

FOR THE
IB DIPLOMA

SECOND EDITION

Chemistry

Christopher Talbot,
Richard Harwood
and Christopher Coates



 **DYNAMIC**
LEARNING

 **HODDER**
EDUCATION

Chemistry

FOR THE IB DIPLOMA

SECOND EDITION

**Christopher Talbot,
Richard Harwood and
Christopher Coates**



**HODDER
EDUCATION**

AN HACHETTE UK COMPANY

All proprietary drug names and brand names in Chapters 22–25 are protected by their respective registered trademarks.

Although every effort has been made to ensure that website addresses are correct at time of going to press, Hodder Education cannot be held responsible for the content of any website mentioned in this book. It is sometimes possible to find a relocated web page by typing in the address of the home page for a website in the URL window of your browser.

Hachette UK's policy is to use papers that are natural, renewable and recyclable products and made from wood grown in sustainable forests. The logging and manufacturing processes are expected to conform to the environmental regulations of the country of origin.

Orders: please contact Bookpoint Ltd, 130 Milton Park, Abingdon, Oxon OX14 4SB. Telephone: (44) 01235 827720. Fax: (44) 01235 400454. Lines are open from 9.00 - 5.00, Monday to Saturday, with a 24 hour message answering service. You can also order through our website www.hoddereducation.com

© Christopher Talbot, Richard Harwood and Christopher Coates 2015

First edition published in 2010

This second edition published 2015

by Hodder Education

An Hachette UK Company

Carmelite House, 50 Victoria Embankment, London EC4Y 0DZ

Impression number 5 4 3 2 1

Year 2019 2018 2017 2016 2015

All rights reserved. Apart from any use permitted under UK copyright law, no part of this publication may be reproduced or transmitted in any form or by any means, electronic or mechanical, including photocopying and recording, or held within any information storage and retrieval system, without permission in writing from the publisher or under licence from the Copyright Licensing Agency Limited. Further details of such licences (for reprographic reproduction) may be obtained from the Copyright Licensing Agency Limited, Saffron House, 6–10 Kirby Street, London EC1N 8TS.

Cover photo © ESA/Herschel/PACS/MESS Key Programme Supernova Remnant Team; NASA, ESA and Allison Loll/Jeff Hester (Arizona State University)

Illustrations by Ken Vail Graphic Design and Aptara Inc.

Typeset in Goudy Oldstyle 10/12 pt by Aptara inc.

Printed in Slovenia

A catalogue record for this title is available from the British Library

ISBN: 978 1471 829055

Contents

Introduction	vii
Acknowledgements	ix
Core	
Chapter 1 Stoichiometric relationships	1
1.1 Introduction to the particulate nature of matter and chemical change	1
1.2 The mole concept	20
1.3 Reacting masses and volumes	31
Chapter 2 Atomic structure	52
2.1 The nuclear atom	52
2.2 Electron configuration	66
Chapter 3 Periodicity	85
3.1 Periodic table	85
3.2 Periodic trends	96
Chapter 4 Chemical bonding and structure	114
4.1 Ionic bonding and structure	114
4.2 Covalent bonding	125
4.3 Covalent structures	129
4.4 Intermolecular forces	144
4.5 Metallic bonding	158
Chapter 5 Energetics/thermochemistry	165
5.1 Measuring energy changes	165
5.2 Hess's Law	178
5.3 Bond enthalpies	187
Chapter 6 Chemical kinetics	199
6.1 Collision theory and rates of reaction	199
Chapter 7 Equilibrium	223
7.1 Equilibrium	223
Chapter 8 Acids and bases	250
8.1 Theories of acids and bases	250
8.2 Properties of acids and bases	256
8.3 The pH scale	261
8.4 Strong and weak acids and bases	265
8.5 Acid deposition	274

Chapter 9 Redox processes	283
9.1 Oxidation and reduction	283
9.2 Electrochemical cells	311
Chapter 10 Organic chemistry	322
10.1 Fundamentals of organic chemistry	322
10.2 Functional group chemistry	350
Chapter 11 Measurement and data processing	375
11.1 Uncertainties and errors in measurement and results	375
11.2 Graphical techniques	395
11.3 Spectroscopic identification of organic compounds	408
Additional higher level (AHL)	
Chapter 12 Atomic structure	435
12.1 Electrons in atoms	435
Chapter 13 The periodic table – the transition metals	451
13.1 First-row d-block elements	451
13.2 Coloured complexes	471
Chapter 14 Chemical bonding and structure	489
14.1 Further aspects of covalent bonding and structure	489
14.2 Hybridization	497
Chapter 15 Energetics/thermochemistry	522
15.1 Energy cycles	522
15.2 Entropy and spontaneity	535
Chapter 16 Chemical kinetics	552
16.1 Rate expression and reaction mechanism	552
16.2 Activation energy	575
Chapter 17 Equilibrium	585
17.1 The equilibrium law	588
Chapter 18 Acids and bases	606
18.1 Lewis acids and bases	606
18.2 Calculations involving acids and bases	612
18.3 pH curves	625
Chapter 19 Redox processes	643
19.1 Electrochemical cells	643
Chapter 20 Organic chemistry	671
20.1 Types of organic reactions	673
20.2 Synthetic routes	692
20.3 Stereoisomerism	699
Chapter 21 Measurement and analysis	719
21.1 Spectroscopic identification of organic compounds	719

Options

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

Option A

Chapter 22 Materials

- 22.1 Materials science introduction
- 22.2 Metals and inductively coupled plasma (ICP) spectroscopy
- 22.3 Catalysts
- 22.4 Liquid crystals
- 22.5 Polymers
- 22.6 Nanotechnology
- 22.7 Environmental impact – plastics
- 22.8 Superconducting metals and X-ray crystallography (AHL)
- 22.9 Condensation polymers (AHL)
- 22.10 Environmental impact – heavy metals (AHL)

Option B

Chapter 23 Biochemistry

- 23.1 Introduction to biochemistry
- 23.2 Proteins and enzymes
- 23.3 Lipids
- 23.4 Carbohydrates
- 23.5 Vitamins
- 23.6 Biochemistry and the environment
- 23.7 Proteins and enzymes (AHL)
- 23.8 Nucleic acids (AHL)
- 23.9 Biological pigments (AHL)
- 23.10 Stereochemistry in biomolecules (AHL)

Option C

Chapter 24 Energy

- 24.1 Energy sources
- 24.2 Fossil fuels
- 24.3 Nuclear fusion and fission
- 24.4 Solar energy
- 24.5 Environmental impact – global warming
- 24.6 Electrochemistry, rechargeable batteries and fuel cells (AHL)
- 24.7 Nuclear fusion and nuclear fission (AHL)
- 24.8 Photovoltaic and dye-sensitized solar cells (AHL)

Option D

Chapter 25 Medicinal chemistry

- 25.1 Pharmaceutical products and drug action
- 25.2 Aspirin and penicillin
- 25.3 Opiates
- 25.4 pH regulation of the stomach
- 25.5 Anti-viral medications
- 25.6 Environmental impact of some medications
- 25.7 Taxol – a chiral auxiliary case study (AHL)

- 25.8 Nuclear medicine (AHL)
- 25.9 Drug detection and analysis (AHL)

Index

737

Answers and glossary

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

Introduction

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

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

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

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

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

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

- Each chapter begins with **Essential Ideas** that summarize the concepts on which it is based.
- The text is written in **straightforward language**, without phrases or idioms that might confuse students for whom English is a second language. The text is also suitable for students of all abilities.
- The **depth of treatment** of topics has been carefully planned to accurately reflect the objectives of the IB syllabus and the requirements of the examinations.
- Photographs and full-colour illustrations support the relevant text, with annotations which elaborate on the context, function, language, history or applications of chemistry.
- The **Nature of Science** is an important new aspect of the IB Chemistry course, which aims to broaden students' interests and knowledge beyond the confines of its specific chemistry content. Throughout this book we hope that students will develop an appreciation of the processes and applications of chemistry and technology. Some aspects of the *Nature of Science* may be examined in IB Chemistry examinations and important discussion points are highlighted in the margins.
- The **Utilizations** and **Additional Perspectives** sections also reflect the Nature of Science, but they are designed to take students beyond the limits of the IB syllabus in a variety of ways. They may, for example, provide a historical context, extend theory or offer an interesting application. They are sometimes accompanied by more challenging, or research style, questions. They do not contain any knowledge which is essential for the IB examinations.
- Science and technology have developed over the centuries with contributions from scientists from all around the world. In the modern world science knows few boundaries and the flow of information is usually quick and easy. Some international applications of science have been indicated with the **International Mindedness** icon.
- **Worked examples** are provided in each chapter whenever new equations are introduced. A large number of **self-assessment questions** and some research questions are also placed throughout the chapters close to the relevant theory. They are phrased in order to assist comprehension and recall, and to help familiarize students with the assessment implications of the command terms.
- It is not an aim of this book to provide detailed information about experimental work or the use of computers. However, our **Applications and Skills** icon has been placed in the margin to indicate wherever such work may usefully aid understanding.
- A selection of **IB examination-style questions** are provided at the end of each chapter, as well as some past IB Chemistry examination questions. **Answers** to these are provided on the website accompanying this book.

Nature of Science



- Extensive links to the interdisciplinary **Theory of Knowledge (ToK)** element of the IB Diploma course, including ethics, are made in all chapters.
- Comprehensive **glossaries** of words and terms, including IB command terms, for Core and AHL topics are included in the website which accompanies this book.



- This icon denotes links to material available on the website that accompanies this book: www.hoddereducation.com/IBextras.

■ Using this book

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

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

■ Author profiles

Christopher Talbot

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

Richard Harwood

Richard was a Biochemistry researcher at Manchester Medical School and University College, Cardiff, before returning to teaching science in England and Switzerland. Most recently he has been involved in projects with various Ministries of Education evaluating science courses and providing teacher training nationally, and in individual schools, in Mongolia, Kazakhstan, Zimbabwe, India and Ghana.

Christopher Coates

Chris has previously taught in Suffolk, Yorkshire and Hong Kong at King George V School, and is currently Head of Science in the Senior School at the Tanglin Trust School, Singapore. He has taught A-level and IB Chemistry as well as ToK and MYP Science.

■ Authors' acknowledgements

We are indebted to the following lecturers who reviewed early drafts of the chapters for the second edition: Dr David L. Cooper, University of Liverpool (Chapters 2 and 14), Professor Mike Williamson, University of Sheffield (Chapters 21 and 23), Professor James Hanson, University of Sussex (Chapter 20), Professor Laurence Harwood, University of Reading (Chapter 20), Professor Robin Walsh, University of Reading (Chapters 6 and 16), Professor Howard Maskill, University of Newcastle (Chapter 20), Dr Norman Billingham, University of Sussex (Chapter 22), Dr Jon Nield, Queen Mary College, (Chapter 23), Professor Jon Cooper, University College London (Chapter 23), Dr Duncan Bruce, University of York (Chapter 22), Professor David Mankoff, University of Pennsylvania (Chapter 25), Dr Philip Walker, University of Surrey and Dr Eli Zysman-Colman (University of St Andrews (Chapter 22), and Dr Graham Patrick (Chapter 25), University of the West of Scotland.

I also acknowledge the contributions of Dr David Fairley (Overseas Family School, Singapore) who gave me invaluable advice and guidance on the many chemical issues I encountered when writing the book.

A special word of thanks must go to Mr Nick Lee, experienced chemistry and TOK teacher, workshop leader and IB examiner, for his most helpful comments on the final drafts.

Finally, we are indebted to the Hodder Education team that produced this book, led by Eleanor Miles and So-Shan Au at Hodder Education.

Chris Talbot
Singapore, June 2015

Acknowledgements

The Publishers would like to thank the following for permission to reproduce copyright material:

■ Photo credits

All photos by kind permission of Cesar Reyes except:

p.1 *t* Chris Talbot; **p.6** Science photo library/Michael W Davidson; **p.7** *t* Chris Talbot, *b* NASA/Johnson Space Center; **p.10** *t, b* Andrew Lambert Photography/Science Photo Library; **p.22** Chris Talbot; **p.26** Reproduced with permission of the BIPM, which retains full internationally protected copyright (photograph courtesy of the BIPM); **p.56** IBM Research; **p.63** Tim Beddow/Science Photo Library; **p.67** *l* Andrew Lambert Photography/Science Photo Library, *c* David Talbot, *r* Robert Balcer; **p.68** Carlos Santa Maria – Fotolia; **p.73** CERN; **p.87** *tl* Andrew Lambert Photography/Science Photo Library; **p.103** Prof. Mark J Winter/<http://www.webelments.com>; **p.108** *tr* Andrew Lambert Photography/Science Photo Library; **p.111** *b* JoLin/istockphoto.com; **p.122** Robert Balcer; **p.124** Se7enimage – Fotolia; **p.129** Chris Talbot; **p.141** *t* Chris Talbot; **p.143** Harry Kroto and used with the permission of The Sussex Fullerene Research Centre and photographer Nicholas Sinclair; **p.144** Public Domain/[Http://Commons.Wikimedia.Org/Wiki/File:Graphene-3D-Balls.Png](http://Commons.Wikimedia.Org/Wiki/File:Graphene-3D-Balls.Png); **p.162** Dirk Wiersma/Science Photo Library; **p.167** *t, b* David Talbot; **p.192** NASA/Goddard Space Flight Center; **p.199** *l* Roger Harris/Science Photo Library, *r* NOAA/Science Photo Library; **p.204** J C Revy /Science Photo Library; **p.205** *t* Dr Colin Baker; **p.223** *t* Anh Ngo – Fotolia, *b* Gigi200043 – Fotolia; **p.224** *t, b* Richard Harwood; **p.226** Richard Harwood; **p.233** Richard Harwood; **p.235** Science Photo Library; **p.237** Richard Harwood; **p.239** *t, b* Richard Harwood; **p.245** Bettmann/CORBIS; **p.255** Juan Gartner/Science Photo Library; **p.259** *l, r* Richard Harwood; **p.274** Leungchopan – Fotolia; **p.283** *b* David Talbot; **p.285** Phil Degginger/Alamy; **p.295** *b* sequence Chris Talbot; **p.304** *t, b* Chris Talbot; **p.305** *t* Chris Talbot; **p.306** Dr Colin Baker; **p.307** Andrew Lambert Photography/Science Photo Library; **p.309** Martyn F. Chillmaid/Science Photo Library; **p.317** Frank Scullion/http://www.franklychemistry.co.uk/electrolysis_lead_bromide_video.html; **p.322** Klaus Boller/Science Photo Library; **p.323** *t, b* Chris Talbot; **p.324** Richard Harwood; **p.328** *t* Mandritoiu – Fotolia, *b* David Talbot; **p.331** Richard Harwood; **p.332** Richard Harwood; **p.336** Rasmol Library/Richard Harwood; **p.339** Richard Harwood; **p.342** Chris Talbot; **p.346** *t* Geraint Lewis/Rex, *b* Richard Harwood; **p.348** *t* Richard Harwood, *b* IBM Research; **p.351** *t* Chris Talbot, *c* Full Image – Fotolia, *bl* Science Photo Library/David Taylor/Cordelia Molly, *br* David Taylor/Science Photo Library; **p.352** *t* Science Photo Library/Paul Rapson, *b* CSIRO/Science Photo Library; **p.353** *l* Eye Ubiquitous/Alamy, *r* Robert Brook/Science Photo Library; **p.354** Paul Rapson/Science Photo Library; **p.355** Chris Talbot; **p.356** David Talbot; **p.358** Chris Talbot; **p.359** Chris Talbot; **p.360** Andrew Lambert/Science Photo Library; **p.361** David Talbot; **p.365** *l* Roger Job/Science Photo Library, *r* Vanessa Vick/Science Photo Library; **p.366** Andrew Lambert/Science Photo Library; **p.368** *t* Chris Talbot, *b* Andrew Lambert/Science Photo Library; **p.370** Chris Talbot; **p.375** Ted Kinsman/Science Photo Library; **p.381** SciLabware; **p.395** JPL/NASA; **p.405** Chris Talbot; **p.408** Chris Talbot; **p.410** Dr Jon Hare; **p.423** Dr Jon Hare; **p.427** Mikhail Basov – Fotolia; **p.430** James Steidl/Fotolia.Com; **p.431** *t* Zephyr/Science Photo Library, *b* Dr Jon Hare; **p.441** CNRI/Science Photo Library; **p.458** Roger-Viollet/Topfoto; **p.460** Mark A. Wilson (Department Of Geology, The College Of Wooster)/Public Domain (<http://Commons.Wikimedia.Org/Wiki/File:Qtubironpillar.JPG>); **p.469** Chris Talbot; **p.474** Chris Talbot; **p.480** *t* Bruce Balick (University of Washington), Vincent Icke (Leiden University, The Netherlands), Garrelt Mellema (Stockholm University), and NASA/ESA, *c* Jose Ignacio Soto – Fotolia; **p.481** Andrew Lambert Photography/Science Photo Library; **p.516** Charles D. Winters/Science Photo Library; **p.526** Richard Harwood; **p.537** *t* David Talbot; **p.549** Public Domain/[Http://Schneider.Ncifcrf.Gov/Images/Boltzmann/Boltzmann-Tomb-3.Html](http://Schneider.Ncifcrf.Gov/Images/Boltzmann/Boltzmann-Tomb-3.Html); **p.585** *t* TUDGAY, Frederick, 1841–1921, The “Dunedin” off the English Coast, 1875, oil on canvas: 487 x 790 mm, accession: 02/01, Hocken Collections, Uare Taoka o Hākena, University of Otago, *b* Everett Collection/Rex; **p.586** Chris Talbot; **p.587** Treestreet – Fotolia; **p.597** Claude Nuridsany and Marie Perennou/Science Photo Library; **p.603** Sovereign, ISM/Science Photo Library; **p.628** Richard Harwood;

p.632 Chris Talbot; **p.652** CNRI/Science Photo Library; **p.659** Chris Talbot; **p.661** David Talbot; **p.667** David Talbot; **p.668** David Talbot; **p.671** Chris Talbot; **p.678** Andrew Lambert Photography/Science Photo Library; **p.684** Richard Harwood; **p.685** Andrew Lambert Photography/Science Photo Library; **p.688** Richard Harwood; **p.701** Richard Harwood; **p.706** David Talbot; **p.707** *t* Chris Talbot, *b* David Talbot; **p.708** David Talbot; **p.715** Richard Harwood; **p.727** Sovereign, ISM/Science Photo Library; **p.730** Chris Talbot; **p.733** Science Source/Science Photo Library; **p.735** Chris Talbot.

■ Artwork credits

p.25 Fig. 1.37 Jon Harwood; **p.37** Fig. 1.48 Kim Gyeoul; **p.54** Fig. 2.3 Kirstie Gannaway; **p.56** Fig. 2.10, **p.91** Fig. 3.16 Jon Harwood; **p.245** Fig. 7.26 Jon Harwood; **p.252** Fig. 8.1 Kim Gyeoul; **p.253** Fig. 8.2, **p.322** Fig. 10.2, **p.494** Fig. 14.17, **p.495** Fig. 14.20, **p.575** Fig. 16.21 Jon Harwood.

■ Examination questions credits

Examination questions have been reproduced with kind permission from the International Baccalaureate Organization.

Every effort has been made to trace all copyright holders, but if any have been inadvertently overlooked the Publishers will be pleased to make the necessary arrangements at the first opportunity.

1

Stoichiometric relationships

ESSENTIAL IDEAS

- Physical and chemical properties depend on the ways in which different atoms combine.
- The mole makes it possible to correlate the number of particles with the mass that can be measured.
- Mole ratios in chemical equations can be used to calculate reacting ratios by mass and gas volume.

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

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

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



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



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

■ States of matter

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

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

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

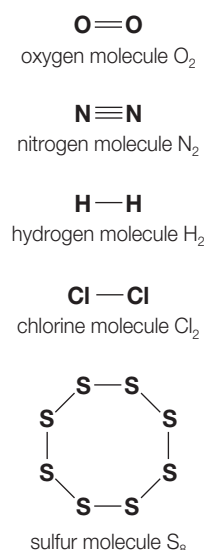
A gas (Figure 1.1), at a given temperature, has neither a definite shape nor a definite volume. It will take up the shape of any container and will spread out evenly within it, by a process known as diffusion. The volumes of gases are greatly affected by changes in temperature.

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

■ Elements

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

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



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

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

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

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

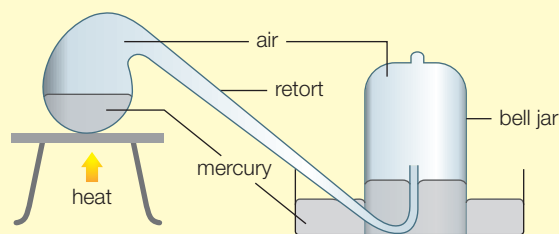
ToK Link

Priestley and Lavoisier's discovery of oxygen

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

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

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



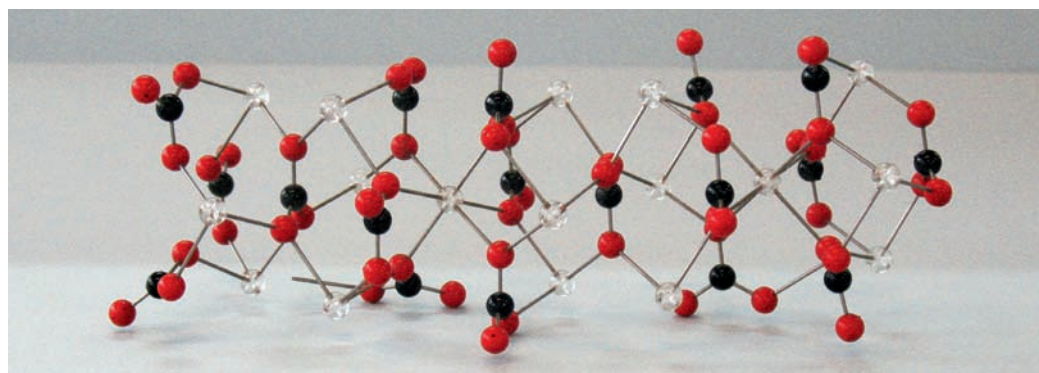
■ **Figure 1.4** Lavoisier's preparation of oxygen

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

■ Compounds

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

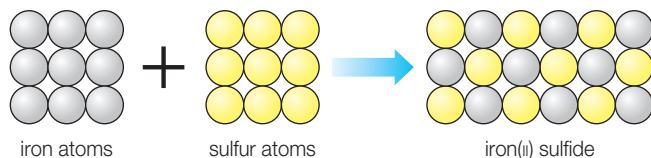
■ **Figure 1.5**
A model showing the structure of the compound calcium carbonate, CaCO_3 (black spheres represent carbon, red oxygen and clear calcium)



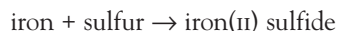
When a mixture of iron and sulfur is heated, large amounts of heat energy are released as the compound iron(II) sulfide, FeS, is formed (Figure 1.6). (Synthesis reactions like this are examples of redox reactions – see Chapter 9). Figure 1.6 describes this reaction in terms of atoms in iron and sulfur (Figure 1.7) reacting to form iron(II) sulfide.

■ **Figure 1.6**

A description of the formation of iron(II) sulfide in terms of atoms



The word equation for this reaction is:



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

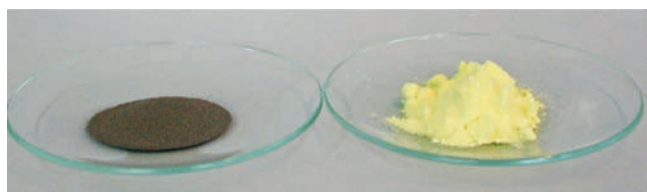
However, when a compound is formed the atoms it contains are chemically bonded together, so the compound will have different physical and chemical properties from the constituent elements (Table 1.1). For example, iron is magnetic, but the compound iron(II) sulfide is non-magnetic (Figure 1.8). A compound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

■ **Table 1.1**

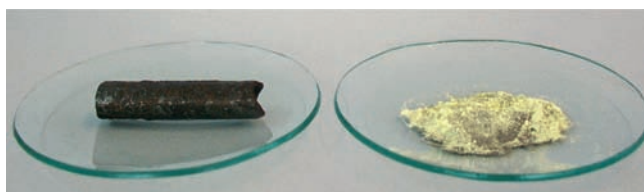
A summary of the different properties of iron, sulfur, an iron/sulfur mixture and iron(II) sulfide

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

The splitting of a chemical compound into its constituent elements is termed **decomposition**. This process requires an input of energy, either heat (**thermal decomposition**) or electricity (electrolysis) (Chapter 9).



■ **Figure 1.7** The elements iron and sulfur



■ **Figure 1.8** A sample of iron(II) sulfide and a mixture of iron and sulfur

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

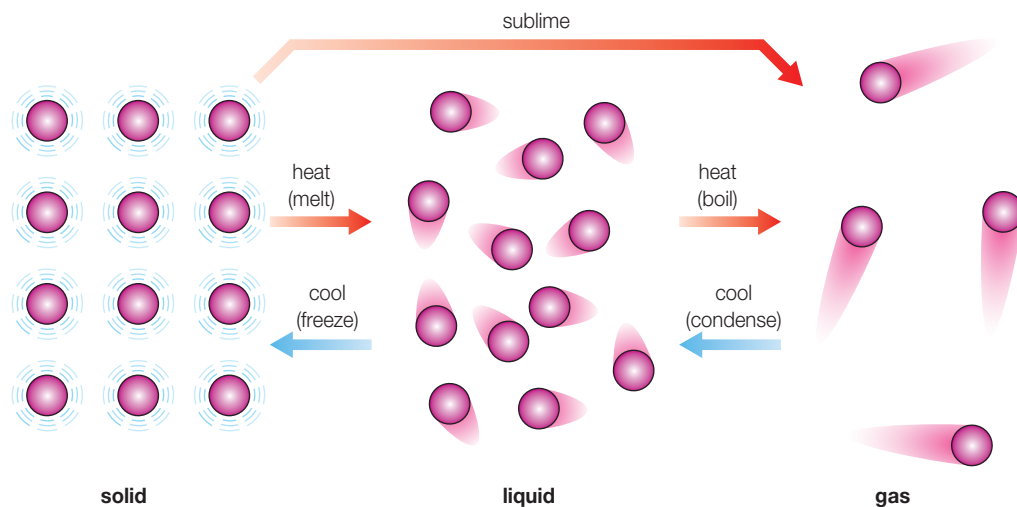
■ Molecular kinetic theory

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



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

Figure 1.9
The three states of matter and their interconversion



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

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



Changes of state

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

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

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

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

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

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

Substance and formula	Melting point/°C	Boiling point/°C
Aluminium, Al	661	2467
Ethanol, C ₂ H ₅ OH	-117	79
Mercury, Hg	-30	357
Oxygen, O ₂	-218	-183
Sodium chloride, NaCl	801	1413
Water, H ₂ O	0	100

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

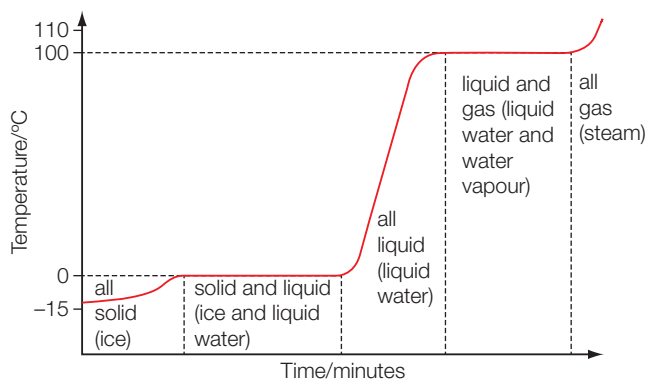
Eventually, a temperature is reached (the boiling point) at which the particles are trying to escape from the liquid so quickly that bubbles of gas form inside the bulk of the liquid. This is known as boiling. At the **boiling point** the pressure of the gas created above the liquid equals that in the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

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

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

The changes of state are physical changes: no new chemical substances are formed. Ice, water and steam all contain molecules with the formula H₂O. Whenever a change in state occurs the temperature remains constant during the change.

- 2 Identify the change of state which describes the following processes:
- Solid ethanol changing to liquid ethanol
 - Molten metal solidifying in a mould
 - Water changing to steam at 100°C
 - Bubbles of ethanol gas forming in liquid ethanol
 - Ice forming from water vapour on the freezer compartment of a fridge
 - Solid aluminium chloride forming a gas on gentle heating



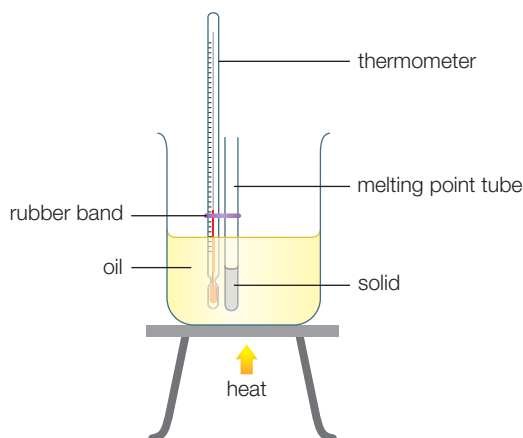
■ **Figure 1.10** Graph of temperature against time for the change from ice at -15°C to water to steam at 120°C

Heating and cooling curves

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

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

For ice to melt, the molecules must obtain sufficient kinetic energy to overcome the forces of attraction between the water particles to allow relative movement to take place. This is what the heat energy is doing. The temperature will begin to rise again only after all the ice has melted. Generally, the heating curve for a pure solid always stops rising at its melting point and gives rise to a sharp melting point. The addition or presence of impurities lowers the melting point. Figure 1.11 shows a simple apparatus used to find the melting point of a solid. Commercial melting point apparatus uses a heating block to melt the sample.



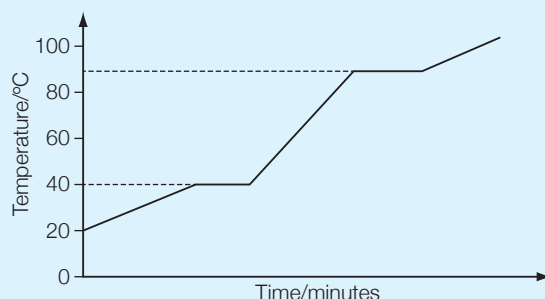
■ **Figure 1.11** Simple apparatus to find the melting point of a solid (in the melting point tube)

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

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

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

- Deduce the melting and boiling points of substance X.
- State the physical state of X at 25, 50 and 100°C.
- Explain what is happening during the melting and boiling of X.

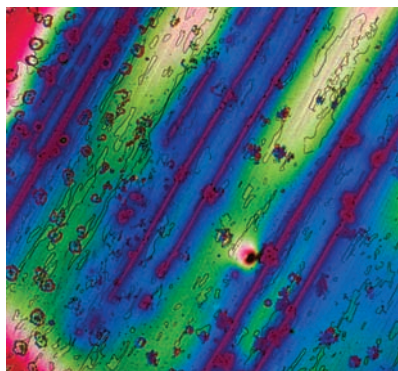


Utilization: Unusual states of matter

Liquid crystals



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



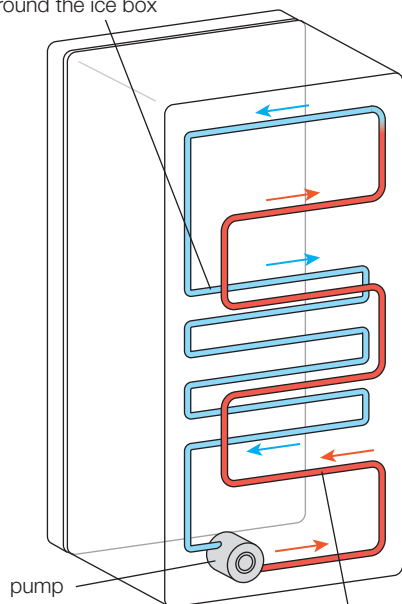
■ **Figure 1.12** A polarized light micrograph of liquid crystals

Plasma

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



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



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

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



■ **Figure 1.14** A domestic air conditioner



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

Utilization of heat changes during changes of state: Refrigeration

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

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

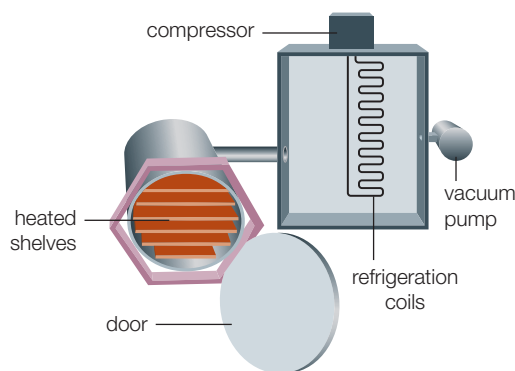
Continuing around the circuit, the vapour (gas) is compressed by the pump as it flows out at the bottom of the refrigerator. The compressed vapour is hot. As it flows through the pipes at the back of the refrigerator the vapour cools and condenses back to a liquid, releasing heat energy and heating up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

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

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

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

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



■ **Figure 1.16** Freeze-drying machine

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

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

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

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

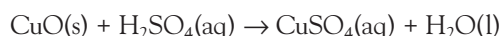
The heating units supply a small amount of heat to the shelves, causing the ice to change phase. Since the pressure is so low, the ice turns directly into water vapour. The water vapour flows out of the freeze-drying chamber, past the freezing coil. The water vapour condenses onto the freezing coil in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

Utilization: Atom economy in chemical reactions

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

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

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



Mass of starting atoms is

$$\text{CuO} = 63.5 + 16 = 79.5 \text{ g}$$

$$\text{H}_2\text{SO}_4 = 2 + 32 + 64 = 98 \text{ g}$$

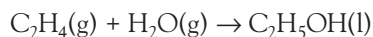
$$\text{Total} = 177.5 \text{ g}$$

Mass of desired product is

$$\text{CuSO}_4 = 63.5 + 32 + 64 = 159.5 \text{ g}$$

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

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



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

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

Green chemistry

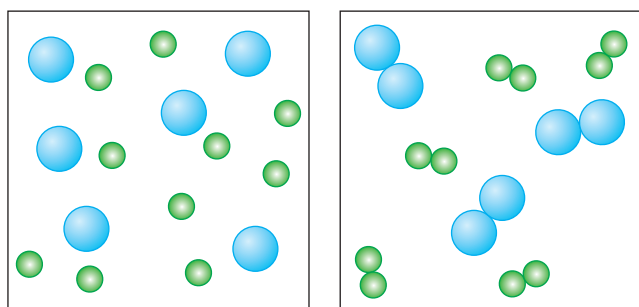
Green chemistry consists of chemicals and chemical processes designed to reduce or eliminate impacts on the environment. The use and production of these chemicals may involve reduced waste products, non-toxic chemicals, and improved efficiency. Industrial chemists evaluate chemical pathways and their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

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

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

Mixtures

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



a mixture of two elements existing as atoms

b mixture of two elements existing as molecules

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

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

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

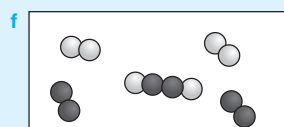
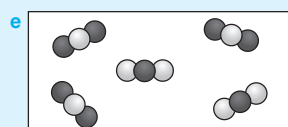
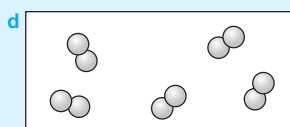
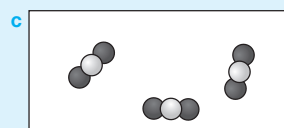
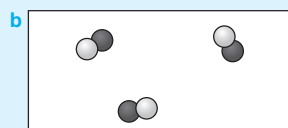
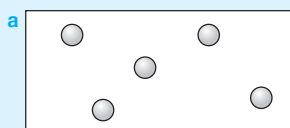
The major differences between mixtures of elements and compounds are summarized below in Table 1.3.

■ **Table 1.3**

The major differences between mixtures of elements and compounds

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

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



Types of mixtures

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