If larger crystals are required, only some of the water should be boiled off and then the solution should be left to crystallise.

Solvation

Solvation can be used to separate a mixture of two or more substances, due to differences in solubility.

For example, a mixture of solid copper(II) oxide (insoluble in water) and sodium chloride (soluble in water) can be separated by putting the mixture into a beaker of warm distilled/deionised water and stirring to make sure that all the sodium chloride has dissolved. The mixture is filtered: copper(II) oxide is the residue and sodium chloride solution is the filtrate. Copper(II) oxide is washed with distilled water to remove any traces of sodium chloride solution and then dried in a warm oven (distilled water will evaporate). Solid sodium chloride can be obtained from the solution by heating it in an evaporating dish until all the water evaporates.

Note that distilled/deionised water must be used because tap water contains dissolved solids and, when heated, will leave a residue of these solids, so that the copper(II) oxide and sodium chloride obtained will not be pure.

Application of solvation to the extraction of caffeine

A common laboratory experiment is the extraction of caffeine from tea. The basic principles of the technique are that tea leaves are boiled with water to make an aqueous solution, which is shaken with dichloromethane (an organic liquid with the formula CH_2Cl_2) in a separatory funnel (Figure 1.8). Dichloromethane is not soluble in water and remains as a separate phase in the separatory funnel. Caffeine is more soluble in dichloromethane than in water and distributes itself between the water layer and the dichloromethane layer, with much more in the dichloromethane layer. The dichloromethane layer can then be run off and the solvent evaporated to leave caffeine (a white solid).

This technique is called solvent extraction, and we talk about a solute partitioning itself between two solvents.

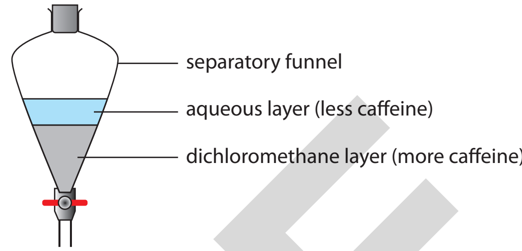


Figure 1.8: A separatory funnel is used in the extraction of caffeine from tea.

Link

Caffeine is more soluble in dichloromethane than in water. A general rule for solubility is 'like dissolves like'. The intermolecular forces are more similar between caffeine (dipole–dipole interactions) and dichloromethane (dipole–dipole interactions) than between caffeine (dipole–dipole interactions) and water (hydrogen bonds). Intermolecular forces will be discussed in Chapter 7.

EXAM TIP

The word solvation is used here because it is used on the IB syllabus, but it is not actually the correct word. Solvation will be discussed further in Chapter 7. The process here is probably best described as dissolving.

Distillation

Distillation could be used, for example, to separate water from a sodium chloride solution. The sodium chloride solution is heated, water evaporates and condenses again in the condenser, so that it can be collected in the collection vessel (it is called the distillate).

Suitable apparatus for distillation is shown in Figure 1.9.

If heating is continued for a long enough time, only solid sodium chloride will be left in the round-bottomed flask, and all the water will be in the collection vessel.

The difference between distillation and evaporation is that, in distillation, the solvent is boiled off, but then condensed again, so that it can be collected. So, evaporation would generally be used when it is the solute that is the desired product and distillation when it is the solvent that is required.

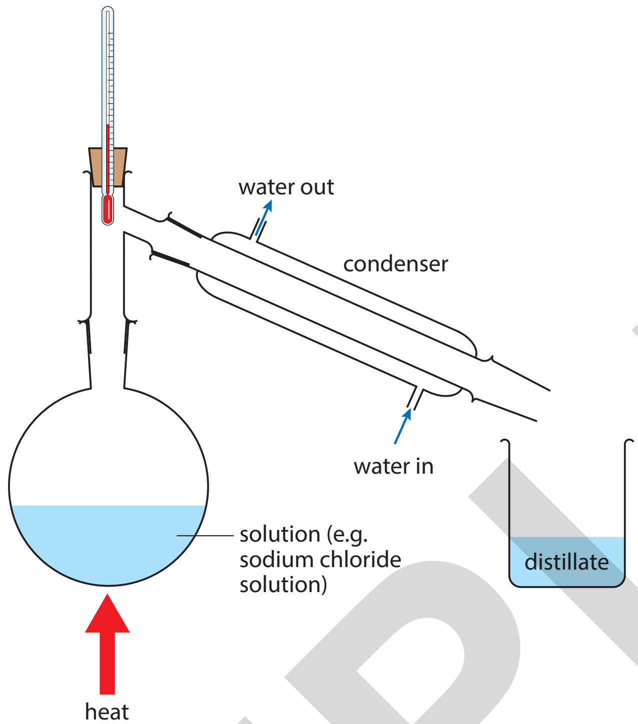


Figure 1.9: The experimental set-up for distillation.

KEY POINT

Distillation (simple distillation) can be used to separate the solute and solvent from a solution (where the solute was a solid) or to separate a mixture of two liquids with sufficiently different boiling points.

INTERNATIONAL MINDEDNESS

Fresh water

Seawater is a mixture, and distillation can be used to obtain water without salt from seawater. The process of removing salt from seawater is called desalination. Desalination is very important in some parts of the world, where sufficient freshwater is not available, 'for example, in parts of Southwest Asia and North Africa. Water obtained by desalination can be used for human consumption, agriculture or in industry.

Distillation can also be used to separate a mixture of two liquids, as long as there is a large enough difference between their **boiling points** (about 70 °C).

The liquid with the lower boiling point (more volatile) will go into the vapour phase more easily and will be collected in the collection vessel (Figure 1.9), whereas the liquid with the higher boiling point will be left in the round-bottomed flask. If the boiling points of the two liquids are too close, then complete separation will not be obtained, and a mixture will distil over.

This technique is used extensively in organic chemistry for extracting the more volatile liquid product of a reaction from the reaction mixture and for purifying a liquid product of a reaction. When used for purification, the pure liquid is collected in the collection vessel, and any non-volatile impurities will be left in the round-bottomed flask. The purity of the liquid could be tested by using chromatography (see the Paper chromatography section below).

Paper chromatography

Paper chromatography may be used, for example, to separate the various dyes in coloured inks, to separate a mixture of sugars or amino acids or to test the purity of a substance. The experimental set-up for paper chromatography is shown in Figure 1.10.

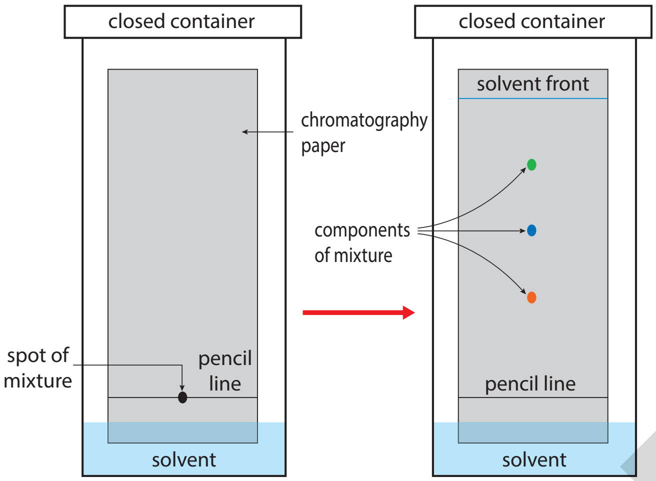


Figure 1.10: A paper chromatography experiment. The process of the solvent travelling up the paper to produce a chromatogram is called development.

To carry out a paper chromatography experiment:

- A line is drawn with a pencil (not a pen, as the inks may move with the solvent) across a piece of chromatography paper about 1 cm from the bottom.
- A sample of the mixture is placed on the pencil line and allowed to dry.
- The paper is suspended in a container with a small amount of solvent at the bottom, so that the end of the paper dips into the solvent (the original sample spot must be above the top of the solvent; otherwise it will just dissolve into the solvent).
- The container is closed, so that the atmosphere becomes saturated with the solvent – this prevents evaporation of the solvent from the surface of the paper.
- The solvent is drawn up the paper by capillary action.
- The process is stopped when the solvent front is about 1 cm from the top of the paper. A pencil line is drawn to record the position of the solvent front and the paper is dried.

At the simplest level, the number of spots present on a chromatogram indicates the number of components of the mixture (although other tests might need to be done, to check whether a particular spot is indeed a pure substance).

KEY POINT

Chromatography can be used to test the purity of the product of a reaction. The presence of more than one spot indicates that the substance is impure.

It can take quite a bit of research and trial and error to find a suitable solvent for chromatography that provides good separation of the components of the mixture. The polarity (see Chapter 7) of the substances influences the choice of solvent, but there are also other factors involved. The solvent does not have to be a pure liquid, and very often mixtures are used.

All chromatography techniques involve a **stationary phase** and a **mobile phase**. The components in a mixture are separated because of their differences in affinity for the stationary and mobile phases. This is explained in Chapter 7.

Location of spots

If the substances to be separated are colourless (e.g. amino acids or sugars), then some method must be used to locate the spots on the paper or TLC plate. The spots may be located using a locating agent. Amino acids, which are colourless, may be located by spraying with ninhydrin, which makes them show up as pink or purple spots. Other methods that are useful for organic solutes are exposing the paper or plate to iodine vapour (the spots become brown) or spraying the plate with concentrated sulfuric acid then heating it (the spots appear as brown–black). Spots may also often be located by the use of an ultraviolet lamp, as some substances fluoresce under ultraviolet light.

TEST YOUR UNDERSTANDING

4 Select a technique that could be used to separate the components of the following mixtures:

- **a** sand from water
- **b** potassium chloride from a potassium chloride solution
- **c** different indicators in universal indicator solution
- **d** a mixture of ethoxyethane ($CH_3CH_2OCH_2CH_3$, boiling point 34 °C) and 1-(hexyloxy)hexane ($CH_3(CH_2)_5O(CH_2)_5CH_3$, boiling point 220 °C).

5 Explain how you would separate a mixture of potassium bromide (soluble in water) and calcium carbonate (insoluble in water).

6 Explain how you would separate iodine from an aqueous iodine solution, given that iodine is much more soluble in hexane than in water, and hexane is immiscible with water. *Immiscible* means that hexane and water do not mix – they form separate layers.

1.2 Kinetic molecular theory

Kinetic molecular theory (often just called kinetic theory) is a model that was developed originally to explain the properties of gases, but it is usually also extended to describe liquids and solids. Within this model, we describe all matter as being made up of individual particles that are in constant motion (hence, the word 'kinetic').

NATURE OF SCIENCE

Models are used throughout science. A model is a way of making sense of the world around us. Models may either be qualitative, as here, or quantitative (involving numbers and equations). The validity of a particular model can be tested by looking at how closely predictions made using the model agree with experimental observations.

The three states of matter most commonly encountered are solid, liquid and gas, and these differ in terms of the arrangement and movement of particles. The particles making up a substance may be individual atoms or molecules or ions. Simple diagrams of the three states of matter are shown in Figure 1.11, in which the individual particles are represented by spheres.

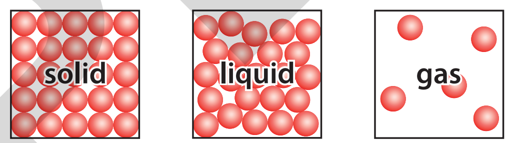


Figure 1.11: The three states of matter.

EXAM TIP

In diagrams showing the states of matter, remember the following:

- Solid: the particles should be arranged regularly and touching.
- Liquid: the particles are arranged randomly but still mostly touching.
- Gas: the particles are arranged randomly and are shown far apart.

Solid: the particles are generally regularly arranged and, due to relatively strong forces of attraction between them, are only able to vibrate about mean positions. As the forces between the particles are relatively strong, solids have fixed shapes.

Liquid: the particles have weaker forces between them, and so are able to move around each other. As the forces between the particles are weaker than those in solids, liquids take the shape of the container they are in. There are, however, still forces between the particles, so they stay together and do not fill the container.

Gas: the particles are assumed to have no forces between them (see Chapter 5) and move around randomly in all directions. There are no forces between the particles, so they are free to move around anywhere in a container, and thus, 'fill' the container. To give you some idea of how quickly the particles in a gas are moving: the average speed of molecules in air (mostly nitrogen and oxygen) at 25 °C and atmospheric pressure is almost 500 ms^{-1}; the particles collide, on average, every 150 ps (1.5×10^{-10} s) and only travel about 7×10^{-8} m between collisions.

The properties of the three states of matter are summarised in Table 1.1.

	Solids	Liquids	Gases
Distance between particles	close together	close but further apart than in solids	far apart
Arrangement	regular	random	random
Shape	fixed shape	no fixed shape; take the shape of the container	no fixed shape; fill the container
Volume	fixed	fixed	not fixed
Movement	vibrate	move around each other	move around in all directions
Speed of movement	slowest	faster	fastest
Energy	lowest	higher	highest
Forces of attraction	strongest	weaker	weakest

Table 1.1: Properties of the three states of matter.

Temperature

There are two temperature scales that are used commonly in everyday life: the Fahrenheit scale (melting point of ice = 32 °F and boiling point of water = 212 °F), which is used predominantly in the USA and a few other countries, and the Celsius or centigrade scale (melting point of ice = 0 °C and boiling point of water = 100 °C), which is used in the rest of the world. In science, however, we much more commonly use the absolute, or Kelvin, scale of temperature. For calculations involving temperatures in science, it is usually essential to use temperatures in kelvin.

KEY POINT

The kelvin is the SI unit of temperature.

SI stands for Système International and is the internationally accepted system of units used in science. Within the SI, seven base units are defined by reference to seven fundamental constants (such as the speed of light in a vacuum and the Planck constant), which have agreed specific values. Other SI base units include the second, the metre and the mole.

The absolute, or Kelvin, scale of temperature starts at **absolute zero**, which is the lowest temperature possible. All molecular motion does not actually stop at absolute zero (this would contravene the Heisenberg uncertainty principle), but it is the temperature at which everything would be in its lowest energy state. It is not possible to actually reach absolute zero, but scientists have managed to get very close – below one nanokelvin!

KEY POINT

Absolute zero the lowest temperature possible, corresponds to 0 K or –273.15 °C (usually taken as –273 °C).

A change of 1 °C is the same as a change of 1 K.

EXAM TIP

A temperature change in °C is the same as one in K.

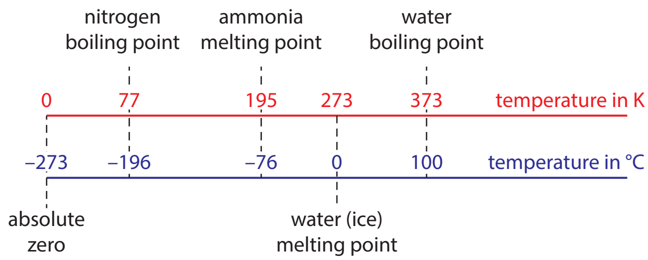


Figure 1.12 Some temperatures in K and °C.

Figure 1.12 compares some temperatures in K and °C. The fact that a change of 1 °C is the same as a change of 1 K makes it quite straightforward to convert temperatures between the two scales.

KEY POINTS

To convert °C into K, add 273.

To convert K into °C, subtract 273.

WORKED EXAMPLE 1.1

Convert a temperature of 25 °C into kelvin.

Answer

To do this, we add 273 to the temperature in °C:
25 + 273 = 298 K

WORKED EXAMPLE 1.2

Convert a temperature of 350 K into °C.

Answer

To do this, subtract 273 from the temperature in K:
350 − 273 = 77 °C

1.3 Temperature and kinetic energy

KEY POINT

Temperature is a measure of the average (mean) kinetic energy (E_k) of the particles in a substance.

- The higher the temperature, the higher the average kinetic energy of the particles.

The particles in gases and liquids are constantly colliding and, therefore, the particles will not all be moving at the same speed, and there will be a spread of kinetic energies for the particles, which is why we use the term *average* **kinetic energy**. The distribution of kinetic energies in a sample of gas at two different temperatures is shown in Figure 1.13. At higher temperature, there are fewer particles with lower kinetic energy and more particles with higher kinetic energy, and so, the average kinetic energy of the particles is greater. This will be explored in Chapter 17.

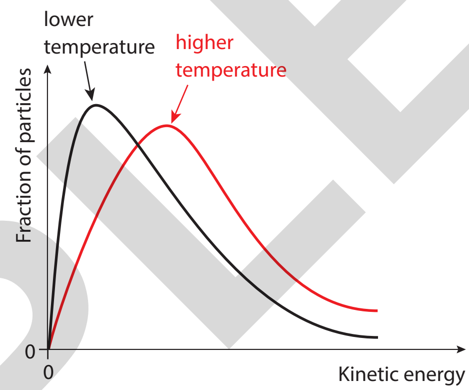


Figure 1.13: The distribution of kinetic energies in a sample of gas is called the Maxwell–Boltzmann distribution.

If two gases are at the same temperature, their particles will have the same average kinetic energy. This does not mean that the average speed of the particles is the same. Kinetic energy is calculated using the following equation:

$$E_k = \tfrac{1}{2}mv^2$$

where m is the mass of the particle and v is the speed.

This means that, the lighter the particles, the higher the average speed at a particular temperature. The average speed of carbon dioxide molecules (relative mass 44.01) at 25 °C is about 380 m s^{-1}, whereas the average speed of hydrogen molecules (relative mass 2.02) is about 1770 m s^{-1}.

TEST YOUR UNDERSTANDING

7 Convert each of the following temperatures in °C to temperatures in K.

a 25 °C

b 500 °C

c −100 °C

d −145 °C

CONTINUED

8 Convert each of the following temperatures in K to temperatures in °C.

- **a** 323 K
- **b** 100 K
- **c** 50 K
- **d** 500 K

9 What is wrong with the temperatures –50 K and –300 °C?

1.4 Changes of state

When one state of matter becomes another state of matter, we describe this as a change of state. Changes of state are summarised in Figure 1.14. Converting one state of matter into another usually involves heating (the change of state is an endothermic process) or cooling the substance (the change of state is an exothermic process) but can also be achieved by changing pressure. Endothermic and exothermic processes will be considered in Chapter 12.

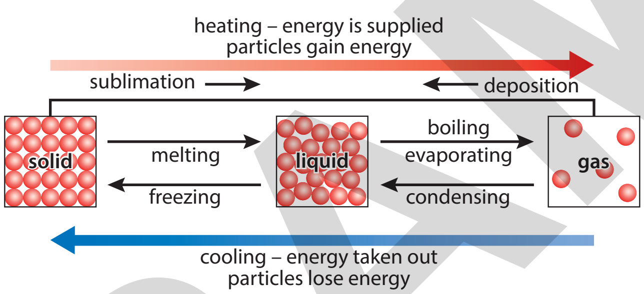


Figure 1.14: Changes of state. Note that evaporation can occur at any temperature, but boiling occurs at a fixed temperature.

Sublimation is the change of state when a substance goes directly from the solid state to the gaseous state, without going through the liquid state. Both iodine and solid carbon dioxide (dry ice) sublime at atmospheric pressure. The reverse process is called deposition.

The temperatures at which a substance changes state are called its melting point (change from solid to liquid) and boiling point (change from liquid to gas).

KEY POINT

A substance will be:

- a solid if the temperature is below its melting point
- a liquid if the temperature is between its melting point and its boiling point
- a gas if the temperature is above its boiling point

So, for example, bromine melts at –7.2 °C and boils at 58.8 °C; therefore, below –7.2 °C bromine will be a solid, between –7.2 °C and 58.8 °C it will be a liquid, and above 58.8 °C it will be a gas. There is no universally accepted definition of 'room temperature', but it is often taken as 25 °C, and so bromine is one of only two elements that is a liquid at room temperature.

Boiling and evaporation both involve a change in state from liquid to gas, but they are not the same thing – boiling only occurs at a certain temperature (the boiling point), but evaporation of the liquid can occur at any temperature between the melting and boiling points.

Using state symbols in equations

State symbols are used in chemical equations to indicate the physical state that the substances are in.

KEY POINT

The state symbols are:

(s) = solid

(l) = liquid

(g) = gas

(aq) = aqueous (dissolved in water)

We saw the following chemical equation in the previous section on Filtration

$$CuO(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l)$$

This indicates that solid copper(II) oxide (CuO) reacts with sulfuric acid, which is an aqueous solution (dissolved in water) to form an aqueous solution of copper(II) sulfate and liquid water.

Changes of state may be described using equations including state symbols, for example:

Melting of ice to form water: $H_2O(s) \rightarrow H_2O(l)$

Boiling/evaporation of liquid bromine: $Br_2(l) \rightarrow Br_2(g)$

Sublimation of iodine: $I_2(s) \rightarrow I_2(g)$

How to write balanced chemical equations will be discussed in Chapter 16.

Temperature during changes of state

If a pure substance is heated slowly, from below its melting point to above its boiling point, a graph of temperature against time can be obtained (Figure 1.15).

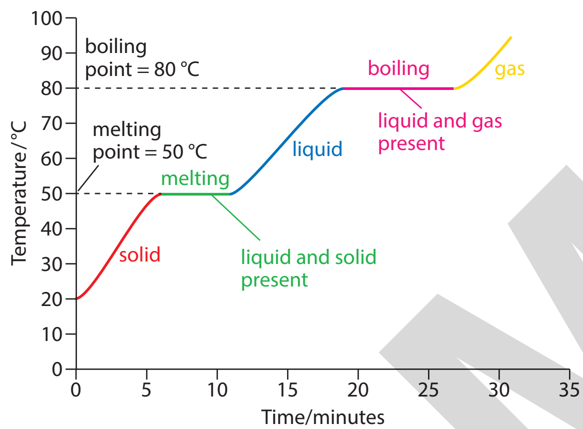


Figure 1.15: A heating curve showing changes of state.

As a solid is heated, its particles vibrate more violently. The particles gain kinetic energy and the temperature of the solid rises. At 50 °C, the solid in Figure 1.15 begins to melt – at this stage, there is solid and liquid present together, and the temperature remains constant until all the solid has melted. All the heat energy being supplied is used to partially overcome the forces of attraction between particles, so that they can move around each other. Another way of saying this is that, at the melting point, all the heat energy being supplied goes into increasing the potential energy of the substance (overcoming forces between particles) and not to increasing the kinetic energy, so the temperature does not change.

When all the solid has melted, the continued supply of heat energy causes the kinetic energy of the particles to increase, so that the particles in the liquid move around each other more quickly and the temperature increases. The average kinetic energy of the particles increases, until the boiling point of the liquid is reached. At this point (80 °C), the continued supply of heat energy is used to overcome the forces of attraction between the particles completely and the temperature of the substance remains constant, until all the liquid has been converted into gas. The continued supply of heat energy then increases the average kinetic energy of the particles and, therefore, the temperature of the gas. The particles move around faster and faster, as the temperature of the gas increases.

TEST YOUR UNDERSTANDING

10 State the names of the following changes of state:

- **a** from solid to liquid
- **b** from solid to gas
- **c** from gas to liquid
- **d** from gas to solid.

11 Use data in the table to determine whether each of the elements will be a solid, liquid or gas at the specified temperature:

Substance	Melting point / °C	Boiling point / °C
Magnesium	650	1090
Fluorine	–220	–188
Polonium	254	962
Mercury	–39	357

- **a** magnesium at 100 °C
- **b** fluorine at –200 °C
- **c** polonium at 1000 °C
- **d** mercury at 25 °C.

SELF-ASSESSMENT CHECKLIST

Think about the topics covered in this chapter. Which parts are you most confident with? Which topics require some extra practice?

I can...	Section	Needs more work	Nearly there	Confident to move on
explain the terms element, compound and mixture, and distinguish between them	1.1			
explain the difference between heterogeneous and homogeneous mixtures and give examples of each	1.1			
explain the different methods for separating the components of a mixture and suggest a suitable method for separating a particular mixture	1.1			
explain the properties of solids, liquids and gases in terms of kinetic molecular theory	1.2			
state the relationship between temperature in K and the average kinetic energy of particles	1.3			
convert temperatures between K and °C	1.3			
use state symbols in chemical equations	1.4			
explain changes of state in terms of kinetic molecular theory.	1.4			

REFLECTION

To what extent do you feel that you have met many of the ideas in this chapter before? Can you highlight specific areas that are new to you? Are you confident with these areas? Can you use your knowledge to identify heterogeneous and homogeneous mixtures around your home or school? Do you think that you could explain the difference between elements, compounds and mixtures to another student?

EXAM-STYLE QUESTIONS

You can find questions in the style of IB exams in the digital coursebook.

Exam-style questions

A periodic table is required to answer some of these questions.
The multiple-choice questions can all be answered without the aid of a calculator.

Chapter 1

1 Which of the following contains an element, a compound and a mixture? **[1]**

A $H_2O(l)$, $H_2(g)$, $FeS(s)$

B $Cl_2(aq)$, $Br_2(g)$, $NaBr(l)$

C $CH_4(g)$, $I_2(l)$, $CO_2(l)$

D $NaCl(aq)$, $CO(g)$, $H_2S(g)$

2 Which of the following is a homogeneous mixture? **[1]**

A a mixture of sand and sodium chloride

B a sodium chloride solution

C a mixture of hexane and water

D a mixture of sulfur and iron

3 Consider the following process: $I_2(g) \rightarrow I_2(s)$

The name of this process is **[1]**

A condensation

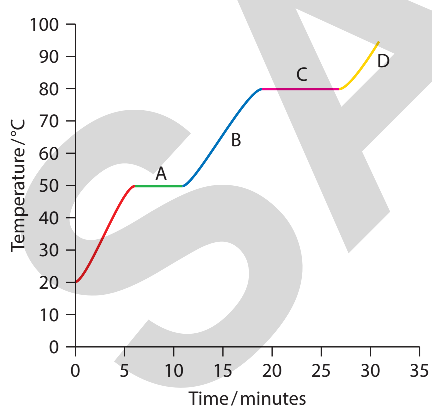
B sublimation

C deposition

D vaporisation

4 A substance, X, which is a solid at room temperature, is heated and the temperature monitored.

The graph of temperature against time is shown.



At which point are a solid and a liquid present? **[1]**

A

B

C

D

5 Rubidium has a melting point of 39 °C and a boiling point of 688 °C.
What are the melting and boiling points of rubidium in kelvin? **[1]**

	Melting point / K	Boiling point / K
A	–234	415
B	234	415
C	312	961
D	39	688

6 In which of the following is the temperature in K higher than the temperature in °C? **[1]**

A 100 °C 250 K
B 150 °C 500 K
C –100 °C 100 K
D 0 °C 250 K

7 Lead(II) iodide (PbI_2) can be made by adding a solution of potassium iodide to a solution of lead(II) nitrate.

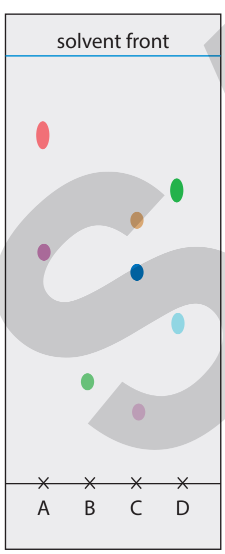
The equation for the reaction is:

$$2KI(aq) + Pb(NO_3)_2(aq) \rightarrow PbI_2(s) + 2KNO_3(aq)$$

Which method could be used to most easily separate lead iodide from the reaction mixture? **[1]**

A distillation
B filtration
C evaporation
D solvation

8 Spots of four substances were put on the baseline of a piece of chromatography paper in the positions marked with an × in the diagram. The resulting chromatogram is shown.



Which of the substances is pure? **[1]**

A **C**
B **D**

9 Ethanol (boiling point 78 °C) is a very good solvent for many organic compounds.

Which liquid is most likely to be completely separated from an ethanol solution using distillation? **[1]**

A hexane (boiling point 69 °C)

B 3-ethylpentane (boiling point 93 °C)

C propan-2-ol (boiling point 82 °C)

D propane-1,2,3-triol (boiling point 289 °C)

10 Which of the following statements about kinetic molecular theory is correct? **[1]**

A In gases, the particles vibrate about mean positions.

B In liquids, there are no forces between particles.

C The particles in a gas all have the same kinetic energy.

D The average kinetic energy of particles increases as temperature increases.

11 The melting and boiling points of some substance are shown in the table. **[1]**

	Melting point / K	Boiling point / K
ethyl benzoate ($C_9H_{10}O_2$)	238	486
anthracene ($C_{14}H_{10}$)	489	614
propanone (C_3H_6O)	178	329
ethene (C_2H_4)	104	169

a Explain which substances are liquids at 25 °C. **[2]**

b Ethyl benzoate is soluble in propanone. Molly is given a solution containing 10 g of ethyl benzoate and 100 g of propanone. Explain how she could separate the components of this mixture. **[3]**

c Anthracene is only slightly soluble in propanone with its solubility being about 1 g per 100 g of propanone. 10 g of anthracene and 100 g of propanone are shaken together for a few minutes.

Explain how all the anthracene could be extracted from this mixture. **[2]**

d Rosalie makes the statement: 'there is no temperature at which all four substances will be liquids'. Evaluate this statement. **[1]**

12 Zinc sulfate can be made by reacting excess zinc with dilute sulfuric acid according to the following equation:

$$Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$$

Excess means that there will still be some zinc left in the reaction mixture when the reaction has finished, but all of the sulfuric acid should have reacted.

a Classify each of the reactants as an element, compound or mixture. **[1]**

b Explain how you could obtain a solid sample of zinc sulfate from the reaction mixture. **[2]**

c Kinetic molecular theory is a model that can be used to explain the properties of solids, liquids and gases. Describe the difference in the motion of the particles in zinc and hydrogen at room temperature. **[2]**

d The melting point of zinc is 420 °C. Sketch a graph showing how the temperature of a sample of zinc varies with time, as it is heated slowly from 400 °C to 440 °C. Identify the physical state of zinc in each region of your graph. **[2]**

13 A student has been given 50 cm^3 of a solution that contains 1 g of caffeine and 1 g of sodium chloride.

a Explain why pure caffeine cannot be extracted from this mixture by heating to evaporate off the water. **[1]**

b Some data about three solvents is given in the table.

Solvent	Caffeine solubility / g per 100 g of solvent	Solvent miscibility with water	Sodium chloride solubility in solvent
water	2.3	miscible	soluble
propanone	1.5	miscible	insoluble
trichloromethane	11.6	immiscible	insoluble

i Give two reasons why trichloromethane can be used to extract caffeine from the mixture but propanone cannot. **[2]**

ii Describe how caffeine can be extracted from the mixture. **[3]**

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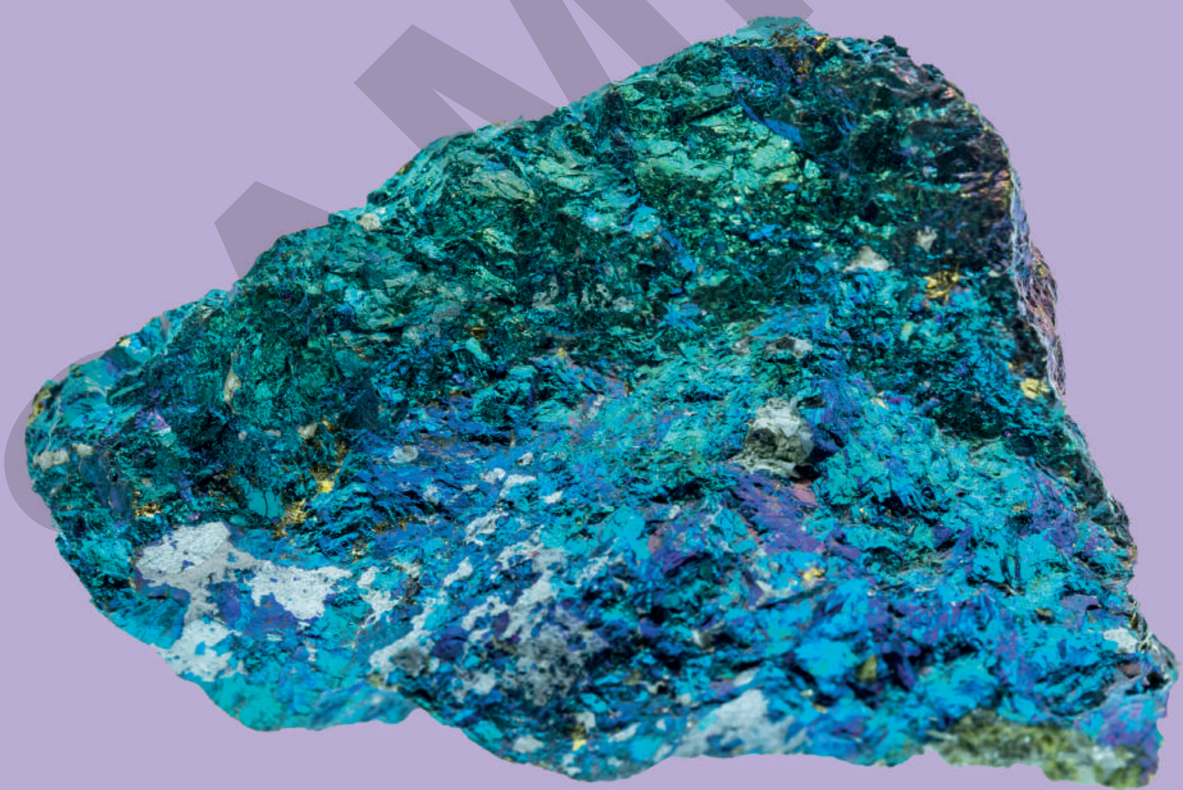
Chemistry

for the IB Diploma

WORKBOOK

Jacqueline Paris



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Contents

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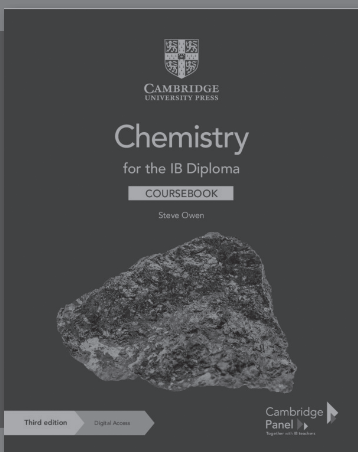
This suite of resources supports students and teachers of the Chemistry course for the IB Diploma programme. All of the books in the series work together to help students develop the necessary knowledge and scientific skills required for this subject.



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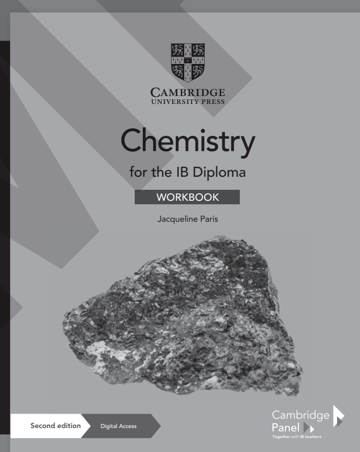

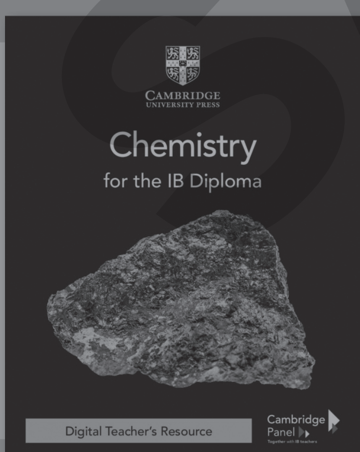

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How to use this book

A chapter outline appears at the start of every chapter to introduce the learning aims and help you navigate the content.

CHAPTER OUTLINE

In this chapter you will:

- describe the structure of the atom and the relative charges and masses of protons, neutrons and electrons
- describe how protons, neutrons and electrons behave in electric fields
- deduce the number of protons, neutrons and electrons in atoms and ions.

KEY TERMS

Definitions of key vocabulary are given at the beginning of each chapter.

You will also find definitions of these words in the glossary.

Exercises

Exercises help you to practice skills that are important for studying Standard Level and Higher Level Chemistry.

TIP

Tip boxes will help you complete the exercises, and give you support in areas that you might find difficult.

EXAM-STYLE QUESTIONS

Questions at the end of each chapter are more demanding exam-style questions, some of which may require use of knowledge from previous chapters. Answers to these questions can be found in digital form on Cambridge GO.

Visit Cambridge GO and register to access these resources at www.cambridge.org/GO

KEY EQUATIONS

In these boxes you find chemical equations in the form of symbols and formula.

Unit 1

The nature of matter

Chapter 1

The particulate nature of matter

CHAPTER OUTLINE

In this chapter you will:

- understand the terms *element*, *compound* and *mixture*
- understand the differences between heterogeneous and homogeneous mixtures
- understand how to separate the components of a mixture
- use kinetic molecular theory to understand the properties of solids, liquids and gases
- understand that temperature in K is proportional to the average kinetic energy of particles
- understand how to convert temperatures between K and °C
- use state symbols in chemical equations
- use kinetic molecular theory to explain changes of state.

KEY TERMS

Make sure you understand the following key terms before you do the exercises.

atom: the smallest part of an element that can still be recognised as that element; in the simplest picture of the atom, the electrons orbit around the central nucleus; the nucleus is made up of protons and neutrons (except for a hydrogen atom, which has no neutrons)

element: a chemical substance that cannot be broken down into a simpler substance by chemical means. Each atom has the same number of protons in the nucleus

compound: a pure substance formed when two or more elements combine chemically in a fixed ratio

mixture: two or more substances mixed together. The components of a mixture can be mixed together in any proportion (although there are limits for solutions). The components of a mixture are not chemically bonded together, and so, retain their individual properties. The components of a mixture can be separated from each other by physical processes

CONTINUED

molecule: an electrically neutral particle consisting of two or more atoms chemically bonded together

heterogeneous mixture: a mixture of two or more substances, that does not have uniform composition and consists of separate phases. A heterogeneous mixture can be separated by mechanical means. An example is a mixture of two solids

chemical properties: how a substance behaves in chemical reactions

chromatography: a technique used to separate the components of a mixture due to their different affinities for another substance and/or solubility in a solvent

deposition: the change of state from a gas to a solid

filtration: a separation technique used to separate insoluble solids from a liquid or solution

physical properties: properties such as melting point, solubility and electrical conductivity, relating to the physical state of a substance and the physical changes it can undergo

solvation: a process used to separate a mixture of two or more substances, due to differences in solubility

states of matter: solid, liquid and gas

state symbols: used to indicate the physical state of an element or compound; these may be either written as subscripts after the chemical formula or in normal type: (aq) = aqueous (dissolved in water); (g) = gas; (l) = liquid; (s) = solid

boiling: change of state from a liquid to a gas at the boiling point of the substance

boiling point: the temperature at which a liquid boils under a specific set of conditions - usually we will be considering the boiling point at atmospheric pressure

distillation: a separation technique used to separate the solvent from a solution or separate liquid components of a mixture that have different boiling points

sublimation: the change of state from a solid to a gas

melting: the change of state from a solid to a liquid

freezing: the change of state from a liquid to a solid

melting point: the temperature at which melting occurs

homogeneous mixture: a mixture of two or more substances with the same (uniform) composition throughout the mixture – it consists of only one phase. Examples are solutions or a mixture of gases

solution: that which is formed when a solute dissolves in a solvent

CONTINUED

evaporation: the change of state from a liquid to a gas that can occur at any temperature above the melting point

solute: a substance that is dissolved in another substance (the solvent) to form a solution

solvent: a substance that dissolves another substance (the solute); the solvent should be present in excess of the solute

temperature: a measure of the average kinetic energy of particles

Exercise 1.1 Elements, compounds and mixtures

This exercise will check you understand the key terms **element**, **compound**, **mixture**, **atom** and **molecule**, which are important fundamental ideas in chemistry.

1 Approximately how many different elements are there?

2 Some elements exist as individual atoms, some as a small group of atoms bonded together into a molecule and others are bonded together into a giant structure.

- **a** Name two elements that exist as giant structures at 25 °C.
- **b** Name an element that exists as a single atom.
- **c** Name an element that exists as a molecule made of two atoms joined together (a diatomic molecule).

3 Identify which of the following formulas represent atoms and which represent molecules:

- **a** He
- **b** O_2
- **c** H_2O
- **d** C

4 Identify which of the following formulas represent elements and which represent compounds:

- **a** He
- **b** O_2
- **c** H_2O
- **d** C

TIP

An atom is a single particle.

A molecule is made up of more than one atom.

The atoms in a molecule can be of the same element.

5 This statement is incorrect, explain why:

Elements are made of atoms and compounds are made of molecules.

TIP

Question 5 is linked to ideas in Chapter 6.

6 An alloy is a mixture of a metal and other elements. Give one way in which the composition of an alloy differs from that of a compound.

7 Compounds have both different **chemical properties** and **physical properties** from the elements from which they are formed.

a What is meant by the term *physical properties*?

b What is meant by the term *chemical properties*?

8 Most everyday substances are mixtures although they are often labelled as pure. Pure orange juice is a common example. The manufacturers simply mean that nothing has been added to the orange juice. In chemistry, the term *pure* is not used in the same way.

In chemistry, what is meant by the term *pure*?

9 Why do the components of a mixture retain their individual properties?

10 Group the following substances into elements, mixtures and compounds:

air, water, sodium chloride solution, sodium chloride crystals, iron, chlorine gas, carbon dioxide gas.

11 What name is given to a mixture that has a uniform composition and only consists of one phase?

TIP

Solid, liquid, gas and solution are all examples of phases.

12 What name is given to a mixture that does not have a uniform composition and consists of separate phases?

13 Why is a mixture of the solids sodium chloride and sand not a **homogeneous mixture**?

14 When a small amount of salt and water is mixed together, it forms a homogeneous mixture, but this is not true when flour is mixed with water, why?

15 Are chemical or physical processes typically used to separate the components of a mixture?

16 Match the name of the separation technique with the type of mixture it can be used to separate.

Technique	Types of mixture
A **filtration**	**1** substances with very different solubilities in a **solvent**
B **distillation**	**2** an insoluble solid from a liquid
C **evaporation**	**3** a **solute** with very different solubilities in two different solvents
D **solvation**	**4** the solute from a **solution**
E solvent extraction	**5** liquids with a large difference in their solubilities in different solvents
F paper **chromatography**	**6** a mixture of substances with small differences in their solubilities in a solvent

Exercise 1.2 Kinetic molecular theory

Kinetic molecular theory is used to explain the observed properties of solids, liquids and gases.

1 Complete Table **1.1**, which describes the arrangement and movement of particles in solids, liquids and gases.

	Solids	Liquids	Gases
diagram showing the arrangement of the particles			
relative distance of the particles from one another			
relative energy of the particles			
movement of particles			
relative force of attraction between the particles			

Table 1.1: Arrangement and movement of particles.

2 Which of the descriptions of particles in Table **1.1** can explain the fixed shape of solids and the lack of a fixed shape in liquids and gases?

3 Which of the descriptions in Table **1.1** explain why, at a given **temperature**, the volume of a gas is not fixed but the volume of solids and liquids are?

4 Younger students are often confused by the observed properties of a powder. A powder can flow like a liquid and take up the shape of its container but does not completely spread out into a puddle like a liquid.

a How would you explain that a powder is a solid?

b How would you explain the ability of a powder to flow like a liquid?

5 Which scale is the SI scale for temperature?

6 On the kelvin scale, what does zero K (or absolute zero) represent?

7 Complete Table **1.2** to show equivalent temperatures on the kelvin and Celsius scales.

Celsius scale	Kelvin scale
0	
	373
40	
	74
946	
	3
	500

Table 1.2: Equivalent temperatures on the kelvin and Celsius scales.

TIP

Absolute zero equals –273.15°C, but you can use –273 °C for your chemical calculations.

8 Temperature is used in some chemical calculations. When it is, the kelvin scale is always used, unless the calculation involves a temperature change.

Explain why either Celsius or kelvin can be used to measure temperature change.

Exercise 1.3 Temperature and kinetic energy

Not all of the particles in a sample have the same amount of energy, and so, they do not all move with the same speed. In this exercise, you will explore the distribution of kinetic energies at different temperatures.

1 Consider a sample of oxygen at a constant temperature.

a Do all the oxygen particles have the same kinetic energy? Explain your answer.

b Do all the particles of the gas move at the same speed? Explain your answer.

2 Consider a mixture of the gases nitrogen and helium at a constant temperature.

a The average kinetic energy of the particles will be higher for which gas?

b The average speed of the particles will be higher for which gas?

3 Describe how the following change when the temperature of a gas is increased:

a the average kinetic energy of the particles

b the average speed of the particles

c the most probable kinetic energy of the particles

d the fraction of particles with the most probable kinetic energy.

TIP

The most probable kinetic energy is the energy at the peak of a Maxwell–Boltzmann distribution curve.

Exercise 1.4 Changes of state

Heating or cooling a substance can cause it to change state, as these processes involve the breaking or formation of forces of attraction between the particles. In this exercise, you will check that you understand these processes and can work out the state of a substance at a given temperature from its **melting point** and **boiling point**.

Figure **1.1** summarises the changes of state.

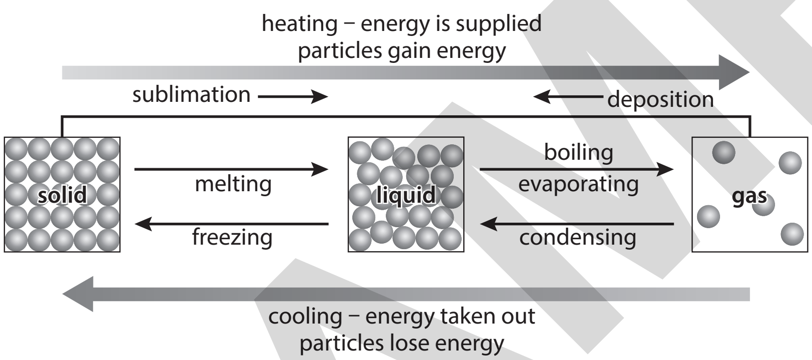

Figure 1.1: The changes of state.

1 Which change of state does not take place only at a fixed temperature for a given pressure?

2 Identify which changes of state are exothermic and which are endothermic.

3 What name is given to the temperature at which a substance changes from a liquid to a solid?

4 What name is given to the temperature at which a substance changes between gas and liquid?

5 Carbon dioxide and iodine are two examples of substances that undergo **sublimation**.

a What is meant by the term *sublimation*?

b What term is used to describe the reverse of sublimation?

TIP

The same name for the temperature at which the change in question 3 happens is used, no matter in which direction the change happens.

6 Complete the table to show whether a substance is a solid, liquid or gas at the temperature stated in the column header.

Substance	Melting point / °C	Boiling point / °C	State at −50 °C	State at 115 °C	State at 200 K
A	15	125			
B	253	578			
C	−83	78			
D	−169	−87			

7 Figure **1.2** shows the cooling curve for a substance.

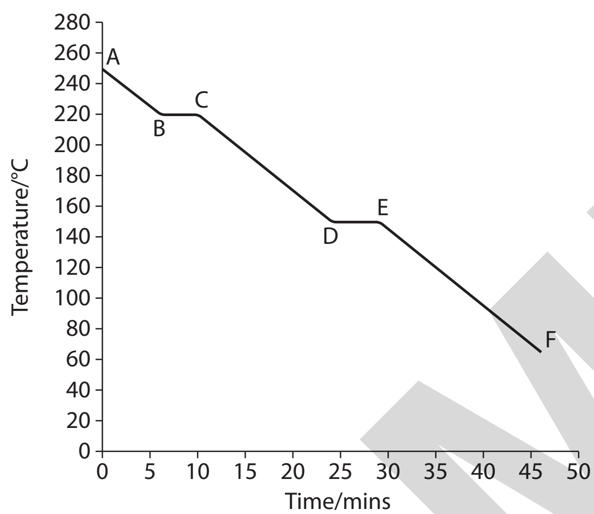

Figure 1.2: The cooling curve for a substance.

a Label the diagram to show the following:

i the region where the substance is a solid

ii the region where the substance is a liquid

iii the region where the substance is a gas

iv the region where the substance is freezing

v the region where the substance is condensing

vi the melting point of the substance

vii the boiling point of the substance.

b Explain, in terms of the movement and arrangement of the particles, why the temperature of the substance remains the same during a change of state.

EXAM-STYLE QUESTIONS

1 Which of the following lists substances that are all made up of molecules?

A C, O_2, CO_2

B Na, Cl_2, NaCl

C H_2, He, Li

D P_4, S_8, O_3 **[1]**

2 Which of the following statements is **true** of heterogeneous mixtures?

A Their components cannot be separated by physical means.

B They have the same composition throughout the mixture.

C The components are in a fixed ratio.

D The components are in separate phases. **[1]**

3 Which of the following is **not** a heterogeneous mixture?

A cola

B tea with milk

C tea with sugar

D milk **[1]**

4 Which of the following shows the correct sequence of the changes of state involved in distillation?

A boiling, condensing

B condensing, boiling

C evaporation, cooling

D boiling, cooling **[1]**

5 What is the name given to the separation technique that is used to separate the components of a mixture that have different solubilities in a solvent at different temperatures?

A distillation

B recrystallisation

C evaporation

D paper chromatography **[1]**

6 Mercury is a liquid at 25 °C, which of the following could be its melting and boiling points?

	Melting point	Boiling point
A	–38.9 °C	83.7 K
B	–38.9 K	629.7 °C
C	–38.9 K	356.7 K
D	–38.9 °C	356.7 °C

[1]

CONTINUED

7 Which is the correct equation for sublimation?

A $CO_2(s) \rightarrow CO_2(g)$

B $CO_2(g) \rightarrow CO_2(s)$

C $H_2O(s) \rightarrow H_2O(l)$

D $CO_2(g) \rightarrow CO_2(aq)$ **[1]**

8 Which statement is correct about melting?

A The average kinetic energy of the particles increases, but the temperature stays the same.

B The average kinetic energy of the particles increases, and the temperature increases.

C The average kinetic energy of the particles stays the same, but the temperature increases.

D The average kinetic energy of the particles stays the same, and the temperature stays the same. **[1]**

9 Ammonia liquid boils at –33 °C and freezes at –78 °C at atmospheric pressure.

a Predict the state of ammonia at

i –50 °C

ii –80 °C

iii 200 K. **[3]**

b Sketch a graph of temperature against time as a sample of ammonia is cooled from 0 °C to –50 °C. **[4]**

10 Some seaweeds accumulate iodide ions in their leaves, and so, are a good source of iodine. The seaweed must first be dried and then heated to burn off the organic matter. The remaining ash is then boiled in water and allowed to cool. The iodide ions dissolve in the water.

a Suggest a suitable technique that could be used to separate the iodide solution from any insoluble impurities. **[1]**

b State the type of mixture that remains after the insoluble impurities have been removed. **[1]**

c When dilute sulfuric acid and hydrogen peroxide are added to the mixture, an aqueous solution of iodine is produced:

$2H^+ + H_2O_2 + 2I^- \rightarrow I_2 + 2H_2O$

Give the state symbols for I_2 and H_2O in the equation above. **[2]**

d Iodine is not particularly soluble in water. It is much more soluble in organic solvents such as cyclohexane. Outline a method that could be used to separate the iodine from the solution. **[3]**

11 The statements below describe the analysis of a mixture of amino acids by paper chromatography.

a Place the statements in the correct order:

i spray the plate with a locating agent

ii mark the position of the solvent front

iii place a small sample of the unknown sample on the bottom of a piece of chromatography paper

iv place the paper into a tank containing a suitable solvent

v allow the solvent to rise up the paper **[3]**

CONTINUED

b Figure **1.3** shows the results of the experiment.

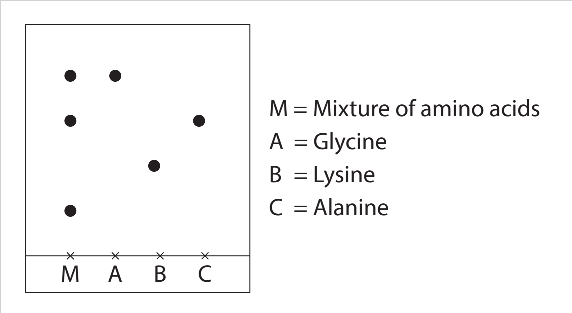

Figure 1.3: Chromatogram of an amino acid mixture.

Which amino acids did the mixture contain? **[1]**

c Why do substances A, B and C each only produce one spot on the chromatogram? **[1]**

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Chemistry

for the IB Diploma

Digital Teacher's Resource

Cambridge Panel
Together with IB teachers

Table of contents

Unit 3 Classification of matter

Unit 4 What drives chemical reactions?

Unit 5 How much, how fast, how far?

Digital resources

The following items are available on Cambridge GO. For more information on how to access and use your digital resource, please see inside front cover.

1 The particulate nature of matter

Teaching plan

Sub-chapter	Approximate number of learning hours	Learning content	Resources
1.1 Elements, compounds and mixtures	2–3	Recall the definitions of elements, compounds and mixtures. Distinguish between the properties of an element, compound or mixture. Understand the difference between homogeneous and heterogeneous mixtures. Describe experimental techniques to separate mixtures.	**Coursebook** Section 1.1 Test your understanding Questions 2 and 3 **Workbook** Exercise 1.1 **Teacher's resource** PowerPoint 1 Slides 2–5 Worksheet 1.1 Questions 1, 3 and 5, End of chapter test Questions 1–6, 9, 10
1.2 Kinetic molecular theory 1.3 Temperature and kinetic energy 1.4 Changes of state	2 2 2	Determine the state symbols in chemical equations. Recall the names of the changes of state. Explain the physical properties of matter and changes of states using kinetic molecular theory. Understand that temperature in kelvin is a measure of average kinetic energy of particles. Know how to convert between Celsius and kelvin scales.	**Coursebook** Sections 1.2–1.4 Test your understanding Question 12 **Workbook** Exercises 1.2–1.4 **Teacher's resource** PowerPoint 1 Slide 6 Worksheet 1.1 Questions 2 and 4 End of chapter test Questions 7 and 8

BACKGROUND KNOWLEDGE

- Understand how to classify substances as elements, compounds or mixtures.
- Describe simple techniques for separating mixtures (filtration, distillation, evaporation and paper chromatography).
- Draw particle diagrams and use them to explain the properties of solids, liquids and gases.
- Know the names of the interconversions of the three states of matter.
- Know there are two different scales for measuring temperatures: degree Celsius and kelvin.
- Know how to use a data logger and a temperature probe, plot line graphs and draw lines of best fit.

Syllabus overview

- The first part of the syllabus covers the concepts of elements, compounds, mixtures and the application of kinetic molecular theory to explain the particle models of states of matter. Chemists should know the differences between compounds and mixtures and how to construct names and formulas of compounds. This will facilitate the study of chemical reactions using balanced symbol equations and how to solve problems using molar ratios of reactants and products (Chapters 4 and 16).
- There are many opportunities for students to practice fundamental laboratory techniques, covering the various methods for separating mixtures. Students should be encouraged to think about how to test for the purity of products after separation and research into how to purify products further. When measuring melting/cooling curves of substances, students also practice mathematical skills of presenting their data graphically and analysing the results to extract information on melting/boiling points.
- Simulations can be used to illustrate molecular movement of particles. This gives an introduction on how the kinetic energy of particles is distributed in a sample of gas at a fixed temperature and the concept of activation energy in a chemical reaction (Chapter 17).

1.1 Elements, compounds and mixtures

LEARNING PLAN

Learning objectives	Success criteria
Understand the terms element, compound and mixture	Students should be able to explain the terms element, compound and mixture and distinguish between them.
Understand the differences between heterogeneous and homogeneous mixtures	Students should be able to explain the difference between heterogeneous and homogeneous mixtures and give examples of each.
Understand how to separate the components of a mixture	Students should be able to explain the different methods for separating the components of a mixture and suggest a suitable method for separating a particular mixture.

Common misconceptions

Misconception	How to identify	How to overcome
Students confuse the meaning of compounds and molecules	Ask students to assign various names and formulas of elements and compounds to a Venn diagram of two circles labelled compounds and molecules.	Draw particle diagrams to show which names/formulas are molecules or compounds. Molecules can be elements (O_2) or compounds (H_2O) and only covalent compounds are molecules.
Students confuse physical and chemical changes	Show students pictures of different processes (for example, physical processes, including melting, freezing and sublimation, and chemical processes, including rusting, fireworks and cooking an egg) and ask them to distinguish the physical from the chemical changes.	A physical change is one in which no new chemicals are formed, for example, dissolving and changes of states (in separating mixtures). A chemical reaction involves making new substances. Teachers can demonstrate some examples when elements are combined in chemical reactions to form compounds. For example, burning Na in Cl_2 or Mg in O_2. Use particle diagrams to show that the microscopic make-up of the reactants and products is different, and the atoms are bonded together differently.

Starter ideas

1 Recap prior knowledge from pre-IB (10 minutes)

Resources: Test your understanding questions 2 and 3 in the Coursebook.

Description and purpose: Students define element, compound and mixture. They should then sort out the listed substances and diagrammatic representations into the three categories. This activity assesses students' prior knowledge.

What to do next: If most of the students can define element, compound and mixture and identify them correctly, teachers can ask them to give more examples of each. Make sure to emphasise the keywords in the definitions. If students find it difficult to distinguish amongst the three categories, help by pointing out that elements can be found in the periodic table, elements combine chemically to form compounds and give examples of names and formulas of various compounds. Most of things we meet daily are mixtures and can be separated by physical methods.

› **Language focus:** Learners are encouraged to pay attention to definitions of the key terms.

Main teaching ideas

2 Teacher demonstrations on the formation of compounds from constituent elements (20 minutes)

Resources: Search for the websites mentioned in the Description and purpose line for apparatus/chemicals required for each demonstration.

Description and purpose: Iron and sulphur (search the 'Royal Society of Chemistry' website with the keywords 'iron and sulfur reaction')

Sodium and chlorine (search the 'University of Washington' website with the keywords 'sodium and chlorine reaction')

These experiments can be performed to show how elements retain their properties in a mixture but change their properties when forming compounds. Ask students to record the observations (changes in physical states, colours, endothermic vs exothermic etc.) during the reaction and write word and symbol equations to represent the chemical processes.

Differentiation ideas

Support: Provide students with a table to record the appearance of the reactants and products of the reactions, and their observations during the reactions. Stress the importance of forming new substances in chemical reactions to form compounds.

Stretch and challenge: Students can be asked to construct balanced chemical equations for these reactions with state symbols.

Language focus: Recording observations. This is one way of checking the correct use of terminology.

3 Student practical (2 × 45 minutes)

Resources: A mixture of sand and water, sodium chloride solution, a mixture of ink and water, a mixture of food dyes. Apparatus required for filtration (for example, funnel, filter paper, clamp, boss head, stand, beaker), simple distillation (for example, round-bottomed flask, thermometer, bung, Liebig condenser, beaker, Bunsen burner, heat-proof mat), evaporation (evaporating basin, gauze, Bunsen burner, heat-proof mat) and paper chromatography (chromatography paper, beaker, pencil, ruler, small capillary tube).

Description and purpose: Ask students to separate various mixtures, including sand and water (filtration), table salt dissolved in water (simple distillation to keep the water or evaporation to obtain only the salt crystal), ink and water (simple distillation), a mixture of food dyes (paper chromatography). The practical could be run at different stations set up around a laboratory.

Differentiation ideas

Support: Providing exact step-by-step methods with diagrams to guide students through the practical.

Stretch and challenge: Students design their own methods and carry out the experiments once their methods are approved by a teacher.

Plenary ideas

1 How to separate mixtures (10 minutes)

Resources: Fill in the information in the following table on how to separate mixtures. The first two rows have been completed as an example.

Separation of …	Homogeneous or heterogeneous mixtures	Technique	Example
two liquids	homogeneous	simple distillation: the liquid with a lower boiling point will boil first	ink and water: water will boil first
two liquids	heterogeneous	two liquids that are immiscible and have different densities can be separated into layers in a separatory funnel	water and dichloromethane: dichloromethane is denser, so it will come out of the funnel first
a solid and a liquid	homogeneous		
a solid and a liquid	heterogeneous		
two solids	heterogeneous		

Description and purpose: This exercise gives students an opportunity to summarise, recall and apply their knowledge.

Language focus: Take note of the language used when making a summary of the experimental methods.

1.2 Kinetic molecular theory; 1.3 Temperature and kinetic energy and 1.4 Changes of state

LEARNING PLAN

Learning objectives	Success criteria
Use kinetic molecular theory to understand theproperties of solids, liquids and gases	Students should be able to explain the properties of solids, liquids and gases in terms of kinetic molecular theory.
Understand that temperature in K is proportional to the average kinetic energy of particles	Students can recall that temperature in K is proportional to the average kinetic energy of particles.
Understand how to convert temperatures between K and °C	Students should be able to convert temperatures between K and °C.
Use state symbols in chemical equations	Students can apply state symbols in chemical equations.
Use kinetic molecular theory to explain changes of state	Students should be able to explain changes of state using kinetic molecular theory.

Common misconceptions

Misconception	How to identify	How to overcome
Students confuse boiling and evaporation	Ask students to explain the differences between boiling and evaporation.	In both processes, liquids change to gases. Boiling occurs at a particular temperature (boiling point) and throughout the whole of the liquid. Evaporation can occur at all temperatures but only on the surface of the liquid.
Changes in temperature on kelvin and Celsius scales are muddled up	Ask students what is $\Delta T = 30$ °C when converted to kelvin scale.	The intervals on both the temperature scales are the same, so the changes in temperature can have either K or °C as units but the numerical values remain the same. A change of 20 °C to 50 °C (30 °C) has the same value as a change of 293 K to 323 K (30 K) when converted to the kelvin scale.
Students struggle to understand what is in the space between particles in the particle model	Show students a model of the giant ionic lattice of NaCl or a model of the molecular structure of ice and ask them what is in between the particles in the model.	Students often have the misconception that the space is filled with air. Air is a mixture of many gas molecules/atoms, and these entities are themselves too big to fit into the space between particles in the NaCl ions or H_2O molecules in the solid state.

Starter ideas

1 Solid, liquid and gas particle diagrams (10 minutes)

Resources: A piece of A4 paper showing three equal-sized square boxes.

Description and purpose: Students complete diagrams showing the arrangement of particles in a solid, liquid and gas and then name the processes for the changes of state (including sublimation and deposition). This activity recaps students' knowledge from pre-IB.

What to do next: Students should be clearly aware of how the particle arrangements are represented in these diagrams.

2 Recognise the states of matter based on melting and boiling points (10 minutes)

Resources: Test your understanding Question 12 in the Coursebook.

Description and purpose: Ask students to identify the states of matter at given temperatures. Students should be able to apply their knowledge of melting and boiling points to recognise the states of matter.

What to do next: Show the melting/boiling points on a number line to order them, if students find this activity difficult.

Main teaching ideas

1 Practical on the freezing of stearic acid (45 minutes)

Resources: A detailed list of apparatus and chemicals can be found by searching the 'Royal Society of Chemistry' website with the keywords 'freezing of stearic acid'. Graph paper is required for analysing the results.

Description and purpose: Students need to heat up 3 spatulas of stearic acid until it melts. Then allow the acid to cool and take a temperature reading every 10 seconds with a temperature probe and a data logger. Plot a graph of temperature of stearic acid (after it completely melts) against time. Ask students to explain the shape of the cooling curve and identify the freezing point of the acid.

Support: Provide a step-by-step method with a titled table to write down results. The temperature against time graph could be plotted using Google sheets.

Challenge: Students can design their own method to carry out the experiment and plot data on a piece of graph paper.

2 Explaining the changes of states of matter in terms of the changes in the arrangement, movement and energy of the particles and the bonds in between the particles (20 minutes)

Resources: Demonstrations showing changes of states (for example, ice melting, water boiling, steam condensing, dry ice subliming)

Description and purpose: Ask students to apply the concept of intermolecular forces/bonds to explain the changes of states of water. Students will self-assess their explanations with key words.

Differentiation ideas

Support: Teachers can help students to review their answers and provide feedback on the use of keywords.

Stretch and challenge: Students can identify which processes are endothermic and which are exothermic.

Students can look into the different types of intermolecular forces and other types of bonding between particles.

Students can find out why sublimation occurs for some substances using the phase diagram (search on Chemguide.co.uk with the keywords 'phase diagram').

Language focus: Using scientific terminology and constructing logical long answers.

Plenary ideas

1 True or false (5 minutes)

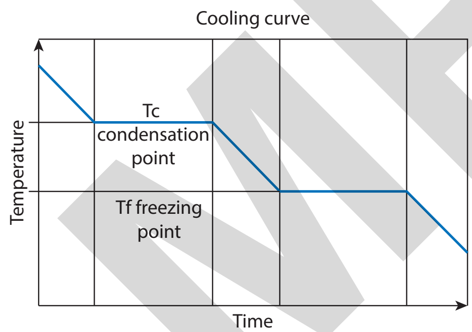
Resources: A table listing various properties of solids liquids and gases, for example (Worksheet question 4):

	True	False
solids can be compressed		
liquids have no fixed volume		
gases have no fixed volume		
gases have a high density		
solids have fixed shape		
liquids can diffuse		

Description and purpose: Students mark true/false in the table. This activity allows students to apply their knowledge of the particle models to draw conclusion on the macroscopic properties of different states of matter.

2 Labelling a cooling curve and explaining the stages of temperature change using the kinetic molecular model (10 minutes)

Resources: A cooling curve, for example, Figure 1.1.



Description and purpose: Students label the cooling curve with states of matter, changes of states, and identify the condensation and freezing points of the substance. Students also need to provide explanations on why temperature of the liquid goes down as cooling occurs but stays constant during freezing.

Assessment ideas

- Ask students to give examples of elements, compounds, mixtures, solids, liquids and gases around the classroom/lab.
- Suggest a suitable method for separating different types of mixtures.
- Label diagrams of lab apparatus and set-ups for filtration, evaporation, distillation and reflux.
- Students can design and carry out an experiment to obtain pure salt from rock salt.
- Calculations involving conversions between Celsius and kelvin temperature scales.
- Label the different stages of a melting and boiling curve.
- Test your understanding questions from the Coursebook.
- Define key words from the chapter.
- Explain the changes of states that occur during separation of mixtures. Ask students to use Post-it notes to assess their peers' answers.
- Give students explanations (containing common mistakes, missing out keywords) on the changes in the states of matter using the kinetic molecular theory and ask them to mark against a mark scheme.

Homework ideas

- Exam-style questions from the Coursebook, for example, questions 11–13.
- Exercises 1.1–1.4 from the Workbook.
- Carry out a paper chromatography experiment at home to separate the dyes in sweets. An example can be found by searching the 'Royal Society of Chemistry' website with the keywords 'chromatography of sweets'.
- Use Word Art to create an image for all the keywords in this chapter.
- Create flashcards on definitions of elements, compounds and mixtures and the different techniques used for separating components of mixtures.
- Sorting cards into solids, liquids and gases, or elements, compounds and mixtures. An example can be found by searching the 'Royal Society of Chemistry' website with the keywords 'lesson plans' and 'particle models'.

Links to digital resources

- Demonstrations on chemical changes (forming compounds from elements): Iron and sulphur – search the 'Royal Society of Chemistry' website with the keywords 'iron and sulfur reaction'
- Sodium and chlorine – search the Royal Society of Chemistry website with the keywords 'sodium and chlorine'
- Experiment on freezing stearic acid: search the 'Royal Society of Chemistry' website with the keywords 'freezing of stearic acid'
- Simulations on particle movements during changes of states: search on phet.colorado.edu for 'states of matter simulation'
- Home experiment to separate dyes in sweets using paper chromatography: search the 'Royal Society of Chemistry' website with the keywords 'chromatography of sweets'
- Revision notes on kinetic molecular theory and the states of matter (2016 syllabus): search on ibchem.com with the keywords 'kinetic molecular theory' and 'states of matter'
- Introduction to the phase diagrams: search on chemguide.co.uk with the keywords 'phase diagram'

CROSS-CURRICULAR LINKS

- Maths: Basic arithmetic calculations, plotting and interpreting graphs.
- Physics: Use and convert between kelvin and Celsius temperature scales. Molecular theory of solids, liquids and gases. Describe phase changes using particle behaviour.
- TOK: How does scientific knowledge progress?