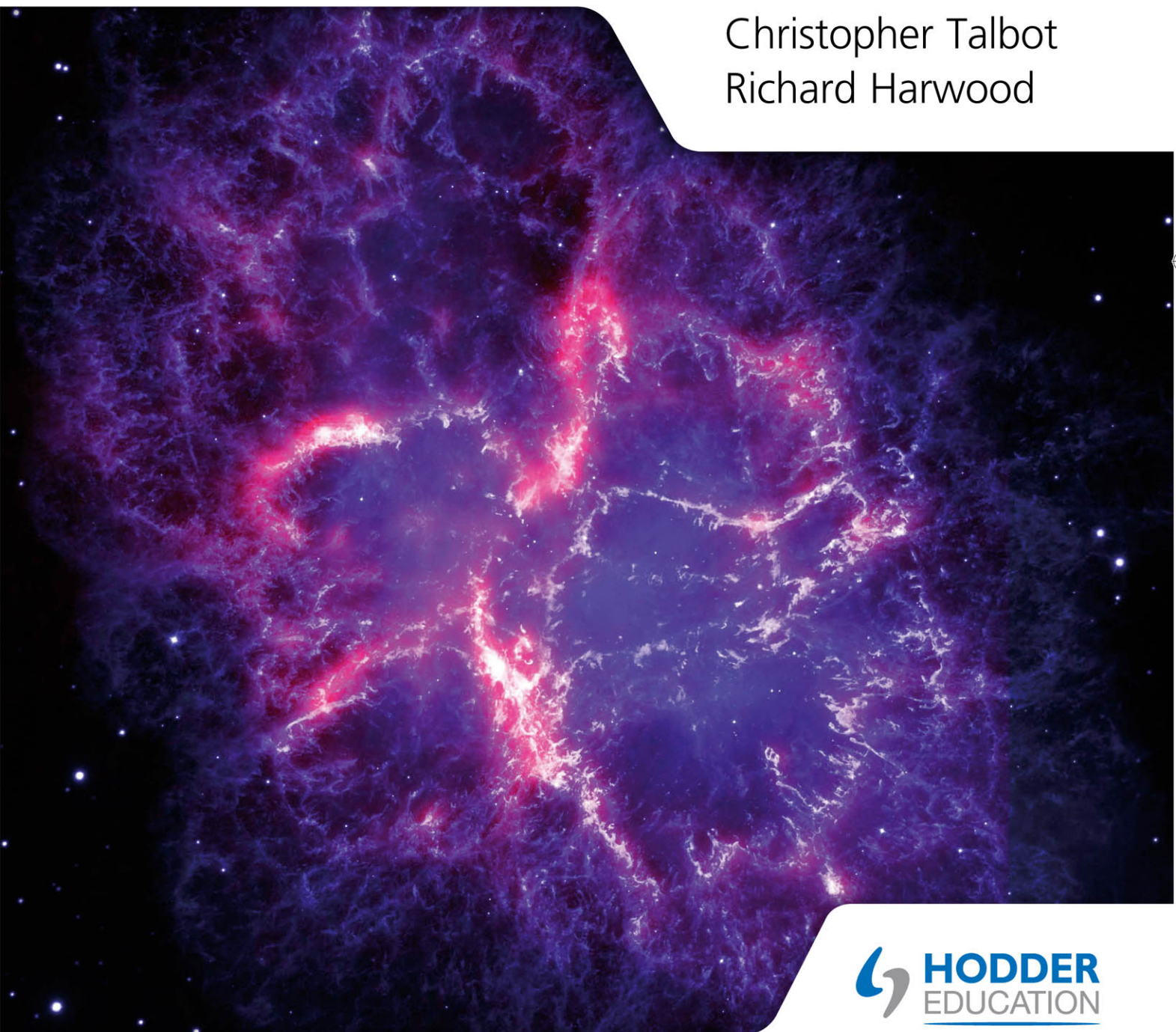


FOR THE  
IB DIPLOMA

# Chemistry

## Study and Revision Guide

Christopher Talbot  
Richard Harwood



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## Option chapters and answers

Option chapters and answers appear on the website accompanying this book:  
[www.hoddereducation.com/IBextras](http://www.hoddereducation.com/IBextras)

# How to use this revision guide

Welcome to the Chemistry for the IB Diploma Revision Guide!

This book will help you plan your revision and work through it in a methodological way. The guide follows the Chemistry syllabus topic by topic, with revision and practice questions to help you check your understanding.

## ■ Features to help you succeed

### Expert tip

These tips give advice that will help you boost your final grade.

### Common mistake

These identify typical mistakes that candidates make and explain how you can avoid them.

### Key definitions

The definitions of essential key terms are provided on the page where they appear. These are words that will help you have a clear understanding of important ideas. A comprehensive **glossary** of chemical terms should be downloaded from the Hodder Plus website: <https://www.hoddereducation.co.uk/ibextras/chemistryfortheib>

### ■ QUICK CHECK QUESTIONS

Use these questions provided throughout each section to make sure you have understood a topic. They are short knowledge-based questions that use information directly from the text.

### Worked example

Some parts of the course require you to carry out mathematical calculations, plot graphs, and so on: these examples show you how.

You can keep track of your revision by ticking off each topic heading in the book.

Tick each box when you have:

- revised and understood a topic
- tested yourself using the **Quick check questions**.

Online material can be found on the website accompanying this book [www.hoddereducation.com/IBextras/](http://www.hoddereducation.com/IBextras/)

Online material included:

- option chapters
- answers to Quick check questions
- glossary

Use this book as the cornerstone of your revision. Don't hesitate to write in it and personalize your notes. Use a highlighter to identify areas that need further work. You may find it helpful to add your own notes as you work through each topic. Good luck!

## Getting to know Papers 1, 2 and 3

At the end of your two year IB Chemistry course you will sit three papers – Papers 1, 2 and 3. Paper 1 is worth 20% of the final marks, Paper 2 is worth 36% of the final marks and Paper 3 is worth 20% of the final marks.

The other assessed part of the course (24%) is made up of the Internal Assessment (practical work), which is marked by your teacher and then moderated by the IBO.

Here is some general advice for the exams:

- Make sure you have learnt the command terms (e.g. evaluate, explain and outline): there is a tendency to focus on the content in the question rather than the command term, but if you do not address what the command term is asking of you then you will not be awarded marks.
- Answer all questions and do not leave gaps.
- Do not write outside the answer boxes provided – if you do so this work will not be marked.
- If you run out of room on the page, use continuation sheets and indicate clearly that you have done this on the cover sheet. The fact that the question continues on another sheet of paper needs to be clearly indicated in the text box provided.
- Plan your time carefully before the exams – this is especially important for Papers 2 and 3.

## 1.1 Introduction to the particulate nature of matter and chemical change

Revised

**Essential idea:** Physical and chemical properties depend on the ways in which different atoms combine.

### Elements and compounds

Revised

- Atoms of different elements combine in fixed ratios to form compounds, which have different properties from their component elements.

### Elements

Revised

There are 92 naturally occurring **elements**. An element consists of one type of particle. These particles may be atoms or molecules (atoms of the same type bonded together covalently). Most elements consist of single atoms, and their formulas are simply the symbol for the element. However, some elements exist as diatomic molecules, having the formula  $X_2$ . Seven elements behave like this under normal conditions; hydrogen, oxygen, nitrogen and the halogens (group 17 – fluorine, chlorine, bromine and iodine).

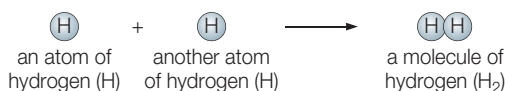
#### Key definition

**Element** – a substance that cannot be broken down by any chemical reaction into simpler substances.

One possible way to remember them is by learning the mnemonic involving their symbols:

**I** **H**ave **N**o **B**right **O**r **C**lever **F**riends!

Atoms of different elements have different sizes and different masses (Figure 1.1). Elements can be classified into metals, non-metals or metalloids.

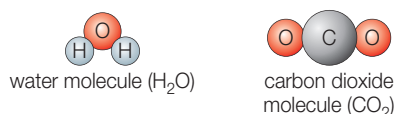


**Figure 1.1** Representation of atoms

### Compounds

Revised

Some pure substances are made up of a single element, although there may be more than one atom of the element in a particle of the substance. For example, as we have seen above, hydrogen is diatomic: molecules of hydrogen contain two hydrogen atoms and have the formula  $H_2$  (H–H).



**Figure 1.2** Representation of the molecular compounds water and carbon dioxide

**Compounds** are made up of two or more different atoms (or ions) that have chemically bonded in fixed ratios; but not always in molecules – we cannot use the term molecule when considering ionic compounds. For example, molecules of carbon dioxide ( $\text{CO}_2$ ) contain one carbon atom and two oxygen atoms. However, zinc iodide is also a compound and contains zinc and iodide ions. It has the chemical formula,  $\text{ZnI}_2$  [ $\text{Zn}^{2+} 2\text{I}^-$ ].

Zinc iodide is a compound, and can be broken down chemically into its constituent elements: iodine and zinc. This can be carried out by heating (thermal decomposition) or electrolysis of molten zinc iodide (electrolytic decomposition).

## ■ Laws of chemical combination

The law of conservation of mass states that during any physical or chemical change, the total mass of the products remains equal to the total mass of the reactants. The law of definite proportions states that a compound always contains the same elements combined together in the same proportion by mass. The law of multiple proportions states that when two elements combine with each other to form one or more compounds, the masses of one of the elements which combine with a fixed mass of the other, are in an integer ratio to one another.

### NATURE OF SCIENCE

Making quantitative measurements with replicates to ensure reliability – definite and multiple proportions. The various laws of chemical combination were formulated after repeated accurate measurements of mass and gas volumes were performed on a variety of chemical reactions.

### Key definition

**Compound** – a substance that contains more than one element combined chemically in a fixed ratio (Figure 1.2).

### Expert tip

Compounds usually have very different chemical and physical properties from their component elements.

## Mixtures

Revised

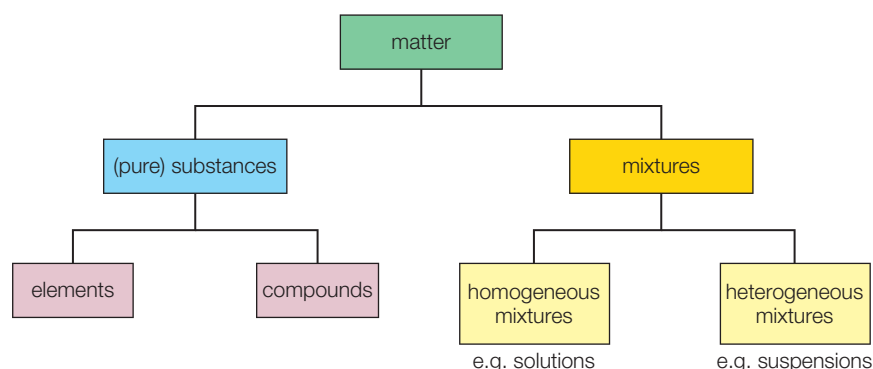
- **Mixtures** contain more than one element and/or compound that are not chemically bonded together and so retain their individual properties.
- Mixtures are either homogeneous or heterogeneous.

The components of a mixture (Figure 1.3) may be elements or compounds. These components are not chemically bonded together and therefore the components of a mixture keep their individual properties.

If all the components of a mixture are in the same phase the mixture is *homogeneous*. In a homogeneous mixture, the components remain mixed with each other and the composition is uniform. If the components of a mixture are in different phases the mixture is *heterogeneous*. There is a phase boundary between two phases.

### Key definition

**Mixture** – consisting of two or more substances (elements and/or compounds) mixed together.



**Figure 1.3** Classification of matter

Homogeneous mixtures can be separated by certain physical methods, e.g. distillation/crystallization for a sodium chloride solution, fractional distillation for liquid air, petroleum, or an ethanol/water mixture.

Heterogeneous mixtures can be separated by a different set of physical (mechanical) methods, e.g. filtration for a suspension of fine sand in water.

### ■ QUICK CHECK QUESTIONS

- 1 Classify the following substances into elements, compounds and mixtures. Indicate whether the mixture is homogeneous or heterogeneous.  
Smoke, mercury, bronze, sucrose, dry ice, carbonated (fizzy) drink, magnesium iodide in solution in distilled water, air, magnesium powder with sulfur powder, petrol and water, and chalk and water.
- 2 When two substances **X** and **Y** are mixed together **Z** is formed with the evolution of heat. The properties of **Z** are different from those of **X** and **Y**. Is **Z** an element, compound or a mixture?

## States of matter and changes of state

Revised

One of the most fundamental ideas in chemistry is that all matter consists of particles since we use this idea to explain many of the chemical and physical behaviours that are observed (the *kinetic model* of matter).

Figure 1.4 shows the interconversions between the three states of matter and the arrangement of their particles (ions, atoms or molecules). These changes of state are physical changes and occur at constant pressure. Table 1.1 summarizes the properties of the three states of matter and their description according to kinetic molecular theory.

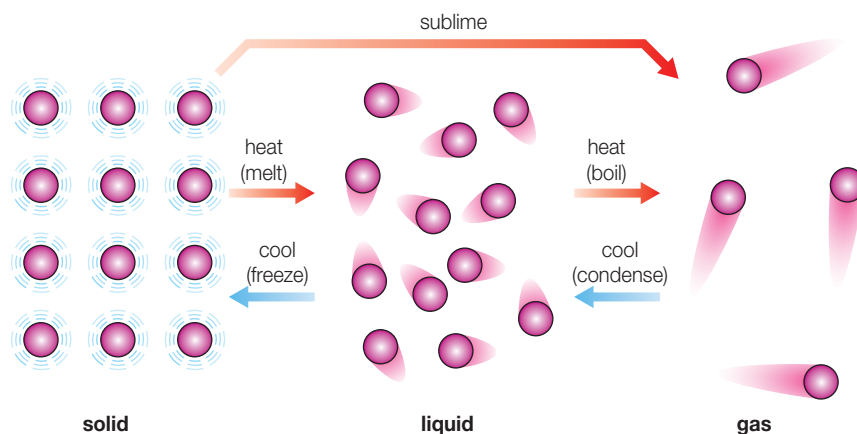


Figure 1.4 The three states of matter and their interconversion

Table 1.1 Properties of the three states of matter and their description by kinetic molecular theory

Solids	Liquids	Gases
<ul style="list-style-type: none"> <li>• Particles are close together.</li> <li>• Particles have lower kinetic energy than in the other two states.</li> <li>• Particles can only vibrate about fixed positions.</li> <li>• There are strong forces of attraction between particles.</li> </ul>	<ul style="list-style-type: none"> <li>• Particles are usually slightly further apart than in the solid.</li> <li>• Particles have greater values of kinetic energy than those in the solid state.</li> <li>• Particles can move (translation) so diffusion may occur.</li> </ul>	<ul style="list-style-type: none"> <li>• Particles are far apart from one another.</li> <li>• Particles have much more kinetic energy than the other two states.</li> <li>• Particles move at high speeds in straight lines between collisions.</li> <li>• There is very little attraction between the particles (no intermolecular forces operating between the particles of an ideal gas).</li> </ul>
Properties		
Fixed shape	Variable shape	Variable shape (occupies whole container due to diffusion)
Fixed volume	Fixed volume	Variable volume (affected by temperature and pressure)
Incompressible	Incompressible	Highly compressible
High density	Intermediate density	Very low density



Substances can change their state of matter as a result of temperature changes (or changes in pressure). Table 1.2 shows the changes in energy and movement that occur during a change in state (the shaded cells) as a result of changes in temperature. Remember that the temperature of a substance is directly related to the average kinetic of all its particles.

When the temperature of a crystalline solid is raised, the particles vibrate increasingly vigorously until they can no longer be held in an ordered arrangement (lattice) by the forces of attraction and the solid melts. Raising the temperature of a liquid increases the average speed of the particles until their kinetic energy is sufficient to overcome the forces of attraction between them: the liquids boils. Evaporation is the conversion of the liquid to the gaseous state, below the boiling point. It occurs at the surface of a liquid, but boiling involves the entire liquid and bubble formation.

**Table 1.2** The physical changes of state taking place as a substance is heated (left hand column) or cooled down (right hand column). These transitions are often represented by a heating or cooling curve.

Cooling (exothermic)			
Gas	Particles gain kinetic energy and vibrate, rotate and move faster, move randomly, and move further apart (gas expands). <i>Temperature of the gas rises.</i>	Particles lose kinetic energy and vibrate, rotate and translate less, come closer together (gas contracts). <i>The temperature decreases to condensation point.</i>	Gas
Evaporation, boiling	Particles use the thermal energy supplied to overcome attractive forces between particles and escape the liquid. <i>Temperature remains the same.</i>	Particles come closer together and release energy (latent heat). <i>Temperature remains the same.</i>	Condensation
Liquid	Particles gain kinetic energy and vibrate, rotate and move faster. <i>The temperature rises to boiling point.</i>	Particles lose kinetic energy and vibrate, rotate and translate less. <i>The temperature decreases to freezing point.</i>	Liquid
Melting	Particles use the thermal energy supplied to overcome attractive forces between them. <i>Temperature remains the same.</i>	Particles lose energy (latent heat) as attractive forces are formed and the particles return to the lattice or regular arrangement. <i>Temperature remains the same.</i>	Freezing
Solid	Particles gain kinetic energy and vibrate more and move further apart. Substance expands but maintains its shape. <i>Temperature increases to melting point.</i>	Particles lose kinetic energy – <i>temperature decreases</i> – and its particles vibrate less and move closer together.	Solid
Heating (endothermic)			

On cooling the reverse changes occur. The particles of the gas gradually slow down as the temperature falls until the forces of attraction are able to condense them together and form a liquid. Cooling the liquid causes further loss of kinetic energy until eventually the particles form a crystalline solid where they vibrate about fixed points.

#### Expert tip

Sublimation is the conversion of a solid to gas at constant temperature; vapour deposition is the reverse process.

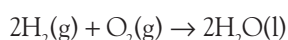
## Chemical equations

Revised

During a chemical reaction new substances are formed. There is an enthalpy change between the reacting chemicals (system) and its surroundings: heat will be released or absorbed. This is due to bonds in the reactants being broken and bonds in the products being formed. There is a fixed relationship between the masses of the *reactants* and *products*, known as the stoichiometry of the reaction.

Chemical reactions can be described by balanced equations showing the formulas of the substances and their physical states (under standard conditions). Equations need to be balanced with (stoichiometric) *coefficients* because atoms cannot be created or destroyed during a chemical reaction.

Consider the following balanced equation:



*Qualitatively*, an equation states the names of the reactants and products (and gives their physical states using state symbols).

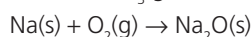
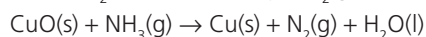
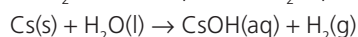
*Quantitatively*, it expresses the following relationships:

- the *relative number of molecules* of the reactants and products – here 2 molecules of hydrogen react with 1 molecule of oxygen to form 2 molecules of water
- the (*amounts*) in moles of the particles that form the reactants and products – 2 moles of hydrogen molecules react with 1 mole of oxygen molecules to form 2 moles of water molecules (one mole is  $6.02 \times 10^{23}$  particles)
- the *relative masses* of reactants and products – 4.04 g of molecular hydrogen reacts with 32.00 g of molecular oxygen to form 36.04 g of water molecules; the law of conservation of mass is obeyed
- the *relative volumes* of gaseous reactants and products (Avogadro's law) – 2 volumes of hydrogen react with 1 volume of oxygen to form liquid water. For example, 2 dm<sup>3</sup> of hydrogen would react with 1 dm<sup>3</sup> of oxygen.

Simple equations are balanced by a process of 'trial and error' (inspection). However, it is helpful to balance the formula which contains the maximum number of elements first.

### ■ QUICK CHECK QUESTION

- 3 Balance the following equations.



## ■ Ionic equations

An ionic equation may be written when a reaction occurs in solution in which some of the ions originally present are removed from solution, or when ions not originally present are formed. Usually, ions are removed from solution by one or more of the following processes:

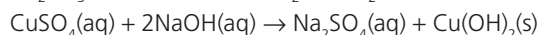
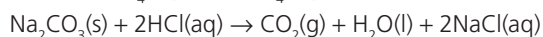
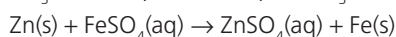
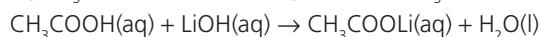
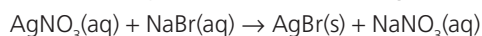
- formation of an insoluble ionic compound (a precipitate)
- formation of molecules containing only covalent bonds (water, for instance)
- formation of new ionic species and
- formation of a gas (e.g. carbon dioxide).

To deduce an ionic equation follow these steps:

- 1 Write down the balanced molecular equation.
- 2 Convert those chemicals that are ions in solution into their ions.
- 3 Cross out spectator ions that appear on both sides of the equation in solution as they have not changed during the reaction.
- 4 Check that the atoms and charges balance in the ionic equation.

### ■ QUICK CHECK QUESTION

- 4 Write ionic equations for the following reactions.

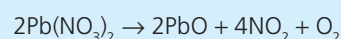


### Common mistake

It should also be noted that some reactions do *not* occur, even though balanced equations can be written, for example,  $\text{Cu}(\text{s}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{CuSO}_4(\text{aq}) + \text{H}_2(\text{g})$ . Hence, the activity series (Topic 9, Redox processes) should be consulted before equations for replacement reactions are written.

### Expert tip

If you balance an equation in a multiple choice question and deduce the sum of the coefficients, do not forget the 1's in your calculation. So for the equation



the answer is **9**. Note that the *stoichiometric coefficients in a balanced equation are whole numbers*.

### Expert tip

It is good practice to include state symbols in equations (and equilibrium expressions) and they are essential in thermochemical equations.

### Expert tip

An ionic equation should be balanced with respect to atoms (elements) and charge. The total charges on both sides of the ionic equation should be the same.

### Expert tip

Insoluble salts include halides of lead(II), barium sulfate, most silver salts (except silver nitrate) and most carbonates (except group 1). For straightforward precipitation reactions the simplest way to construct the ionic equation is to write down the precipitate on the product side, then fill in the ions that would make that precipitate as the reactants.

## Types of reaction

A direct synthesis reaction involves the combination of two or more elements to produce a single product. Single replacement (displacement) reactions occur when one element replaces another in a compound. These types of reaction are both redox reactions as they involve a transfer of electrons.

Decomposition reactions involve a single reactant being broken down into two or more products. Precipitation (or double displacement) reactions occur between ions in aqueous solution to form an insoluble precipitate. Acid–base (Brønsted–Lowry) reactions involve a transfer of protons ( $\text{H}^+$ ).

### QUICK CHECK QUESTION

5 Write ionic equations for the following reactions and state the type of reaction.

calcium reacting with dilute nitric acid

barium chloride solution reacting with sodium sulfate solution

copper(II) oxide reacting with hydrochloric acid

### Expert tip

Pure water is a liquid and represented as  $\text{H}_2\text{O}(\text{l})$  and not an aqueous solution,  $\text{H}_2\text{O}(\text{aq})$ .

## 1.2 The mole concept

Revised

**Essential idea:** The mole makes it possible to correlate the number of particles with the mass that can be measured.

## Mole concept and the Avogadro constant

Revised

The mole is a fixed number of particles and refers to the amount,  $n$ , of a substance.

As particles such as atoms, ions and molecules are extremely small, chemists measure amounts of substance using a fixed amount called the **mole**. This fixed amount contains a certain number of particles referred to as the **Avogadro constant**, which has a value of  $6.02 \times 10^{23} \text{ mol}^{-1}$ .

The mass of one mole of a substance has units of gram per mole ( $\text{g mol}^{-1}$ ). Multiplying an amount of substance (in number of moles) by Avogadro's constant converts it to the number of particles present. Dividing the number of particles present by the Avogadro constant gives the amount (in mol) (Table 1.3).

**Table 1.3** Amounts of selected substances and the number of particles in multiples of Avogadro's constant

Amount of substance/mol	Mass/g	Number of particles (in terms of $N_A$ or $L$ )
1 mole of aluminium, Al	26.98	$N_A$ or $L$ atoms
1 mole of oxygen molecules, $\text{O}_2$	32.00	$2N_A$ or $2L$ atoms or $N_A$ or $L$ molecules
1 mole of ethene, $\text{C}_2\text{H}_4$	26.04	$6N_A$ or $6L$ atoms or $N_A$ or $L$ molecules
1 mole of calcium nitrate, $\text{Ca}(\text{NO}_3)_2$	164.10	$N_A$ or $L$ formula units of $\text{Ca}(\text{NO}_3)_2$ $N_A$ or $L$ ions of $\text{Ca}^{2+}$ , $2N_A$ or $2L$ ions of $\text{NO}_3^-$ Total: $3N_A$ or $3L$ ions

### Key definitions

**Mole** – that amount ( $n$ ) of substance that contains the same number of particles (atoms, ions, molecules or electrons) as there are atoms in exactly 12 g of isotope carbon-12.

**Avogadro constant** – the number of particles in a mole; found (experimentally) to be  $6.02 \times 10^{23}$  (to 3 s.f.) and has the symbol  $L$  or  $N_A$ .

## Molar mass has the units of $\text{g mol}^{-1}$

Revised

$M$  refers to the *molar mass*. It is the mass of substance in grams that contains one mole of particles (Figure 1.5). It has units of gram per moles,  $\text{g mol}^{-1}$ , and the same numerical value as the relative atomic mass,  $A_r$ , or relative formula mass,  $M_r$ . The molar mass of a compound is the sum of the molar masses of all the atoms in the molecule or formula unit.



**Figure 1.5** The mole concept for selected elements

In each pile there is one mole of atoms ( $6.02 \times 10^{23}$  atoms). The relative atomic masses of carbon, sodium and magnesium are 12.01, 22.99 and 24.31. The molar masses are  $12.01 \text{ g mol}^{-1}$ ,  $22.99 \text{ g mol}^{-1}$  and  $24.31 \text{ g mol}^{-1}$ .

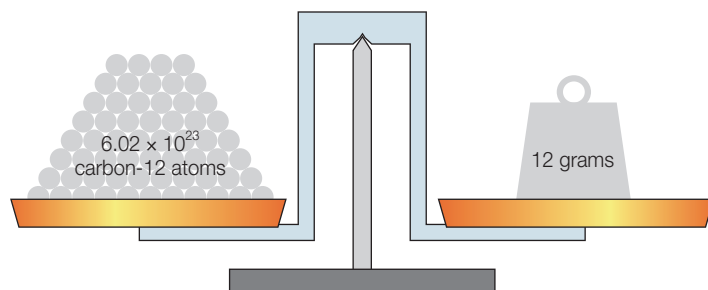
## Masses of atoms

Revised

- Masses of atoms are compared on a scale relative to  $^{12}\text{C}$  and are expressed as relative atomic mass ( $A_r$ ) and relative formula/molecular mass ( $M_r$ ).

**Relative atomic mass** and **relative formula (or molecular) mass** are pure numbers without units. Do not write them with units of grams. Molar mass has units of gram per mole, not gram.

The term relative molecular mass only really applies to simple molecules (a group of covalently bonded atoms); the concept of relative formula mass is used for giant structures like ionic substances (e.g. sodium chloride, copper(II) sulfate) and macromolecules (e.g. silicon dioxide).



**Figure 1.6** Illustration of the Avogadro constant

To calculate the mass of one molecule (or one ion or atom) in grams, calculate the molar mass of the substance and divide by Avogadro's constant.

### Key definitions

**Relative atomic mass,  $A_r$**  (Figure 1.6) – the weighted average (according to their percentage abundances) of the atomic masses of its isotopes compared to  $1/12$ th of (the mass of) of  $^{12}\text{C}$ .

**Relative formula (molecular) mass,  $M_r$**  – the sum of the relative atomic masses of the elements present in the formula unit or molecule.

### Expert tip

Remember that the mass (in grams or kilograms) of a single atom or molecule is going to be extremely small.

### QUICK CHECK QUESTIONS

- Deduce molar masses of the following species: uranium atom, U, sulfur hexafluoride molecule,  $\text{SF}_6$ , lead(IV) ion,  $\text{Pb}^{4+}$ , and ammonium sulfate,  $(\text{NH}_4)_2\text{SO}_4$  [ $2\text{NH}_4^+ \text{SO}_4^{2-}$ ].
- Calculate the number of atoms in 0.5 moles of atoms of nitrogen (N) and 0.2 moles of molecules of nitrogen ( $\text{N}_2$ ).
- Calculate the mass (in milligram) of 0.020 moles of calcium sulfate,  $\text{CaSO}_4$ .
- Calculate the mass (in gram) of one molecule of nitrogen dioxide,  $\text{NO}_2$ .
- Calculate the amount, mass and number of chloride ions in 0.400 mol of magnesium chloride,  $\text{MgCl}_2$ .
- 0.10 mol of a substance has a mass of 4.00 g. Calculate its molar mass.
- One argon atom has a mass of  $6.64 \times 10^{-26} \text{ kg}$ . Calculate the relative atomic mass of argon.

### Common mistake

Relative atomic and molecular masses are not referenced to hydrogen atoms,  $^1\text{H}$ , the lightest atoms. Carbon-12 has been chosen (for practical reasons) as the reference nuclide for comparing the relative masses of atoms and molecules. Remember too that all *relative* masses do not have units, they are simply a number.

## ■ Interconversion between amount and other properties

The amount of a pure substance (solid, liquid or gas) can be found from the mass of the substance and the molar mass:

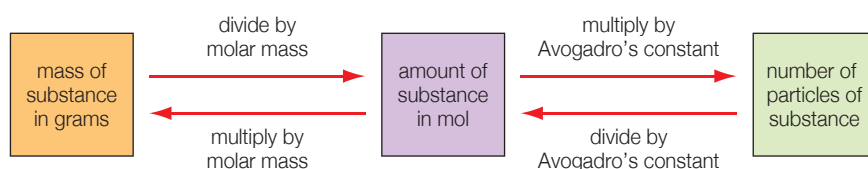
$$\text{amount (mol)} = \frac{\text{mass of substance (g)}}{\text{molar mass (g mol}^{-1}\text{)}}$$

In the laboratory substances are most conveniently measured by weighing for solids and by volume for liquids and gases. However, gases and liquids can both be weighed and masses and volumes can be converted via density ( $\text{density} = \frac{\text{mass}}{\text{volume}}$ ).

## Solution of problems involving the relationships between the number of particles, the amount of substance in moles and the mass in grams

Revised

The relationships between the amount of substance, number of particles (atoms, ions and molecules), mass of pure solid, liquid or gas (and volume of pure gas) are very important (Figure 1.7). Many stoichiometry calculations involve converting from one part of this relationship to another.



**Figure 1.7** Summary of interconversions between amount, mass and number of particles

### ■ QUICK CHECK QUESTIONS

- 13** Calculate the mass of carbon dioxide produced from heating 5.50 g of sodium hydrogen carbonate which on heating will completely decompose and form sodium carbonate, water and carbon dioxide.
- 14** Calculate the mass of titanium(IV) oxide formed when 10.00 g of titanium(IV) chloride reacts with water:  

$$\text{TiCl}_4(\text{s}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow \text{TiO}_2(\text{s}) + 4\text{HCl}(\text{aq})$$
- 15** Determine the mass of fluorine ( $\text{F}_2$ ) required to produce 150.00 g of  $\text{C}_4\text{F}_{10}$ .  

$$\text{C}_4\text{H}_{10}(\text{g}) + 10\text{F}_2(\text{g}) \rightarrow \text{C}_4\text{F}_{10}(\text{g}) + 10\text{HF}(\text{g})$$

### Expert tip

The term amount in chemistry means moles (units of mol). It does not refer to mass or volume.

## Formulas of compounds

Revised

- The empirical formula and molecular formula of a compound give the simplest ratio and the actual number of atoms present in a molecule respectively.

### ■ Empirical formulas

#### ■ Interconversion of the percentage composition by mass and the empirical formula

An empirical formula (obtained by experiments) shows the simplest whole number ratio of atoms of each element. It can be deduced from the percentage composition by mass of the compound. The percentage composition values can be converted directly into mass by assuming 100 g of the pure compound. The amounts of each atom can then be calculated using the molar masses. The numbers may need to be multiplied to obtain integer values.



## ■ Obtaining and using experimental data for deriving empirical formulas from reactions involving mass changes

The empirical formula can also be calculated when given mass data. The mass of a missing element can be calculated by difference. The masses of each element are converted to amounts (mol) via the molar masses. The masses are divided by the smallest value. The numbers may need to be multiplied to obtain integer values.

## ■ Molecular formulas

### ■ Determination of the molecular formula of a compound from its empirical formula and molar mass

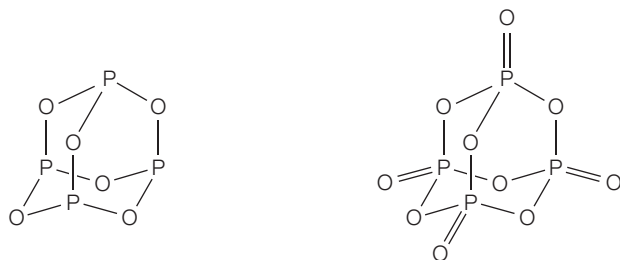
The molecular formula shows the actual number of atoms of each element in a molecule of the substance. It can be obtained from the empirical formula if the molar mass of the compound is known. The empirical and molecular formulas are connected by the following relationship:

$$\frac{\text{molecular formula}}{\text{empirical formula}} = n \text{ (where } n = 1, 2, 3, \text{ etc.)}$$

The empirical formula of organic compounds can be obtained from combustion data. The masses of carbon and hydrogen can be obtained from the masses of water and carbon dioxide (via their percentage by masses of oxygen and carbon).

## ■ Structural formulas

Structural formulas show the arrangement of atoms and bonds within a molecule and are used in organic chemistry. Figure 1.8 shows the structural formulas of phosphorus(III) and phosphorus(V) oxides. The molecular formulas are  $\text{P}_4\text{O}_6$  and  $\text{P}_4\text{O}_{10}$ , and the empirical formulas are  $\text{P}_2\text{O}_3$  and  $\text{P}_2\text{O}_5$ , respectively.



**Figure 1.8** The structures of phosphorus(III) oxide and phosphorus(V) oxide

### Expert tip

The empirical formula of a compound often may not correspond to a stable species. For example, the empirical formula of  $\text{H}_2\text{O}_2$  is  $\text{OH}$ , which does not exist as a stable species.

### ■ QUICK CHECK QUESTIONS

- 16** Deduce the empirical formulas of the molecular formulas of the following compounds:  
 $\text{C}_6\text{H}_6$ ,  $\text{C}_6\text{H}_{12}$ ,  $\text{PH}_3$ ,  $\text{B}_2\text{H}_6$ ,  $\text{N}_2\text{O}_4$ ,  $\text{Fe}_2\text{O}_3$ ,  $\text{C}_2\text{H}_2$
- 17** The composition by mass of a compound is 40.2% potassium, 32.9% oxygen and 26.9% chromium. Determine the empirical formula of the compound.
- 18** When 2.67 g of copper reacts with excess sulfur (in the form of atoms), the mass of the compound formed is 4.01 g. Determine the empirical formula of the compound.
- 19** A compound has an empirical formula of  $\text{CH}_2\text{O}$ . Its molar mass is approximately  $180 \text{ g mol}^{-1}$ . Determine its molecular formula.
- 20** 1.615 g of an anhydrous salt was placed in moist air. After complete hydration its mass increased to 2.875 g. The composition by mass of the salt was  $\text{Zn} = 40.6\%$ ,  $\text{S} = 19.8\%$  and  $\text{O} = 39.6\%$ . Determine the empirical formula of the compound.
- 21** 0.5000 g of an organic compound containing carbon, hydrogen and oxygen formed 0.6875 g of carbon dioxide and 0.5625 g of water on complete combustion. Determine the empirical formula of the compound.

## 1.3 Reacting masses and volumes

Revised

**Essential idea:** Mole ratios in chemical equations can be used to calculate reacting ratios by mass and gas volume.

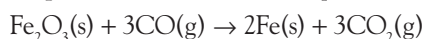
### Reacting masses

Revised

A balanced equation shows the reacting ratios of the amounts (in moles) of the reactants and products (stoichiometry). To find the mass of products formed in a reaction the mass of the reactants, the molar masses of the reactants and the balanced equation are needed. The stoichiometry of the reaction can also be found if the amounts (or masses) of each reactant that exactly react together and the amounts (or masses) of each product formed are known.

The example below shows how to calculate the approximate mass of iron(III) oxide needed to produce 28 g of iron by reduction with carbon monoxide.

■ **Step 1:** write the balanced equation



■ **Step 2:** convert the given mass data into moles

mass of iron required = 28 g

$$\text{number of moles of iron} = \frac{28 \text{ g}}{56 \text{ g mol}^{-1}} = 0.5 \text{ mol of iron}$$

0.5 mol of iron needs 0.25 mol of  $\text{Fe}_2\text{O}_3$

■ **Step 3:** use molar ratio from equation

The equation tells us that to get 2 mol of Fe we need 1 mol of  $\text{Fe}_2\text{O}_3$

■ **Step 4:** convert answer into grams

mass = number of moles  $\times$  molar mass

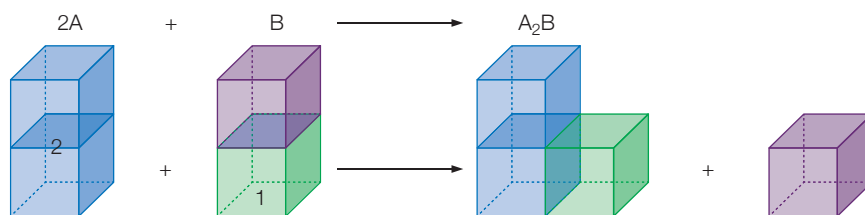
$$= 0.25 \text{ mol} \times 160 \text{ g mol}^{-1}$$

$$= 40 \text{ g}$$

### Limiting and excess reactants

Reactants can be either limiting or excess.

Chemicals react in stoichiometric ratios according to the balanced chemical equation. However, reagents are often not present in stoichiometric amounts; one of the reagents would be present as a *limiting reagent*, while other(s) would be present *in excess* (Figure 1.9). A limiting reactant is the first reagent to be consumed during the reaction, which causes the reaction to stop.



**Figure 1.9** Illustrating how excess reactant B (in purple) remains unused after the reaction has taken place

To identify the limiting reactant write the balanced equation. Calculate the amount of each reagent that is involved in the reaction. Convert the reactant quantities to the corresponding amounts (mol). Taking into account the stoichiometric ratio, compare which of the reactants produces the smallest amount of product to determine the limiting reagent.

#### Expert tip

When finding which reactant is limiting it is useful to do the following simple calculation. If the number of moles of each reactant is divided by its coefficient in the balanced equation, then the smallest number indicates the limiting reactant.

## ■ Percentage yield calculations

The experimental yield can be different from the theoretical yield.

Chemists use percentage yield calculations to assess the efficiency of a chemical reaction or series of chemical reactions. Few reactions are completely efficient and most reactions, especially organic reactions, give relatively low yields.

The reaction may be incomplete so that some of the reactants do not react. There may be side reactions producing by-products instead of the required chemical product. Recovery of all the product from the reaction mixture is usually impossible. Some of the product is usually lost during transfer of the chemicals from one container to another.

The *theoretical yield* is the mass of the product assuming that the reaction goes to completion according to the stoichiometric equation and the synthesis is 100% efficient. The *actual yield* is the mass of the product obtained.

$$\text{percentage yield} = \frac{\text{actual mass of product}}{\text{theoretical yield}} \times 100$$

A percentage yield must be less than 100%.

Often one of the chemicals in a reaction mixture is present in an amount which limits the theoretical yield. The other reactants are added in excess to make sure that the most valuable chemical is converted to the required product.

### ■ QUICK CHECK QUESTIONS

- 22** Ammonia gas reacts with heated copper(II) oxide to form copper, water and nitrogen gas. Formulate a balanced equation. Determine the limiting reagent if 20.00 g of ammonia reacts with 85.00 g of copper(II) oxide. Calculate the percentage yield if 68.50 g of copper is produced.
- 23** State two reasons why the experimental yield from a reaction is less than the theoretical yield.

## Solutions

Revised

The molar concentration of a solution is determined by the amount of solute and the volume of solution.

*Concentration* is the amount of solute (dissolved substance) in a known volume of solution (solute and solvent). It is expressed either in grams per cubic decimetre ( $\text{g dm}^{-3}$ ), or more usually in moles per cubic decimetre ( $\text{mol dm}^{-3}$ ). For very dilute solutions (or gases) it can be expressed in parts per million (ppm).

The amount of a solute can be calculated from the following expression:

$$\text{amount (mol)} = \text{volume of solution (dm}^3\text{)} \times \text{molar concentration (mol dm}^{-3}\text{)}$$

A molar concentration can be converted to a mass concentration by the following expression:

$$\text{mass concentration (g dm}^{-3}\text{)} = \text{molar mass (g mol}^{-1}\text{)} \times \text{molar concentration (mol dm}^{-3}\text{)}$$

### Expert tip

Read questions carefully to establish whether the question is concerned with the solute or one or more of the ions released when the solute dissolves.

### ■ QUICK CHECK QUESTIONS

- 24** Determine the concentration in  $\text{g dm}^{-3}$  and  $\text{mol dm}^{-3}$  of the solution formed by dissolving 4.00 g of solid sodium hydroxide in 125.00  $\text{cm}^3$  of aqueous solution.
- 25** Calculate the concentrations (in  $\text{mol dm}^{-3}$ ) of all the ions in 100.00  $\text{cm}^3$  of 0.020  $\text{mol dm}^{-3}$  aluminium sulfate solution,  $\text{Al}_2(\text{SO}_4)_3(\text{aq})$ .
- 26** Calculate the mass of potassium manganate(VII) needed to prepare 250.00  $\text{cm}^3$  of 0.02  $\text{mol dm}^{-3}$   $\text{KMnO}_4(\text{aq})$ .

## Titration

Titration can be used to find the reacting volumes of solutions. From the volumes and concentrations the equation can be determined. Titration of a solution of unknown concentration against a standard solution, with the equations for the reaction, allows an unknown concentration to be determined. If the solution to be titrated is concentrated then it may need to be diluted before titration. A dilution factor then needs to be applied in the calculation.

A standard solution is one of known concentration.

### ■ QUICK CHECK QUESTIONS

- 27** 1.00 cm<sup>3</sup> of concentrated hydrochloric acid was transferred with a graduated pipette to a 100.00 cm<sup>3</sup> volumetric flask. The volume was made up to 100.00 cm<sup>3</sup> with distilled water. A 10.00 cm<sup>3</sup> portion of the diluted solution from the volumetric flask was titrated with KOH(aq) and was neutralized by 24.35 cm<sup>3</sup> of potassium hydroxide of concentration 0.0500 mol dm<sup>-3</sup>. Calculate the concentration of the original concentrated hydrochloric acid in mol dm<sup>-3</sup>.
- 28** An unknown group 1 metal carbonate reacts with hydrochloric acid:  
 $M_2CO_3(aq) + 2HCl(aq) \rightarrow 2MCl(aq) + CO_2(g) + H_2O(l)$   
 A 3.960 g sample of  $M_2CO_3$  was dissolved in distilled water to make 250.00 cm<sup>3</sup> of solution. A 25.00 cm<sup>3</sup> portion of this solution required 32.85 cm<sup>3</sup> of 0.175 mol dm<sup>-3</sup> hydrochloric acid for complete reaction. Calculate the molar mass of  $M_2CO_3$  and determine the identity of the metal M.
- 29** A 20.00 g sample of a cleaning solution containing aqueous ammonia (NH<sub>3</sub>(aq)) was dissolved in water and the solution was made up to 500.00 cm<sup>3</sup> in a volumetric flask.  
 A 25.00 cm<sup>3</sup> portion of this solution was then neutralized with 26.85 cm<sup>3</sup> of 0.200 mol dm<sup>-3</sup> sulfuric acid. Calculate the percentage by mass of ammonia in the cleaning solution.

### Expert tip

Dilution involves the addition of more to a solution. The amount of solute does not change but the concentration and volume of the solution change. Sampling involves the collection of a portion from a solution. The concentration of the portion is the same as the original solution but the amount of dissolved solute present in the new solution changes.

### Expert tip

A standard solution is a solution of known concentration that remains constant. Sodium hydroxide cannot be used as a standard solution: it absorbs atmospheric carbon dioxide and slowly reacts to form sodium carbonate.

### ■ Back titration

Not all substances can be quantitatively determined by a direct titration method. Back titration (Figure 1.10) is an indirect titration method usually used when a compound is an insoluble solid in which the end point is difficult to detect, or a reaction is too slow.

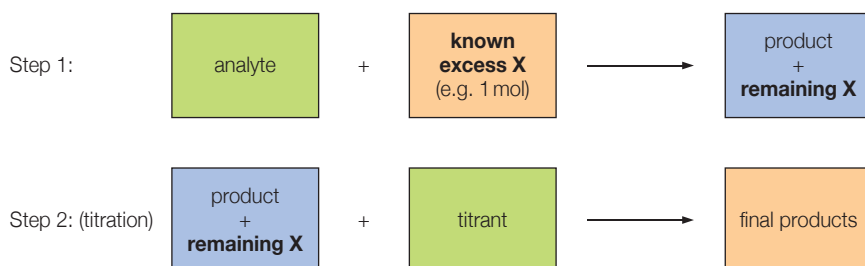
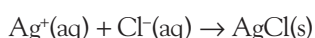


Figure 1.10 Concept of a back titration

Similarly, compounds which contain impurities that may interfere with direct titration or contain volatile substances (ammonia, iodine, etc.) that may result in inaccuracy arising due to loss of substance during titration require a back titration method.

### ■ Redox and precipitation titrations

Calculations for redox titrations are identical to that of simple acid–base titrations. Completion of a redox reaction may be shown by the final disappearance of a reactant. Titrations involving precipitation often involve determining the chloride ion concentration by titration with silver nitrate. As the silver nitrate solution is slowly added, a precipitate of silver chloride forms:



The end point of the titration occurs when all the chloride ions are precipitated.

### ■ QUICK CHECK QUESTION

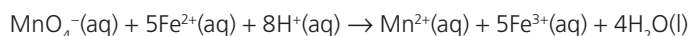
- 30** An impure sample of calcium carbonate with a mass of 0.95 g was reacted with 50.00 cm<sup>3</sup> of 1.00 mol dm<sup>-3</sup> hydrochloric acid (an excess). The resulting solution was transferred to a volumetric flask. The volume was made up to 100.00 cm<sup>3</sup> with distilled water and 10.00 cm<sup>3</sup> of this solution was titrated with 11.10 cm<sup>3</sup> of 0.300 mol dm<sup>-3</sup> sodium hydroxide solution. Determine the percentage purity by mass of the calcium carbonate sample.

Explain why the sodium hydroxide solution has to be freshly prepared.

### ■ QUICK CHECK QUESTIONS

**31** To determine the concentration of chloride ion in a sample of pool water, a 50.00 cm<sup>3</sup> sample of the pool water was titrated with 0.500 mol dm<sup>-3</sup> AgNO<sub>3</sub>(aq) solution. At the equivalence point, it was found that 53.40 cm<sup>3</sup> of AgNO<sub>3</sub>(aq) solution had been added. Calculate the [Cl<sup>-</sup>(aq)] in the sample of pool water.

**32** Determine the volume of 0.020 mol dm<sup>-3</sup> KMnO<sub>4</sub>(aq) solution needed to react with 25.00 cm<sup>3</sup> of 0.100 mol dm<sup>-3</sup> iron(II) sulfate solution:



### ■ Consecutive reactions

It often takes more than one step to get the desired product. Reactions that are carried out one after another in sequence to yield a final product are called consecutive reactions. Any substance that is produced in one step and is consumed in another step of a multistep process is called an intermediate. The overall reaction is the chemical equation that expresses all the reactions occurring in a single overall equation.

### ■ Parts per million

When dealing with the levels of pollution in water, the concentrations are often very low, even though they can be dangerous. An example would be the levels of heavy metals in river water, for instance. In this case, the statutory allowed levels are often quoted in terms of the unit, parts per million (ppm).

The number of milligrams of solute per kg of solution (mg/kg) is equivalent to one ppm, since 1 mg = 10<sup>-3</sup> g and 1 kg = 10<sup>3</sup> g. Assuming the density of water is 1.00 g cm<sup>-3</sup>, 1 dm<sup>3</sup> of solution has a mass of 1 kg and hence, 1 mg dm<sup>-3</sup> = 1 ppm. This is approximately true for dilute aqueous solutions. Parts per million concentrations are essentially mass ratios (solute to solution) × a million (10<sup>6</sup>).

To convert concentrations in mg dm<sup>-3</sup> (or ppm in dilute solution) to molarity (mol dm<sup>-3</sup>), divide the mass in grams by the molar mass of the analyte (chemical under analysis) to convert mass in milligram (mg) into a corresponding number of moles. To convert from molarity to mg dm<sup>-3</sup> (or ppm in dilute solution), multiply by the molar mass of the analyte.

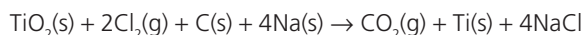
### ■ Atom economy

This term refers to the efficiency of a chemical process in terms of the atoms that are lost as by-products to the product. Industrial processes with poor atom economy are inefficient in terms of resources and often produce undesirable waste. The atom economy is calculated as a percentage of desired product divided by the total mass of products and scaled up by multiplying by 100 to make a percentage.

$$\text{Atom economy} = \frac{\text{molar mass of desired products}}{\text{total molar mass of all reactants}} \times 100$$

### ■ QUICK CHECK QUESTION

**35** There are two methods for extracting titanium – the traditional method and electrolysis. The following reaction is used for the traditional method:



(The desired product is titanium; CO<sub>2</sub> and NaCl are waste products.) The electrolytic method can be summarized by this reaction:



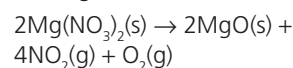
Calculate the atom economy for both of these reactions and identify the more efficient process.

### Expert tip

Any titration data that has an initial trial that is significantly different from other titre volumes should not be included in a calculation of the average titre volume.

### ■ QUICK CHECK QUESTION

**33** A 2.23 g sample of magnesium nitrate was fully decomposed by heating:



The magnesium oxide produced was reacted with hydrochloric acid and required 33.20 cm<sup>3</sup> of hydrochloric acid for complete reaction. Calculate the concentration, in mol dm<sup>-3</sup>, of the hydrochloric acid.

### Expert tip

The mass of solute and solute must have the same units.

### ■ QUICK CHECK QUESTION

**34** The density of seawater is 1027 kg m<sup>-3</sup> and the density of pure water is 1.000 g cm<sup>-3</sup>. Calculate the mass of sodium ions present in 1.000 dm<sup>3</sup> of sea water, assuming only NaCl is present. Calculate the concentration of sodium ions, Na<sup>+</sup>(aq), in mol dm<sup>-3</sup> and ppm.



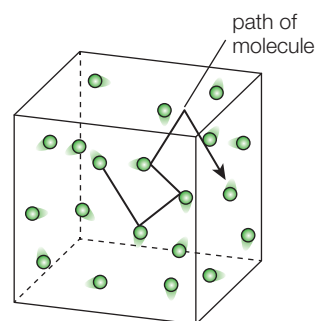
# 1.4 Gases

Revised

## Kinetic theory of gases

Revised

The kinetic theory of gases assumes that a gas (Figure 1.11) is composed of widely spaced tiny particles (atoms or molecules); the particles are in rapid and random motion, moving in straight lines (between collisions); the particles have negligible volume; there are no intermolecular forces; the particles frequently collide with each other and with the walls of the container. Gas pressure arises from the force caused by collisions of gas molecules with the walls of the container. All collisions are perfectly elastic, that is, no momentum is lost on collision; the average kinetic energy of gas particles is directly proportional to the absolute temperature and remains constant at any temperature.



**Figure 1.11** The kinetic theory description of a gas

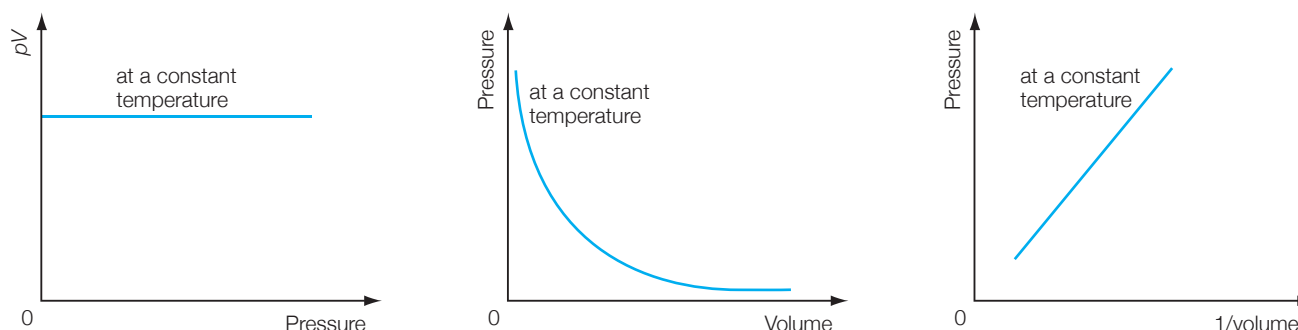
Revised

## Gas laws

The volume of a gas is changed by changes in temperature and pressure. The changes in the volume of a gas do not depend on the chemical nature of the gas. The behaviour of gases can be described by the gas laws. The gas laws were formulated from early experiments with gases, but can also be derived from the ideal gas equation.

### Boyle's law

Boyle's law (Figure 1.12) states that, at a constant temperature, the volume of a fixed mass of an ideal gas is inversely proportional to its pressure. The mathematical expressions of Boyle's law are:  $V \propto \frac{1}{P}$  (at constant temperature) or  $PV = k$  (a constant) or  $P_1V_1 = P_2V_2$ , where  $P_1V_1$  = original pressure and volume, respectively, and  $P_2V_2$  = final pressure and volume, respectively.



**Figure 1.12** Graphical methods of representing Boyle's law

### Charles' law

Charles' law states that, at a constant pressure, the volume of a fixed mass of an ideal gas is directly proportional to its absolute temperature. The mathematical expressions of Charles' law are:  $V \propto T$  ( $T = 273 + t^\circ\text{C}$ ), or  $\frac{V}{T} = \text{constant}$ , or  $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ .

Charles' law can be represented graphically (Figure 1.13) by plotting the volume of the gas against the temperature. On extrapolation, the graph meets the temperature axis at approximately  $-273^\circ\text{C}$  (to 3 s.f.). This temperature is called absolute zero and is adopted as the zero on the absolute or thermodynamic temperature scale (units of kelvin).

#### Expert tip

Any consistent units for pressure and volume may be used in gas law calculations, but the temperature must always be expressed in kelvin, because the gas laws are derived from thermodynamics.

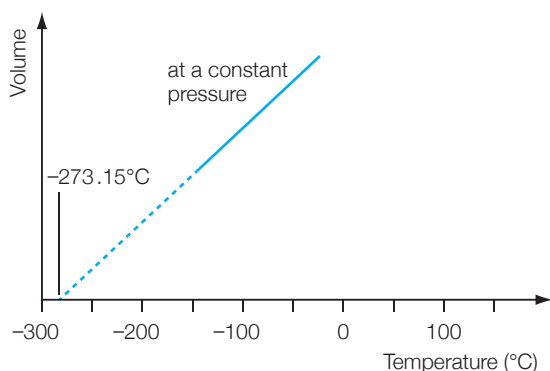


Figure 1.13 Charles' law

## Temperature scales

To demonstrate this proportionality, you must use temperatures in kelvin, K (see Figure 1.13). If you use degrees Celsius (°C), then the graph will still be a straight line, but it does not go through the absolute zero.

When using these relationships for gases, it is important to use temperature values on the absolute temperature scale – where the values are in kelvin. This scale is the SI scale for temperature and is the true thermodynamic scale of temperature.

The Celsius scale is also commonly used in scientific and everyday work, but its basis is related to the physical properties of water (the 'fixed points' of the freezing point and the boiling point). The size of the 'degree' on both scales is the same, so it is easy to convert between the two (Figure 1.14).

The Fahrenheit temperature scale can also be used in everyday terms, but is no longer used in scientific work. This scale also uses the fixed points for water, with the values being 32°F and 312°F for the freezing point and boiling point respectively.

$$\text{temperature (K)} = \text{temperature (°C)} + 273.15$$

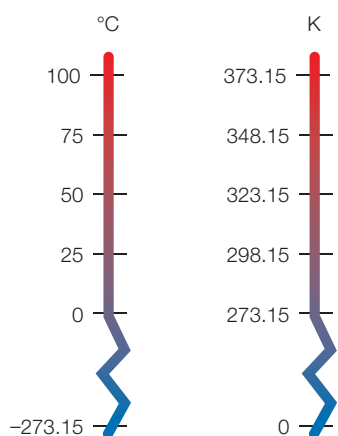


Figure 1.14 Comparison between the absolute temperature scale and the Celsius scale

Boyle's law and Charles' law can be combined to give the general gas equation:

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

This can be used to solve any gas law problem.

## Pressure law

The pressure law states that for a fixed mass of gas (at constant volume) its absolute temperature is directly proportional to pressure. The mathematical expressions of the pressure law are:  $P \propto T$ , or  $\frac{P}{T} = \text{constant}$  or  $\frac{P_1}{T_1} = \frac{P_2}{T_2}$ , where  $P_1$  represents the initial pressure,  $T_1$  represents the initial absolute temperature,  $P_2$  represents the final pressure and  $T_2$  represents the final absolute temperature. Table 1.4 summarizes the various gas laws.

### Common mistake

Note that a doubling of the temperature in degrees Celsius is *not* a doubling of the absolute temperature. For example a doubling of the temperature from 200°C to 400°C is only a rise from  $(200 + 273) = 473 \text{ K}$  to  $(400 + 273) = 673 \text{ K}$ , that is, a ratio of 673/473 or 1.42.

### Expert tip

When using this general gas equation  $\left(\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}\right)$  you will need to be careful with the units you use:

- Most importantly,  $T$  must be in kelvin!
- The units of pressure and volume can be any, so long as they are the same for both the values in the equation.

**Table 1.4** Summary of the gas laws

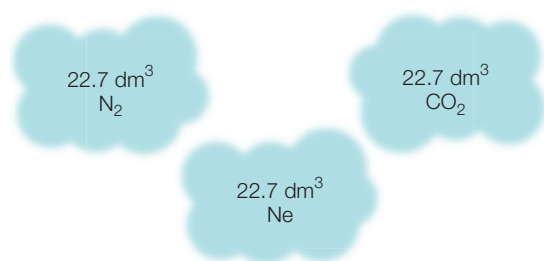
	Constant factor	Relationship	Effect
Boyle's law	Temperature	$V \propto \frac{1}{P}$	$P_1 \times V_1 = P_2 \times V_2$
Charles' law	Pressure	$V \propto T$	$\frac{V_1}{T_1} = \frac{V_2}{T_2}$
Pressure law	Volume	$P \propto T$	$\frac{P_1}{T_1} = \frac{P_2}{T_2}$
Avogadro's law	Volume and amount	$V \propto n$	$\frac{V_1}{n_1} = \frac{V_2}{n_2}$

### ■ QUICK CHECK QUESTIONS

- 36** Indicate whether gas pressure increases or decreases with each of the following changes in a sealed container: increasing the temperature, increasing the volume and increasing the number of gas molecules.
- 37** 50.00 cm<sup>3</sup> of a gas at 100.00 kPa is compressed (at constant temperature) to a volume of 20.00 cm<sup>3</sup>. Determine the final pressure of the gas.
- 38** 50.0 cm<sup>3</sup> of a sample of a gas is heated (at constant pressure) from 25.0 to 200.0°C. Calculate the final volume of the gas if the pressure remains constant.
- 39 a** A container is filled with gas at a pressure of 15 kPa and is cooled from 2°C to -40°C. Calculate the final pressure, assuming the volume remains constant.
- b** The density of a gas at STP is 1.78 g dm<sup>-3</sup>. Calculate the molar mass of the gas.

### ■ Avogadro's law

Avogadro's law states that equal volumes of all gases at the same temperature and pressure contain equal numbers of atoms or molecules (Figure 1.15). The volume occupied by one mole of any gas (molar volume) depends on the temperature and pressure of the gas. The molar volume of any gas at standard temperature and pressure (STP) is 22.7 dm<sup>3</sup> (Figure 1.15). The conditions of STP are: temperature 0°C (273 K); pressure:  $1.00 \times 10^5$  Pa N m<sup>-2</sup> (1 bar).



**Figure 1.15** An illustration of the molar gas volume (at STP). All samples of the gases contain the same number of particles (atoms or molecules)

The molar gas volume allows interconversion (Figure 1.16) between the volume of a pure gas (behaving ideally) and its amount (in moles). The ratios of reacting volumes of gases can be used to deduce the stoichiometry of a reaction.



**Figure 1.16** Summary of interconversions between the amount of gas and volume (at STP)

### Common mistake

The molar gas volume of 22.7 dm<sup>3</sup> mol<sup>-1</sup> only applies to gases (behaving ideally). It can be applied to all gases for IB calculations. It does not apply to pure liquids.

### ■ QUICK CHECK QUESTIONS

- 40** 100 cm<sup>3</sup> of carbon monoxide react with 50 cm<sup>3</sup> of oxygen to form 100 cm<sup>3</sup> of carbon dioxide. Deduce the equation for the reaction.
- 41** 50 cm<sup>3</sup> of hydrogen were exploded with 80 cm<sup>3</sup> of air (80% by volume nitrogen and 20% by volume oxygen). Calculate the composition of the resulting gas, all measurements being made at STP.
- 42** Lead(II) nitrate decomposes to form lead(II) oxide, nitrogen dioxide and oxygen. Determine the total volume of gas (at STP) produced when 33.10 g of lead(II) nitrate is completely decomposed by strong heating.
- 43** Determine the volume of sulfur trioxide produced by the complete reaction of 1000 cm<sup>3</sup> of sulfur dioxide with oxygen. Determine the volume of oxygen required to react with the sulfur dioxide.
- 44** Calculate the volume occupied (at STP) by the following gases: 0.250 mol of chlorine and 106.35 g of chlorine gas.

## The ideal gas equation

Revised

The behaviour of an ideal gas can be described by the ideal gas equation:

$$pV = nRT$$

The ideal gas equation allows the conversion from volumes of gases to amounts (in moles) at any particular temperature and pressure. When feeding values into the ideal gas equation you should be careful to use the correct units:

- pressure,  $p$ , must be in Pa (N m<sup>-2</sup>); if kPa are given, multiply by 10<sup>3</sup>
- volume,  $V$ , must be in m<sup>3</sup>; if dm<sup>3</sup> are given, divide by 10<sup>3</sup>, or if cm<sup>3</sup> are given, divide by 10<sup>6</sup>
- amount of gas (mol),  $n$ : this is often calculated using  $n = \frac{m}{M}$
- absolute temperature,  $T$  must be in kelvin; if °C is given, add 273 (or, very precisely, 273.15)
- gas constant,  $R = 8.31 \text{ J K}^{-1} \text{ mol}^{-1}$ .

The ideal gas equation can be used to determine the relative molecular mass,  $M_r$ , of compounds in the gaseous state, and to determine the density,  $\rho$ , of a gas.

By recording the mass ( $m$ ) of the gas, the pressure ( $p$ ), temperature ( $T$ ) and volume of the gas ( $V$ ) and using this modified form of the ideal gas equation shown below, the relative molecular mass of the compound in its gaseous state can be determined.

If the relative molecular mass ( $M_r$ ), pressure ( $p$ ) and the absolute temperature ( $T$ ) of the gas are known, then an alternative form of the ideal gas equation can be used to find the density ( $\rho$ ) of the compound in its gaseous state.

$$\rho = \frac{pM_r}{RT}$$

### ■ QUICK CHECK QUESTIONS

- 45** A sample of a gas has a mass of 0.112 g and it occupies a volume of 81.80 cm<sup>3</sup> at a temperature of 127 °C and a pressure of 100 kPa. Calculate the relative molecular mass of the gas (to the nearest integer).
- 46** A syringe containing 0.15 g of a volatile liquid was injected into a graduated gas syringe at a temperature of 90.0 °C and a pressure of 101 kPa. The vapour occupied a volume of 62.2 cm<sup>3</sup>. Calculate the relative molecular mass of the gas.
- 47** The density of a gas is 2.615 g cm<sup>-3</sup> at 298 K and 101 kPa. Calculate the molar mass of the gas.

### ■ Comparison between a real gas and an ideal gas

Chemists make predictions about gas behaviour by assuming that the gases are ideal and the gas laws are exactly obeyed. This assumption allows chemists to make very accurate estimations of measurements.

#### Expert tip

Suitable units for a calculation involving any of these forms of the ideal gas equation are SI base units for pressure (Pa) and volume (m<sup>3</sup>). Temperature must be expressed in kelvin.

The molecules of an ideal gas are assumed to be points with no attractive or repulsive forces operating between the particles (atoms or molecules). It is a hypothetical state but gases (especially the noble gases) approach ideal behaviour under certain conditions. Table 1.5 summarizes the difference between the behaviour of real and ideal gases.

**Table 1.5** The behaviour of real and ideal gases

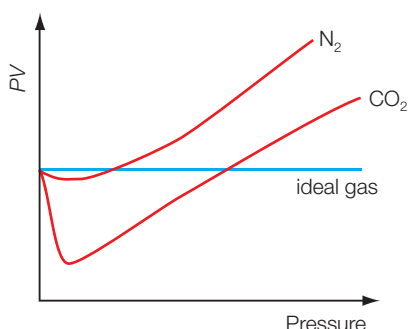
Real gas	Ideal gas
1. Each gas particle has a definite volume and hence, occupies space. So the volume of the gas is the volume of the particles and the spaces between them.	Each gas particle has a negligible (almost zero) volume and hence, the volume of the gas is the volume of the space between the particles.
2. Attractive forces are present between the particles.	There are negligible attractive forces operating between the particles.
3. Collisions between the particles are non-elastic, and hence, there is a change (loss or gain) of kinetic energy from the particles after collision.	Collisions between the particles are elastic, and hence, there is no loss of kinetic energy from the particles after collision.
<ul style="list-style-type: none"> <li>Real gases can be compressed into liquids but molecular kinetic theory predicts that ideal gases cannot be liquefied.</li> <li>Ideal gases obey the ideal gas equation: <math>pV = nRT</math>.</li> <li>A gas behaves most like an ideal gas at high temperatures and low pressures when the particles move fast and are far apart, which minimizes attractive forces between particles.</li> </ul>	

## NATURE OF SCIENCE

Chemists and physicists use simplified models – the three ‘gas laws’ – based on ideal behaviour to describe gases. Deviation of the experimental values from the predicted values led to a refinement of models of gas to include intermolecular forces and the actual volume of gas particles. The gas laws were developed from empirical data (observations and measurements) and were all later explained in terms of molecular kinetic theory.

## Deviation from ideal gas behaviour

The deviation from ideal gas behaviour is shown by plotting  $pV$  against  $p$  (Figure 1.17) or  $\frac{pV}{RT}$  against  $p$ . For an ideal gas,  $pV$  must be constant at all pressures if the temperature is constant. When  $pV$  is plotted against  $p$ , or when  $\frac{pV}{RT}$  is plotted against  $p$ , a straight line is obtained which runs parallel to the  $x$ -axis of pressure. The extent of deviation from ideal behaviour depends on pressure, temperature, the size of the gas particles and the strength of the intermolecular forces.



**Figure 1.17** Deviation from ideal behaviour at high pressure

## ■ QUICK CHECK QUESTIONS

- 48** State three assumptions underlying the kinetic theory of gases.
- 49** Explain Boyle's law using this theory.
- 50** State and explain the conditions at which gases would show almost ideal behaviour.

## ■ QUICK CHECK QUESTION

- 51** State which of the following gases would behave most ideally (under the same conditions of temperature and pressure: hydrogen, fluorine and hydrogen fluoride.  
State and explain which one would show the greatest deviation from ideal behaviour.



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