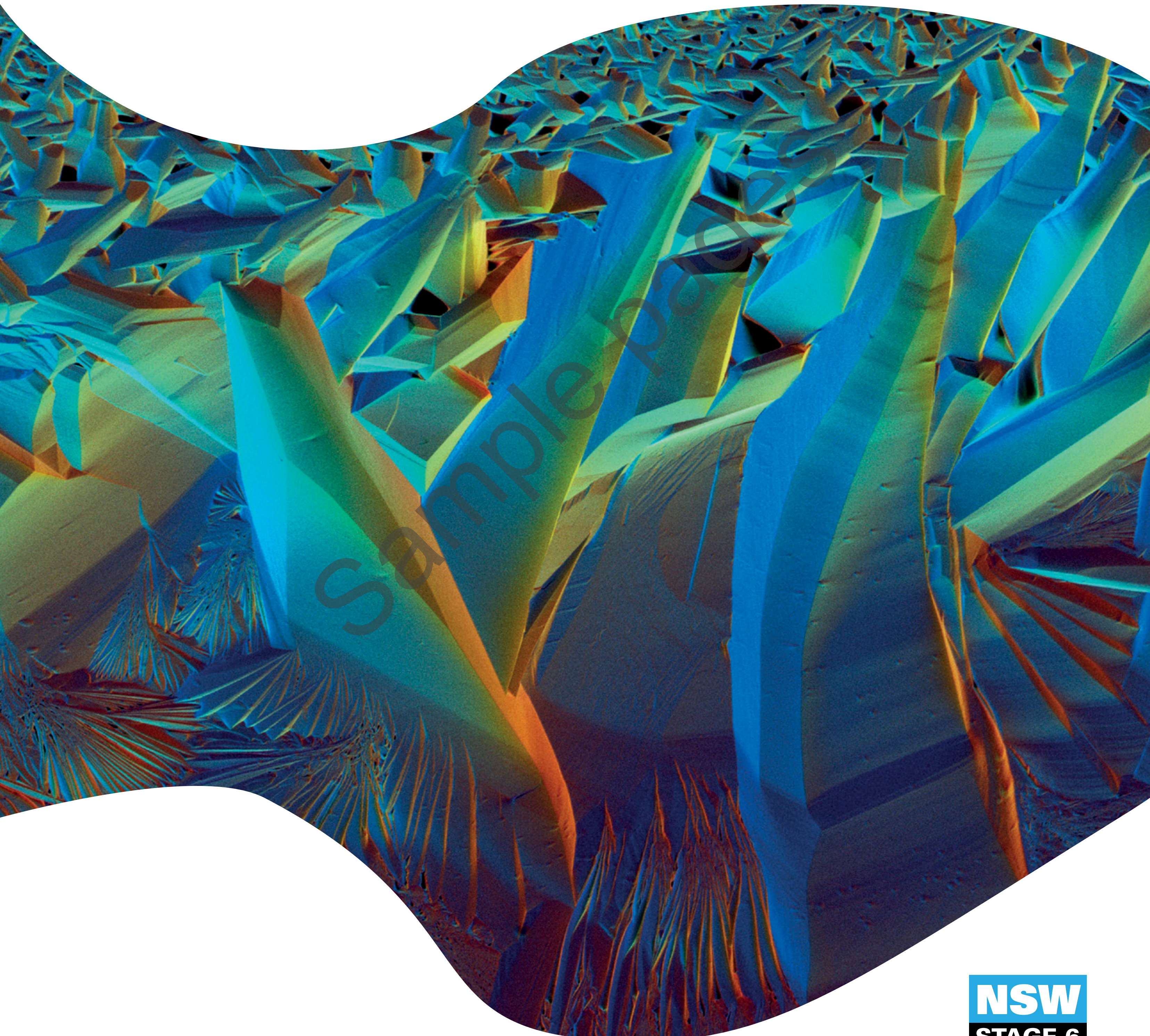
PEARSON

CHEMISTRY

NEW SOUTH WALES

STUDENT BOOK

11



NSW
STAGE 6

Contents

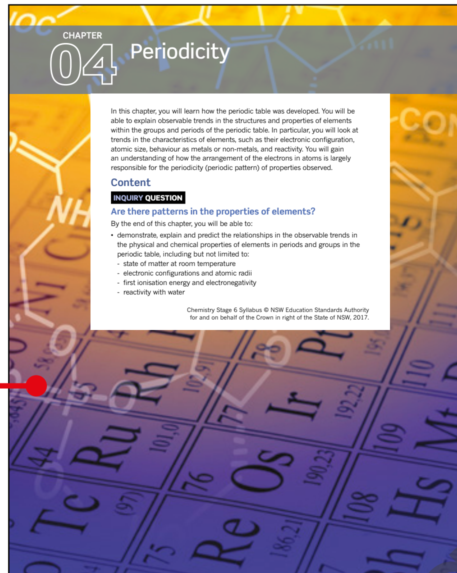
Module 3 Reactive chemistry

Module 4 Drivers of reactions

How to use this book

Pearson Chemistry 11 New South Wales

Pearson Chemistry 11 New South Wales has been written to fully align with the new Stage 6 Syllabus for New South Wales Chemistry. The book covers Modules 1 to 4 in an easy-to-use resource. Explore how to use this book below.

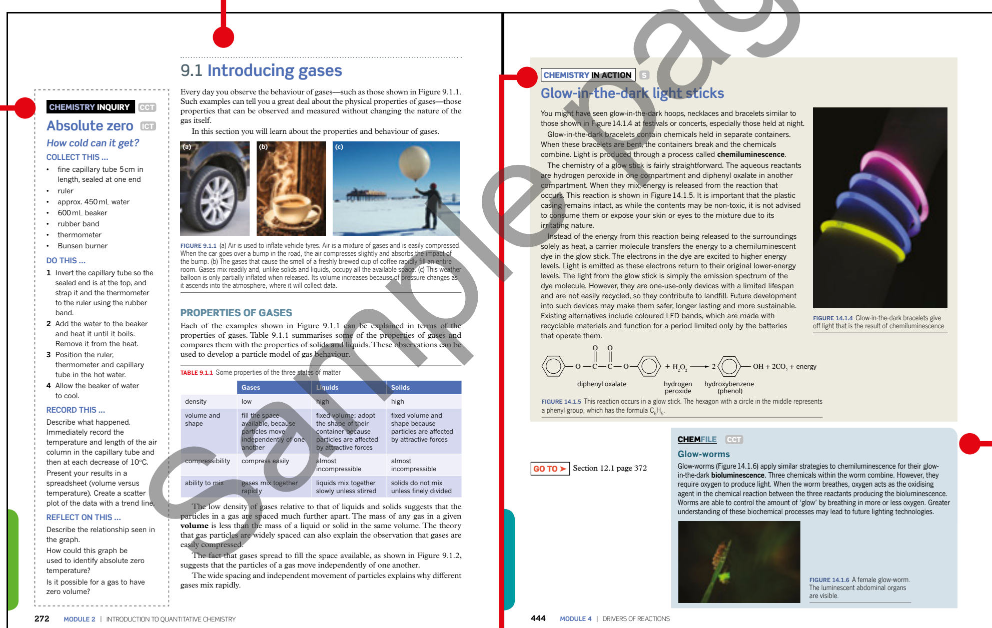


Chapter opener

The chapter opening page links the Syllabus to the chapter content. Key content addressed in the chapter is clearly listed.

Section

Each chapter is clearly divided into manageable sections of work. Best-practice literacy and instructional design are combined with high-quality, relevant photos and illustrations to help students better understand the ideas or concepts being developed.



Chemistry Inquiry

Chemistry Inquiry features are inquiry-based activities that pre-empt the theory and allow students to engage with the concepts through a simple activity that sets them up to 'discover' the science before they learn about it. They encourage students to think about what happens in the world and how science can provide explanations.

Chemistry in Action

Chemistry in Action boxes place chemistry in an applied situation or a relevant context. They refer to the nature and practice of chemistry, its applications and associated issues, and the historical development of its concepts and ideas.

ChemFile

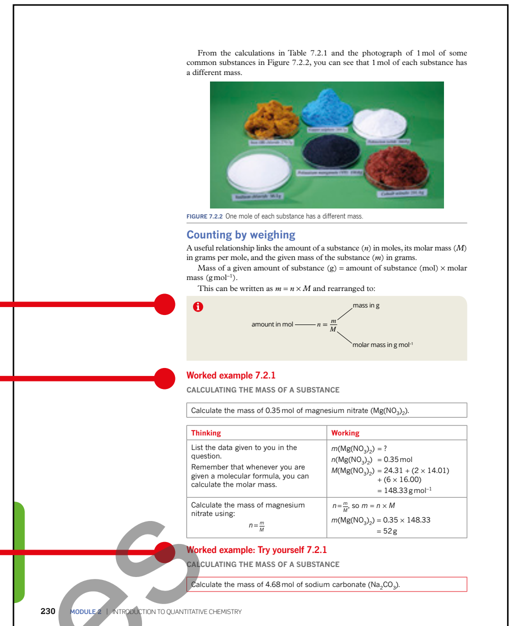
ChemFiles include a range of interesting and real-world examples to engage students.

Highlight box

Highlight boxes focus students' attention on important information, such as key definitions, formulae and summary points.

Worked examples

Worked examples are set out in steps that show thinking and working. This format greatly enhances student understanding by clearly linking underlying logic to the relevant calculations. Each Worked example is followed by a Try yourself activity. This mirror problem allows students to immediately test their understanding. Fully worked solutions to all Worked example: Try yourself activities are available on *Pearson Chemistry 11 New South Wales* Reader+.



Counting by weighing

Worked example 7.2.1

CALCULATING THE MASS OF A SUBSTANCE

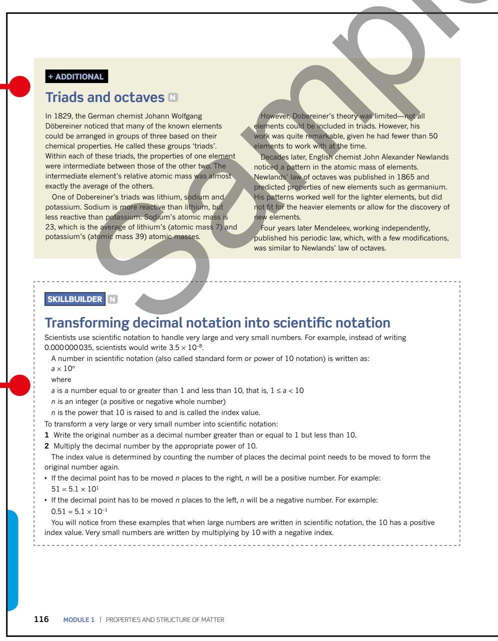
Worked example: Try yourself 7.2.1

Additional content

Additional content features include material that goes beyond the core content of the Syllabus. They are intended for students who wish to expand their depth of understanding in a particular area.

Section summary

Each section has a section summary to help students consolidate the key points and concepts of the section.

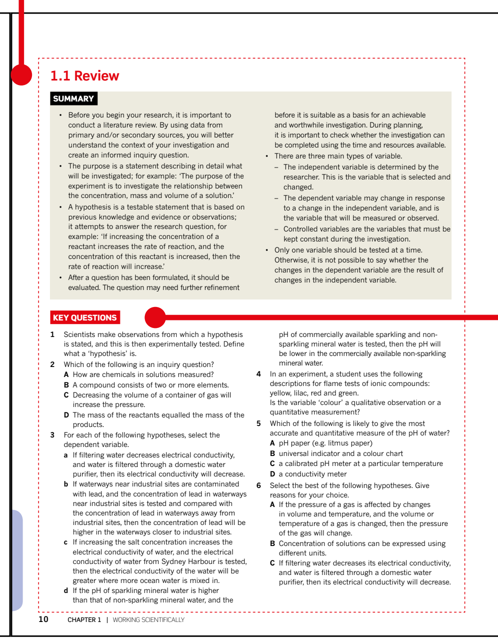


+ ADDITIONAL

Triads and octaves

SKILLBUILDER

Transforming decimal notation into scientific notation



1.1 Review

SUMMARY

KEY QUESTIONS

SkillBuilder

Skillbuilders outline methods or techniques. They are instructive and self-contained. They step students through the skill to support science application.

Section review questions

Each section finishes with key questions to test students' understanding of and ability to recall the key concepts of the section.

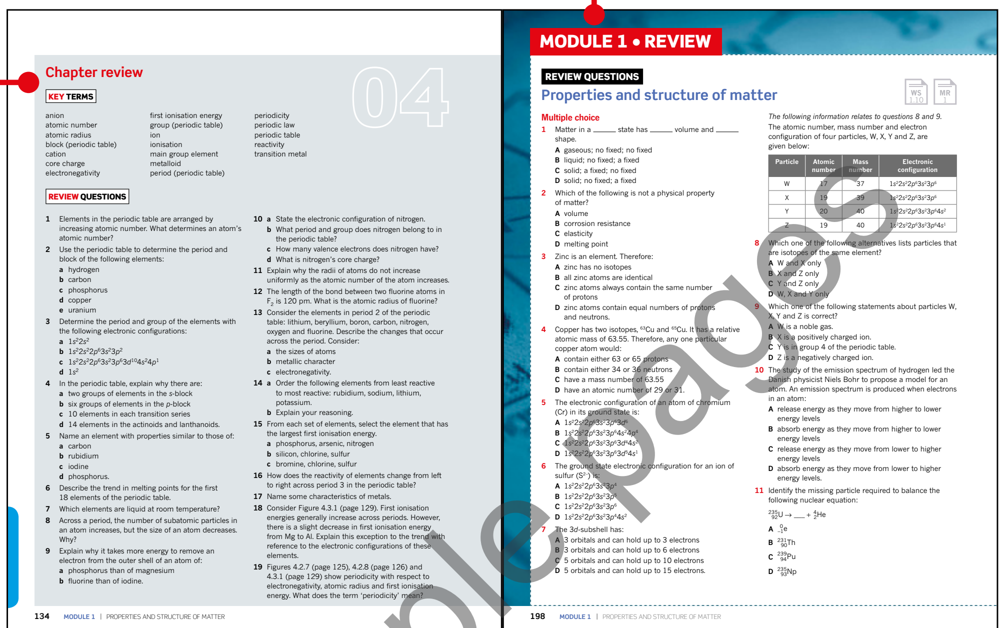
How to use this book

Chapter review

Each chapter finishes with a list of key terms covered in the chapter and a set of questions to test students' ability to apply the knowledge gained from the chapter.

Module review

Each module finishes with a comprehensive set of questions, including multiple-choice, short-answer and extended-response questions. These assist students in drawing together their knowledge and understanding, and applying it to these types of questions.



Icons

The New South Wales Stage 6 Syllabus 'Learning across the curriculum' and 'General capabilities' content are addressed throughout the series and are identified using the following icons.

AHC A CC CCT DD EU ICT
IU L N PSC S WE

'Go to' icons are used to make important links to relevant content within the Student Book.

GO TO ➤

This icon indicates when it is the best time to engage with a worksheet (WS), a practical activity (PA), a depth study (DS) or module review (MR) questions in the *Pearson Chemistry 11 Skills and Assessment Book*.

WS 3.1

This icon indicates the best time to engage with a practical activity on *Pearson Chemistry 11 New South Wales* Reader+.

PA 3.1

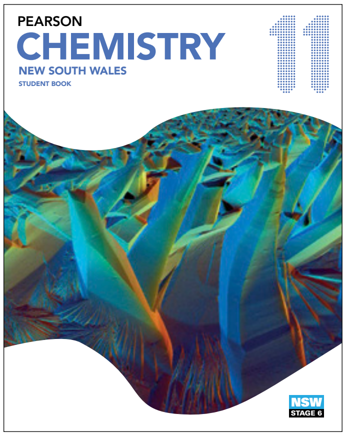
Glossary

Key terms are shown in **bold** in sections and listed at the end of each chapter. A comprehensive glossary at the end of the book includes and defines all the key terms.

Answers

Numerical answers and short-response answers are included at the back of the book. Comprehensive answers and fully worked solutions for all section review questions, Worked examples: Try yourself activities, chapter review questions and module review questions are provided via *Pearson Chemistry 11 New South Wales* Reader+.

Pearson Chemistry 11
New South Wales

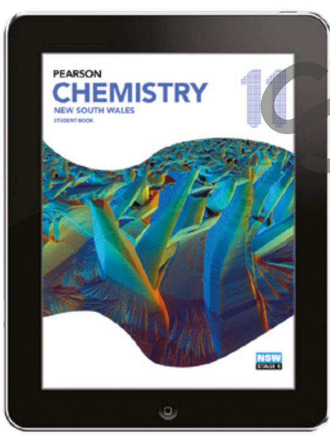


Student Book

Pearson Chemistry 11 New South Wales has been written to fully align with the new Stage 6 Syllabus for New South Wales. The Student Book includes the very latest developments in and applications of chemistry and incorporates best-practice literacy and instructional design to ensure the content and concepts are fully accessible to all students.

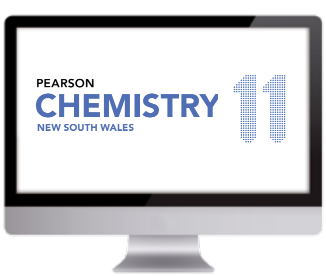
Skills and Assessment Book

The *Skills and Assessment Book* gives students the edge in preparing for all forms of assessment. Key features include a toolkit, key knowledge summaries, worksheets, practical activities, suggested depth studies and module review questions. It provides guidance, assessment practice and opportunities for developing key skills.



Reader+ the next generation eBook

Pearson Reader+ lets you use your Student Book online or offline on any device. Pearson Reader+ retains the look and integrity of the printed book. Practical activities, interactives and videos are available on Pearson Reader+, along with fully worked solutions for the Student Book questions.



Teacher Support

The Teacher Support available includes syllabus grids and a scope and sequence plan to support teachers with programming. It also includes fully worked solutions and answers to all Student Book and *Skills and Assessment Book* questions, including worksheets, practical activities, depth studies and module review questions. Teacher notes, safety notes, risk assessments and a laboratory technician checklist and recipes are available for all practical activities. Depth studies are supported with suggested marking schemes and exemplar answers.

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CHAPTER 07

The mole concept

Chemists in fields as diverse as environmental monitoring, pharmaceuticals and fuel production routinely carry out chemical reactions as part of their work. It is important for them to be able to measure specific quantities of chemicals quickly and easily, in part because the amount of products formed depends on the amount of reactants.

By the end of this chapter, you will have a greater understanding of the way in which chemists measure quantities of chemicals, in particular of the way they can accurately count the number of particles in samples of elements and compounds simply by weighing them. This is essential for designing and producing materials, including cosmetics, fuels, fertilisers, pharmaceuticals and building materials.

Content

INQUIRY QUESTION

How are measurements made in chemistry?

By the end of this chapter, you will be able to:

- conduct a practical investigation to demonstrate and calculate the molar mass (mass of 1 mole) of:
  - an element
  - a compound (ACSCH046) ICT N
- conduct an investigation to determine that chemicals react in simple whole number ratios by moles ICT N
- explore the concept of the mole and relate this to Avogadro's constant to describe, calculate and manipulate masses, chemical amounts and numbers of particles in: (ACSCH007, ACSCH039) ICT N
  - moles of elements and compounds $n = \frac{m}{M}$ (n = number of moles, m = mass in g, M = molar mass in g mol^{-1})
  - percentage composition calculations and empirical formulae
  - limiting reagent reactions

7.1 Introducing the mole

CHEMISTRY INQUIRY CCT N

Moles of household items

How are moles measured in chemistry?

COLLECT THIS …

- half a cup of each of:
  - table sugar (sucrose, $C_{12}H_{22}O_{11}$)
  - water
  - aluminium foil, tightly packed
  - table salt (NaCl)
  - bicarbonate of soda ($NaHCO_3$)
- digital kitchen scales
- measuring cup
- mobile phone or digital camera

DO THIS …

1. Zero the scales with the measuring cup on the balance pan.
2. Measure out half a cup of each substance and weigh it. Record the masses in your table.
3. Take a photo as you weigh the substance.
4. If the substance came in a packet, write down the mass of the contents of the packet when it was unopened.

RECORD THIS …

Draw up the table below in your workbook to record your results.

Use number of moles = $\frac{\text{mass (g)}}{\text{molar mass}}$ to calculate the number of moles in half a cup of each substance.

Calculate the number of molecules or particles in your sample using:

number of molecules = number of moles × 6.022 × 10^{23}.

REFLECT ON THIS …

How do we measure the number of moles of substances?

Rank the substances from largest to smallest number of moles in half a cup.

Do liquids, ionic compounds, molecular substances or elements have the largest number of particles in half a cup?

Identify any patterns in your results and attempt to explain your observations.

What could you do next time to improve your experiment?

Substance	Mass (g)	Molar mass	Number of moles in half a cup	Mass of the full packet	Number of moles in the full packet
table sugar ($C_{12}H_{22}O_{11}$)		342.30			
water		18.02		–	–
aluminium		26.98			
table salt (NaCl)		55.44			
bicarbonate of soda ($NaHCO_3$)		84.01			

It is often essential for chemists to be able to measure the exact number of particles of an element or compound. However, the particles in elements and compounds (atoms, ions and molecules) are so small that it would be difficult to count them individually or even by the thousands of millions. If it were possible to count individual particles, the numbers in even very small samples would be huge and very inconvenient to work with.

The ice cubes shown in Figure 7.1.1 each contain more than 100 000 000 000 000 000 000 000 (or 10^{23}) water molecules (H_2O). As each water molecule is composed of two hydrogen atoms and one oxygen atom, the number of individual atoms in each ice cube is greater than 10^{23}. A quantity that allows chemists to measure accurate amounts of extremely small particles is required.

In this section, you will learn about the very convenient quantity used by chemists: the mole.

FIGURE 7.1.1 Each of these ice cubes contains more than 10^{23} water (H_2O) molecules. Molecules and atoms are so small, and the numbers of them in everyday samples are so large, that it would be very inconvenient to always count them individually.

THE CHEMIST'S COUNTING UNIT

A dozen is a convenient quantity for buying the eggs shown in Figure 7.1.2. For atoms, ions and molecules, which are much smaller than eggs, a quantity that describes a much larger number is needed.

The accepted quantity for chemists is the **mole**. The mole is the unit for **amount of substance** and is given the symbol n and the abbreviation 'mol'.

So n(glucose) = 2 mol is read as 'the amount of glucose molecules is 2 moles'.

Chemists use the mole as a counting measure. Figure 7.1.3 shows some quantities that you would be very familiar with, such as pair, dozen and ream. One dozen is equal to 12, two dozen equals 24, 20 dozen equals 240 and half a dozen equals 6. In the same way, chemists know that 1 mole is equivalent to a certain number and that 2 moles, 20 moles and half a mole are all multiples of that number.

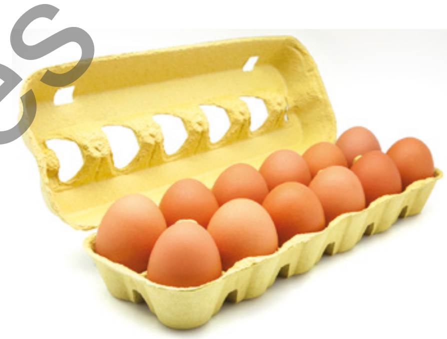

FIGURE 7.1.2 One dozen is 12 eggs, two dozen is 24 eggs and half a dozen is 6 eggs.

FIGURE 7.1.3 Convenient quantities that you would be very familiar with: (a) a pair of shoes equals two, (b) a dozen roses equals 12, (c) a ream of paper equals 500 sheets.

Information provided by molecular formulae

The **molecular formula** of a substance indicates the number of atoms of each element in one molecule of the substance. In Chapter 5, you learnt that an oxygen molecule contains two oxygen atoms joined by a covalent bond.

GO TO ➤ Section 5.3 page 162

When referring to a mole of a substance, it is important to indicate which type of particle is being specified. The expression, '1 mol of oxygen' is ambiguous because it could describe 1 mol of oxygen atoms (O) or 1 mol of oxygen molecules (O_2). As there are two atoms in each oxygen molecule, 1 mol of oxygen molecules will contain 2 mol of oxygen atoms.

Some other examples of the use of the mole as a counting unit are provided in Table 7.1.1.

TABLE 7.1.1 Examples of the use of the mole as a counting unit

Number of moles of element or compound	Information that can be obtained about numbers of particles
1 mol of hydrogen atoms (H)	1 mol of hydrogen atoms (H)
1 mol of hydrogen molecules (H_2)	1 mol of hydrogen molecules (H_2) 2 mol of hydrogen atoms (H)
2 mol of aluminium atoms (Al)	2 mol of aluminium atoms (Al)
2 mol of calcium chloride ($CaCl_2$)	2 mol of Ca^{2+} ions 4 mol of Cl^- ions
10 mol of glucose ($C_6H_{12}O_6$) molecules	10 mol of glucose ($C_6H_{12}O_6$) molecules 60 mol of carbon atoms 120 mol of hydrogen atoms 60 mol of oxygen atoms

AVOGADRO'S CONSTANT

One mole of any substance is defined as the same number of particles as there are atoms in exactly 12 g of carbon-12. This number of atoms has been experimentally determined to be $6.022\,141 \times 10^{23}$ atoms. That's 602 214 100 000 000 000 000 000 atoms!

This number is commonly rounded to 6.022×10^{23} and is referred to as **Avogadro's constant** or Avogadro's number. It is given the symbol N_A.

Avogadro's constant is written in scientific notation.

GO TO ➤ SkillBuilder page 80

FIGURE 7.1.4 One mole of sodium chloride and 1 mol of water contain the same number of particles, 6.022×10^{23}.

$N_A = 6.022 \times 10^{23}\ mol^{-1}$

Therefore, 1 mol of particles contains 6.022×10^{23} particles.

Avogadro's constant is an enormous number, but the extremely small size of atoms, ions and molecules means that 1 mol of most elements and compounds does not have a large mass or volume.

For example, in Figure 7.1.4 you can see that 1 mol of water, that is 6.022×10^{23} water molecules, has a volume of only 18 mL, and 1 mol of table salt (NaCl) has a mass of 58.44 g.

If you know that 1 mol of a substance contains 6.022×10^{23} particles, then:

- 2 mol of a substance contains $2 \times (6.022 \times 10^{23}) = 1.204 \times 10^{24}$ particles
- 0.3 mol of a substance contains $0.3 \times (6.022 \times 10^{23}) = 1.806 \times 10^{23}$ particles
- 4.70×10^{23} particles $= \frac{4.70 \times 10^{23}}{6.022 \times 10^{23}}$ particles $= 0.781$ mol
- 7.35×10^{24} particles $= \frac{7.35 \times 10^{24}}{6.022 \times 10^{23}}$ particles $= 12.2$ mol.

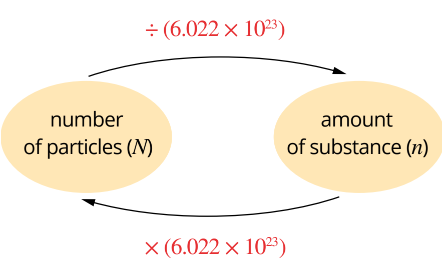

FIGURE 7.1.5 Relationship between number of particles and amount of substance in moles.

As you can see from Figure 7.1.5, a mathematical relationship exists between the number of particles, N, and the amount of substance in moles, n. This relationship can be written as:

$$n = \frac{N}{N_A}$$

Calculations using the mole and Avogadro's constant

Three quantities have been introduced so far:

- the mole, which is given the symbol n and the abbreviation mol
- Avogadro's constant, which is given the symbol N_A and has the value 6.022×10^{23}
- the actual number of particles (atoms, ions or molecules), which is given the symbol N.

The mathematical relationship that links the three quantities is $n = \frac{N}{N_A}$.

Worked example 7.1.1

CALCULATING THE NUMBER OF MOLECULES

Calculate the number of molecules in 3.5 mol of water (H_2O).

Thinking	Working
List the data given in the question next to the appropriate symbol. Include units.	The number of water molecules is the unknown, so: $N(H_2O) = ?$ $n(H_2O) = 3.5$ mol $N_A = 6.022 \times 10^{23}$
Rearrange the formula to make the unknown the subject.	$n = \frac{N}{N_A}$ so $N(H_2O) = n \times N_A$
Substitute in data and solve.	$N(H_2O) = n \times N_A$ $= 3.5 \times 6.022 \times 10^{23}$ $= 2.1 \times 10^{24}$ molecules

Worked example: Try yourself 7.1.1

CALCULATING THE NUMBER OF MOLECULES

Calculate the number of molecules in 1.6 mol of carbon dioxide (CO_2).

Worked example 7.1.2

CALCULATING THE NUMBER OF ATOMS

Calculate the number of oxygen atoms in 2.5 mol of oxygen gas (O_2).

Thinking	Working
List the data given in the question next to the appropriate symbol. Include units.	The number of oxygen atoms is the unknown, so: $N(O) = ?$ $n(O_2) = 2.5$ mol $N_A = 6.022 \times 10^{23}$
Calculate the amount, in mol, of oxygen atoms from the amount of oxygen and the molecular formula.	$n(O) = n(O_2) \times 2$ $= 2.5 \times 2$ $= 5.0$ mol
Rearrange the formula to make the unknown the subject.	$n = \frac{N}{N_A}$ so $N(O) = n \times N_A$
Substitute in data and solve.	$N(O) = n \times N_A$ $= 5.0 \times 6.022 \times 10^{23}$ $= 3.0 \times 10^{24}$ atoms

CHEMFILE N

Avogadro's constant

It is very difficult to imagine just how big Avogadro's constant really is, especially when atoms, ions and molecules are so small. Here are some examples to help.

- 6.022×10^{23} grains of sand, placed side by side, would stretch from Earth to the Sun and back about 7 million times.
- A computer counting 10 billion times every second would take 2 million years to reach 6.022×10^{23}.
- 6.022×10^{23} of the marshmallows shown in Figure 7.1.6 would cover Australia to a depth of 900 km!

FIGURE 7.1.6 One mole of (6.022×10^{23}) marshmallows would cover Australia to a depth of 900 km!

Worked example: Try yourself 7.1.2

CALCULATING THE NUMBER OF ATOMS

Calculate the number of hydrogen atoms in 0.35 mol of methane (CH_4).

Worked example 7.1.3

CALCULATING THE NUMBER OF MOLES, GIVEN THE NUMBER OF PARTICLES

Calculate the amount, in moles, of ammonia molecules (NH_3) represented by 2.5×10^{22} ammonia molecules.

Thinking	Working
List the data given in the question next to the appropriate symbol. Include units.	The number of mol of ammonia molecules is the unknown, so: $n(NH_3) = ?$ $N(NH_3) = 2.5 \times 10^{22}$ molecules $N_A = 6.022 \times 10^{23}$
Rearrange the formula to make the unknown the subject.	$n = \frac{N}{N_A}$ n is the unknown, so rearrangement is not required
Substitute in data and solve.	$n(NH_3) = \frac{N}{N_A}$ $= \frac{2.5 \times 10^{22}}{6.022 \times 10^{23}}$ $= 0.042$ mol

Worked example: Try yourself 7.1.3

CALCULATING THE NUMBER OF MOLES, GIVEN THE NUMBER OF PARTICLES

Calculate the amount, in mol, of magnesium atoms represented by 8.1×10^{20} magnesium atoms.

Worked example 7.1.4

CALCULATING THE NUMBER OF MOLES OF ATOMS, GIVEN THE NUMBER OF MOLES OF MOLECULES

Calculate the amount, in mol, of hydrogen atoms in 3.6 mol of sulfuric acid (H_2SO_4).

Thinking	Working
List the data given in the question next to the appropriate symbol. Include units.	The number of moles of hydrogen atoms is the unknown, so: $n(H) = ?$ $n(H_2SO_4) = 3.6$ mol
Calculate the number, in mol, of hydrogen atoms from the amount of sulfuric acid and the molecular formula.	$n(H) = n(H_2SO_4) \times 2$ $= 3.6 \times 2$ $= 7.2$ mol

Worked example: Try yourself 7.1.4

CALCULATING THE NUMBER OF MOLES OF ATOMS, GIVEN THE NUMBER OF MOLES OF MOLECULES

Calculate the amount, in mol, of hydrogen atoms in 0.75 mol of water (H_2O).

SKILLBUILDER N

Identifying significant figures

When giving an answer to a calculation, it is important to take note of the number of significant figures that you use.

You should give an answer that is as accurate as possible. However, an answer can't be more precise than the data or the measuring device used to calculate it. For example, if a set of scales that measures to the nearest gram shows that an object has a mass of 56 g, then the mass should be recorded as 56 g, not 56.0 g. This is because you do not know whether it is 56.0 g, 56.1 g, 56.2 g or 55.8 g.

The value '56' has 2 significant figures. Recording it with 3 significant figures (e.g. 56.0 g or 55.8 g) would not be scientifically 'honest'. If the mass of 56 g was used to calculate another value, it would also not be 'honest' to give that answer with more than 2 significant figures.

The number of significant figures required in an answer depends on what kind of calculation you are doing.

If you are multiplying or dividing, use the smallest number of significant figures provided in the initial values.

If you are adding or subtracting, use the smallest number of decimal places provided in the initial values.

Working out the number of significant figures

The following rules should be followed to avoid confusion in determining how many significant figures are in a number.

- All non-zero digits are always significant. For example, 21.7 has three significant figures.
- All zeroes between two non-zero digits are significant. For example, 3015 has four significant figures.
- A zero to the right of a decimal point and following a non-zero digit is significant. For example, 0.5700 has four significant figures.
- Any other zero in a number less than one is not significant, as it is used only for locating decimal places. For example, 0.005 has just one significant figure.

7.1 Review

SUMMARY

- A mole is a convenient quantity (unit) for counting particles. The mole is given the symbol *n* and the abbreviation mol.
- One mole is defined as the amount of substance that contains the same number of 'specified' particles as there are atoms in 12 g of carbon-12.
- The number of particles in 1 mol is given the symbol N_A. This is known as Avogadro's constant and has the numerical value of 6.022×10^{23}.
- The formula $n = \frac{N}{N_A}$ can be used or rearranged to calculate the amount or number of specified particles in a sample.

KEY QUESTIONS

1 Calculate the number of:

a atoms in 2.0 mol of sodium atoms (Na)
b molecules in 0.10 mol of nitrogen molecules (N_2)
c atoms in 20.0 mol of carbon atoms (C)
d molecules in 4.2 mol of water molecules (H_2O)
e atoms in 1.0×10^{-2} mol of iron atoms (Fe)
f molecules in 4.62×10^{-5} mol of CO_2 molecules.

2 Calculate the amount of substance, in mol, represented by:

a 3.0×10^{23} molecules of water (H_2O)
b 1.5×10^{23} atoms of neon (Ne)
c 4.2×10^{25} atoms of iron (Fe)
d 4.2×10^{25} molecules of ethanol (C_2H_5OH).

3 Calculate the amount, in mol, of:

a sodium atoms represented by 1.0×10^{20} sodium atoms
b aluminium atoms represented by 1.0×10^{20} aluminium atoms
c chlorine molecules represented by 1.0×10^{20} chlorine molecules.

4 Calculate the amount, in mol, of:

a chlorine atoms in 0.4 mol of chlorine (Cl_2)
b hydrogen atoms in 1.2 mol of methane (CH_4)
c hydrogen atoms in 0.12 mol of ethane (C_2H_6)
d oxygen atoms in 1.5 mol of sodium sulfate (Na_2SO_4).

7.2 Molar mass

Chemical laboratories always contain a balance like the one in Figure 7.2.1, which is used for weighing. If a chemist knows that a specific mass of a substance always contains a specific number of particles, it is possible to easily weigh a sample of the substance and calculate the exact number of particles present in the sample.

In this section, you will learn about how the amount of a substance, measured in moles, is related to the mass of the substance.

FIGURE 7.2.1 A digital balance is a simple piece of laboratory equipment used for weighing.

MOLAR MASS

Chemists have cleverly defined the mole so that you can determine the number of moles of a substance by simply measuring its mass.

The particles of different elements and compounds have different masses. Therefore, the masses of 1 mol of different elements or compounds will also be different. This is like saying that the mass of one dozen oranges will be greater than the mass of one dozen mandarins because one orange is heavier than one mandarin. The mass, in grams, of 1 mol of a particular element or compound is known as its **molar mass**. It is given the symbol M and the unit $g\,mol^{-1}$.

Remember that a mole is defined as the amount of substance that contains the same number of specified particles as there are atoms in 12 g of carbon-12. This is a very convenient definition because:

- 1 atom of ^{12}C has a **relative isotopic mass** (I_r) of 12 exactly
- 1 mol of atoms of ^{12}C has a mass of 12 g exactly.

Naturally occurring carbon is mainly composed of the ^{12}C isotope, so the molar mass of carbon is $12.01\,g\,mol^{-1}$.

Consider an atom of ^{12}C and an atom of ^{24}Mg. ^{12}C has been assigned a relative isotopic mass of 12 exactly. On that scale, the relative isotopic mass of ^{24}Mg is approximately 24. Since 1 mol of ^{12}C atoms weighs exactly 12 g, 1 mol of ^{24}Mg must weigh approximately twice as much, 24 g.

In general, the molar mass of an element is the mass of 1 mol of the element.

The molar mass of a compound is the mass of 1 mol of the compound. It is equal to the relative molecular or **relative formula mass** of the compound expressed in grams.

The molar mass is given the symbol, M, and the unit $g\,mol^{-1}$.

Table 7.2.1 shows you how to calculate the molar masses of some common substances.

TABLE 7.2.1 Calculating the molar mass of a substance by adding the relative atomic masses for each atom present in the substance based on the molecular or ionic formula

Substance	Relative atomic masses	Molar mass of substance
Na	Na: 22.99	$= 22.99\,g\,mol^{-1}$
O_2	O: 16.00	$= 2 \times 16.00$ $= 32.00\,g\,mol^{-1}$
H_2O	H: 1.008 O: 16.00	$= (2 \times 1.008) + 16.00$ $= 18.02\,g\,mol^{-1}$
CO_2	C: 12.01 O: 16.00	$= 12.01 + (2 \times 16.00)$ $= 44.01\,g\,mol^{-1}$
$NaNO_3$	Na: 22.99 N: 14.01 O: 16.00	$= 22.99 + 14.01 + (3 \times 16.00)$ $= 85.00\,g\,mol^{-1}$

From the calculations in Table 7.2.1 and the photograph of 1 mol of some common substances in Figure 7.2.2, you can see that 1 mol of each substance has a different mass.

FIGURE 7.2.2 One mole of each substance has a different mass.

Counting by weighing

A useful relationship links the amount of a substance (n) in moles, its molar mass (M) in grams per mole, and the given mass of the substance (m) in grams.

Mass of a given amount of substance (g) = amount of substance (mol) × molar mass ($g\,mol^{-1}$).

This can be written as $m = n \times M$ and rearranged to:

amount in mol — $n = \frac{m}{M}$

mass in g

molar mass in g mol^{-1}

Worked example 7.2.1

CALCULATING THE MASS OF A SUBSTANCE

Calculate the mass of 0.35 mol of magnesium nitrate ($Mg(NO_3)_2$).

Thinking	Working
List the data given to you in the question. Remember that whenever you are given a molecular formula, you can calculate the molar mass.	$m(Mg(NO_3)_2)$ = ? $n(Mg(NO_3)_2)$ = 0.35 mol $M(Mg(NO_3)_2)$ = 24.31 + (2 × 14.01) + (6 × 16.00) = 148.33 $g\,mol^{-1}$
Calculate the mass of magnesium nitrate using: $n = \frac{m}{M}$	$n = \frac{m}{M}$, so $m = n \times M$ $m(Mg(NO_3)_2)$ = 0.35 × 148.33 = 52 g

Worked example: Try yourself 7.2.1

CALCULATING THE MASS OF A SUBSTANCE

Calculate the mass of 4.68 mol of sodium carbonate (Na_2CO_3).

You are now in a position to count atoms by weighing. When you use the mole, you are effectively counting the number of particles in a substance. The number of particles present in a substance is equal to the number of moles of the substance multiplied by 6.022×10^{23}.

Some calculations require you to use both of the formulae $n = \frac{m}{M}$ and $n = \frac{N}{N_A}$. Worked example 7.2.2 is such a calculation.

Worked example 7.2.2

CALCULATING THE NUMBER OF MOLECULES

Calculate the number of CO_2 molecules present in 22 g of carbon dioxide.

Thinking	Working
List the data given to you in the question. Convert mass to grams if required. Remember that whenever you are given a molecular formula you can calculate the molar mass.	$N(CO_2) = ?$ $M(CO_2) = 12.01 + (2 \times 16.00)$ $= 44.01\ g\ mol^{-1}$ $m(CO_2) = 22\ g$
Calculate the amount, in mol, of CO_2 using: $n = \frac{m}{M}$	$n(CO_2) = \frac{m}{M}$ $= \frac{22}{44.01}$ $= 0.50\ mol$
Calculate the number of CO_2 molecules using: $n = \frac{N}{N_A}$	$n = \frac{N}{N_A}$, so $N = n \times N_A$ $N(CO_2) = 0.50 \times 6.022 \times 10^{23}$ $= 3.0 \times 10^{23}$ molecules

Worked example: Try yourself 7.2.2

CALCULATING THE NUMBER OF MOLECULES

Calculate the number of sucrose molecules in a teaspoon (4.2 g) of sucrose ($C_{12}H_{22}O_{11}$).

WS 2.3

CHEMISTRY IN ACTION N

Big numbers in science and engineering

Astronomers and engineers also deal with very large numbers and have developed units to handle the huge distances and volumes they deal with. An astronomical unit (AU) is equal to the mean distance between the Sun and Earth, and it is used to compare the size of other planetary systems with our own. Distances between stars are measured in light-years: a light-year is the distance that light travels in a vacuum in 365.25 days. A more unusual Australian unit of measurement is a Sydharb, which is a volume defined in the Macquarie Dictionary as equal to the amount of water in Sydney Harbour at high tide (Figure 7.2.3).

News reports often discuss the size of a flood or the capacity of a reservoir in multiples of Sydney Harbour, and one Sydharb is approximately 5×10^{11} L. An astronomical unit is 1.49×10^{11} m and a light-year is 9.46×10^{15} m. In the same way that it is easier to talk about a mole than 6.022×10^{23} particles, these units are easy ways to visualise and discuss large numbers. These units also put the size of a mole into perspective. Even the numerical value of a light-year is 100 000 000 times smaller than a mole!

FIGURE 7.2.3 A Sydharb is an unusual unit of volume equal to the amount of water in Sydney Harbour at high tide. Floods and reservoir capacity are often measured in Sydharbs.

CHEMISTRY IN ACTION N

Measuring a molar mass of a compound without a mass spectrometer

When chemists were first determining the molar masses of molecules in the 19th century, a common method was to dissolve the molecule in a solvent and measure the decrease in the solvent's freezing point. This is called a freezing point depression experiment. It was a successful way to measure a molar mass because the freezing point of a solution depends on the number of particles dissolved in the solution, but not on their identity.

When water is the solvent, this relationship is written as:

T_f(solution) × m(water) = −1.86 n

where:

T_f(solution) is the freezing point of the solution, in °C

m(water) is the mass of water, in kg

n is the number of moles of the dissolved substance.

Therefore, if you know the mass of the substance, you can use the formula $n = \frac{m}{M}$ to work out its molar mass.

7.2 Review

SUMMARY

- The molar mass of an element or compound is the mass, in grams, of 1 mol of that element or compound. Molar mass is given the symbol M and the unit g mol^{-1}.
- The molar mass of an element or compound has the same numerical value as the relative mass of the element or compound.
- The formula $n = \frac{m}{M}$ can be used or rearranged to calculate the mass, amount or molar mass of an element or compound.

KEY QUESTIONS

1 Calculate the molar mass of:
  - **a** nitrogen (N_2)
  - **b** ammonia (NH_3)
  - **c** sulfuric acid (H_2SO_4)
  - **d** iron(III) nitrate ($Fe(NO_3)_3$)
  - **e** acetic acid (CH_3COOH)
  - **f** sulfur atoms (S)
  - **g** vitamin C (ascorbic acid, $C_6H_8O_6$)
  - **h** hydrated copper(II) sulfate ($CuSO_4{\bullet}5H_2O$).

2 Calculate the mass of:
  - **a** 1.0 mol of sodium atoms (Na)
  - **b** 2.0 mol of oxygen molecules (O_2)
  - **c** 0.10 mol of methane molecules (CH_4)
  - **d** 0.25 mol of aluminium oxide (Al_2O_3).

3 Calculate the amount, in mol, of:
  - **a** H_2 molecules in 5.0 g of hydrogen (H_2)
  - **b** H atoms in 5.0 g of hydrogen (H_2)
  - **c** Al atoms in 2.7 g of aluminium (Al)
  - **d** CH_4 molecules in 0.40 g of methane (CH_4)
  - **e** O_2 molecules in 0.10 g of oxygen (O_2)
  - **f** O atoms in 0.10 g of oxygen (O_2)
  - **g** P_4 molecules in 1.2×10^{-3} g of phosphorus (P_4)
  - **h** P atoms in 1.2×10^{-3} g of phosphorus (P_4).

4 Calculate the number of atoms in:
  - **a** 23 g of sodium (Na)
  - **b** 4.0 g of argon (Ar)
  - **c** 0.243 g of magnesium (Mg)
  - **d** 10.0 g of gold (Au).

5 Calculate the:
  - **a** number of molecules in:
    - **i** 16 g of oxygen (O_2)
    - **ii** 2.8 g of nitrogen (N_2)
  - **b** number of oxygen atoms in 3.2 g of sulfur dioxide (SO_2)
  - **c** total number of atoms in 288 g of ammonia (NH_3).

7.3 Percentage composition and empirical formula

In Chapter 2, you learnt that compounds are substances that contain two or more different elements, and that the relative proportions of each element in a compound can be expressed as:

- a percentage in terms of the mass contributed by each element
- a formula, showing either the ratio of atoms contributed by each element in a compound or the actual numbers of atoms in a molecule.

In this section, you will learn how to use the **percentage composition** of elements in a compound to calculate a type of formula called an **empirical formula**.

PERCENTAGE COMPOSITION

In Chapter 2, you were introduced to the idea of the percentage composition of a given compound, which tells you the proportion by mass of the different elements in that compound. The proportion of each element is expressed as a percentage of the total mass of the compound.

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If the chemical formula of a compound is known, the percentage composition can be determined using the molar masses of the elements and compound. In general, this is written as:

$$\%\text{ by mass of an element in a compound} = \frac{\text{mass of the element in 1 mol of the compound}}{\text{molar mass of the compound}} \times 100$$

Worked example 7.3.1

CALCULATING PERCENTAGE COMPOSITION

Calculate the percentage by mass of aluminium in alumina (Al_2O_3).

Thinking	Working
Find the molar mass of the compound.	$M(Al_2O_3) = (2 \times 26.98) + (3 \times 16.00)$ $= 101.96\text{ g mol}^{-1}$
Find the total mass of the element in one mol of the compound.	mass of Al in 1 mol $= 2 \times M(\text{Al})$ $= 2 \times 26.98$ $= 53.96\text{ g}$
Find the percentage by mass of the element in the compound.	% by mass of Al in Al_2O_3 $= \frac{\text{mass of Al in 1 mol of } Al_2O_3}{\text{molar mass of } Al_2O_3} \times 100$ $= \frac{53.96}{101.96} \times 100$ $= 52.92\%$

Worked example: Try yourself 7.3.1

CALCULATING PERCENTAGE COMPOSITION

Calculate the percentage by mass of nitrogen in ammonium nitrate (NH_4NO_3).

EMPIRICAL FORMULA

Atoms or ions are present in compounds in fixed whole number ratios. The empirical formula of a compound gives the simplest whole number ratio of elements in that compound. See Table 7.3.1 for some examples.

TABLE 7.3.1 Empirical formulae of some common compounds

Compound	Empirical formula	Simplest whole number ratio of elements in the compound
water	H_2O	H:O 2:1
ethene	CH_2	C:H 1:2
calcium carbonate	$CaCO_3$	Ca:C:O 1:1:3

Determining empirical formulae

The empirical formula for a compound is determined from the mass of each element present in a given mass of the compound. These masses can be determined experimentally.

Once the masses of elements in a compound are known, the steps in Figure 7.3.1 are followed to convert these masses into a **mole ratio** (a ratio by number of atoms) and then into an empirical formula.

Obtain the mass (m) of each element in a compound. → Calculate the amount, in mol, of each element present using the formula $n = \frac{m}{M}$. → Convert the mol of each element calculated in the previous step to a whole number ratio. → Write the empirical formula.

FIGURE 7.3.1 Follow these steps to calculate an empirical formula.

Worked example 7.3.2

DETERMINING THE EMPIRICAL FORMULA

A compound of carbon and oxygen contains 27.3% carbon and 72.7% oxygen by mass. Calculate the empirical formula of the compound.

Thinking	Working
Write down the mass, in g, of all elements present in the compound. If masses are given as percentages, assume that the sample weighs 100 g. The percentages then become masses in grams.	$m(\text{C}) = 27.3\text{ g}$ $\quad m(\text{O}) = 72.7\text{ g}$
Calculate the amount, in mol, of each element in the compound using: $n = \frac{m}{M}$	$n(\text{C}) = \frac{27.3}{12.01} = 2.28\text{ mol}$ $\quad n(\text{O}) = \frac{72.7}{16.00} = 4.54\text{ mol}$
Simplify the ratio by dividing each number of mol by the smallest number of mol calculated in the previous step. This gives you a ratio of the number of atoms of each element.	$\frac{2.28}{2.28} = 1$ $\quad \frac{4.54}{2.28} = 2$
Find the simplest whole number ratio.	1:2
Write the empirical formula.	CO_2

Worked example: Try yourself 7.3.2

DETERMINING THE EMPIRICAL FORMULA

0.50 g of magnesium is heated and allowed to completely react with chlorine. 1.96 g of white powder is formed. Determine the empirical formula of the compound.

MOLECULAR FORMULA

Molecular compounds have a molecular formula in addition to an empirical formula. The molecular formula gives the actual number of atoms of each element present in a molecule, rather than the simplest whole number ratio.

The molecular formula can be the same as or different from the empirical formula.

The empirical and molecular formulae of some common molecular compounds are shown in Table 7.3.2.

TABLE 7.3.2 Empirical and molecular formulae of some common molecular compounds

Molecule	Molecular formula	Empirical formula
water	H_2O	H_2O
ethane	C_2H_6	CH_3
carbon dioxide	CO_2	CO_2
glucose	$C_6H_{12}O_2$	CH_2O

Ionic compounds do not have molecular formulae because they do not exist as molecules. However, they do have empirical formulae that describe the fixed ratio of ions that exist in their lattices. The formula for calcium chloride ($CaCl_2$) is an example of an empirical formula of an ionic compound.

Determining molecular formulae

A molecular formula can be determined from the empirical formula of a compound if the molar mass of the compound is also known.

The molecular formula of a molecule is always a whole number multiple of the empirical formula. The number of the multiple is determined by the following formula.

$$\text{Number of empirical formula units in a molecule} = \frac{\text{molar mass of the compound}}{\text{molar mass of one empirical formula unit}}$$

The general steps in the determination of a molecular formula are shown in Figure 7.3.2.

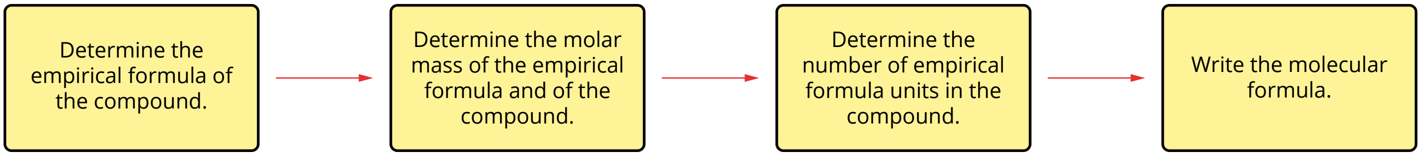


FIGURE 7.3.2 Steps for calculating a molecular formula.

Worked example 7.3.3

DETERMINING THE MOLECULAR FORMULA

A compound has the empirical formula CH. The molar mass of this compound is 78.11 g mol^{-1}. What is the molecular formula of the compound?

Thinking	Working
Calculate the molar mass of one unit of the empirical formula.	Molar mass of a CH unit = 12.01 + 1.008 = 13.02 g mol^{-1}
Determine the number of empirical formula units in the molecular formula.	number of CH units = $\frac{78.11}{13.02}$ = 6
Determine the molecular formula of the compound.	molecular formula = 6 × CH = C_6H_6

Worked example: Try yourself 7.3.3

PA 2.1 PA 2.1 PA 2.2 PA 2.2

DETERMINING THE MOLECULAR FORMULA

A compound has the empirical formula C_2H_5. The molar mass of this compound was determined to be 58.12 g mol^{-1}. What is the molecular formula of the compound?

7.3 Review

SUMMARY

- The percentage, by mass, of an element in a compound can be calculated from the mass of the element in 1 mol of the compound and the molar mass of the compound.
- The empirical formula of a compound gives the simplest whole number ratio of atoms or ions in the compound.
- Ionic compounds only have an empirical formula.
- Molecular compounds have a molecular formula that gives the actual number of atoms of each element in the molecule. It may be the same as, or different from, the empirical formula.

KEY QUESTIONS

1 Calculate the percentage by mass of:

a iron in iron(III) oxide (Fe_2O_3)

b uranium in uranium oxide (U_3O_8)

c nitrogen in ammonium chloride (NH_4Cl)

d oxygen in copper(II) nitrate ($Cu(NO_3)_2$).

2 Determine the empirical formulae of the compounds with the following compositions:

a 2.74% hydrogen, 97.26% chlorine

b 42.9% carbon, 57.1% oxygen

c 10.0 g of a compound of magnesium and oxygen that contains 6.03 g of magnesium

d 3.2 g of a hydrocarbon that contains 2.4 g of carbon

3 Determine the molecular formula of the following compounds.

	Empirical formula	Relative molecular mass
a	CH	78.11
b	HO	34.02
c	CH_2O	90.09
d	NO_2	46.01
e	CH_2	154.29

4 A hydrocarbon contains 85.7% carbon. Its relative molecular mass is 70. Determine the hydrocarbon's:

a empirical formula

b molecular formula.

5 A sample of the carbohydrate glucose contains 1.8 g carbon, 0.3 g hydrogen and 2.4 g oxygen.

a Calculate the empirical formula of the compound.

b Deduce its molecular formula, given that its relative molecular mass is 180.

7.4 Calculations based on the amount of a reactant or product

In this section, you will learn to calculate the amounts of reactants or products involved in chemical reactions. For example, in reactions such as that seen in Figure 7.4.1, you can use your knowledge of the mole concept to calculate the mass of products formed or reactants consumed.

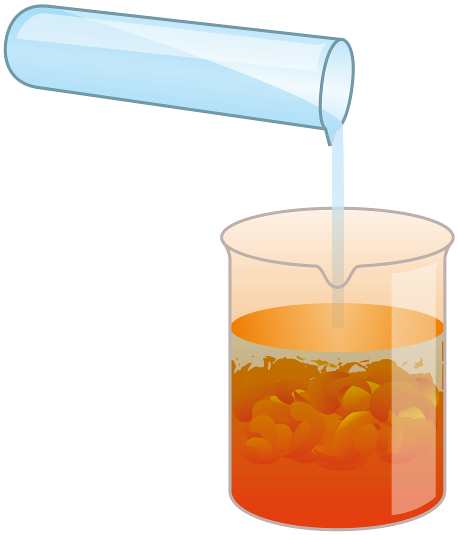


FIGURE 7.4.1 When colourless aqueous solutions of mercury(II) acetate and sodium iodide are mixed, they produce a red solid, mercury(II) iodide. The mass of the solid from this reaction can be used to determine the amount of mercury present in the original sample.

STOICHIOMETRY

Chemists can apply the mole concept to chemical reactions to determine the quantities of reactants involved or products formed. This is useful in industry, as manufacturers need to know the quantities of raw materials required to produce a predicted amount of product. Although particles such as atoms, ions and molecules are so small that they cannot be counted individually, the mole concept allows us to determine the number of individual particles by simply measuring their mass.

Calculations based on the mole are used to make these predictions.

The study of ratios of moles of substances is called **stoichiometry**. **Stoichiometric calculations** are based on the law of conservation of mass.

The total mass of all products is equal to the total mass of all reactants in a chemical reaction.

In a chemical reaction, the total number of atoms of each element in the reactants is always equal to the total number of atoms of each element in the products. The atoms are just rearranged to form new compounds.

Mole ratios

Consider one of the important reactions in breadmaking, where table sugar (or sucrose, $C_{12}H_{22}O_{11}$) is fermented by baker's yeast to make carbon dioxide and ethanol (C_2H_5OH). The carbon dioxide makes the bread rise, and the ethanol evaporates in the baking process. The chemical equation is:

$$C_{12}H_{22}O_{11}(aq) + H_2O(l) \rightarrow 4CO_2(g) + 4C_2H_5OH(l)$$

The coefficients used to balance the equations show the ratios of reactants and products involved in the reaction. The equation indicates that 1 mol of $C_{12}H_{22}O_{11}$ reacts with 1 mol of water to form 4 mol of gaseous CO_2 and 4 mol of C_2H_5OH.

In more general terms, the amount, in moles, of CO_2 produced will always be four times the amount, in moles, of $C_{12}H_{22}O_{11}$ used. The amount, in moles, of C_2H_5OH produced will be equal to the amount, in moles, of CO_2 produced and four times the amount, in moles, of water consumed.

Stoichiometric calculations allow you to use the mole ratio established in a chemical equation to predict the amount of a product that will be formed and how much reactant will be used.

You can use the coefficients of this reaction to write relationships that show the mole ratios of any two chemicals involved in the reaction:

$$\frac{n(C_{12}H_{22}O_{11})}{n(H_2O)} = \frac{1}{1}, \ \frac{n(CO_2)}{n(C_{12}H_{22}O_{11})} = \frac{4}{1} \text{ and } \frac{n(CO_2)}{n(C_2H_5OH)} = \frac{4}{4} = \frac{1}{1}$$

In the stoichiometric calculations that you will perform in this section, the number of mole or mass of one of the reactants or products will always be known (called the 'known chemical') and used to determine the number of mole or mass of one of the other reactants or products involved in the reaction (called the 'unknown chemical'). You can write the relationship between the known and the unknown chemicals using ratios:

$$\frac{n(\text{unknown chemical})}{n(\text{known chemical})} = \frac{\text{coefficient of unknown chemical}}{\text{coefficient of known chemical}}$$

When carrying out any stoichiometric calculation, you must always clearly state the mole ratio you are working with.

Worked example 7.4.1

USING MOLE RATIOS

0.375 mol of methanol (CH_3OH) is burnt in oxygen to make carbon dioxide and water. How many moles of carbon dioxide are formed?

Thinking	Working
Write a balanced equation for the reaction.	$CH_3OH(l) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$
Determine the number of mol of the 'known substance'. The known substance is the one that you are provided information about in the question.	$n(CH_3OH) = 0.375\ mol$
Find the mole ratio: $$\text{mole ratio} = \frac{\text{coefficient of unknown chemical}}{\text{coefficient of known chemical}}$$ The 'unknown' substance is the one for which you need to calculate the number of mol	$\frac{\text{coefficient of } CO_2}{\text{coefficient of } CH_3OH} = \frac{1}{1}$
Calculate the number of mol of the unknown substance using: $n(\text{unknown}) = n(\text{known}) \times (\text{mole ratio})$	$n(CO_2) = 0.375 \times 1$ $= 0.375\ mol$

Worked example: Try yourself 7.4.1

USING MOLE RATIOS

In the breadmaking reaction, baker's yeast ferments table sugar (sucrose, $C_{12}H_{22}O_{11}$) to form carbon dioxide and ethanol (C_2H_5OH). 0.050 mol of CO_2 is needed to make a loaf of bread rise properly. If the chemical reaction is $C_{12}H_{22}O_{11}(aq) + H_2O(l) \rightarrow 4CO_2(g) + 4CH_3CH_2OH(l)$, how much sucrose, in mol, is needed to form this amount of CO_2?

Mass–mass stoichiometry

When you carry out a reaction in the laboratory, you will measure quantities of chemicals in grams, not moles. For this reason, most calculations will require you to start and finish with mass rather than moles of a substance. To calculate the number of moles of the known substance from a mass, you can use the relationship:

$$\text{number of moles} = \frac{\text{mass (in g)}}{\text{molar mass (in g mol}^{-1})}$$

This can be written as:

$$n = \frac{m}{M}$$

To calculate a final answer as a mass, this formula is rearranged:

$$m = n \times M$$

Calculating the mass of a product

Stoichiometry can be used to calculate the mass of product that is expected from any reaction in which the mass of a reactant is known.

There are several steps involved in calculating the mass of a product.

1 Write a balanced equation for the reaction.

2 Calculate the number of moles of the reactant from its mass, using the formula:

$$n = \frac{m}{M}$$

3 Use the mole ratios in the equation to calculate the number of moles of the expected product.

4 Calculate the mass of the product using $m = n \times M$.

FIGURE 7.4.3 Iron metal is formed from iron(III) oxide reacting with carbon. Manufacturers use stoichiometry to calculate how much iron(III) oxide and carbon are needed to produce enough iron for their next order.

Figure 7.4.2 provides a flow chart that summarises this process, and Worked example 7.4.2 will help you to understand these steps.

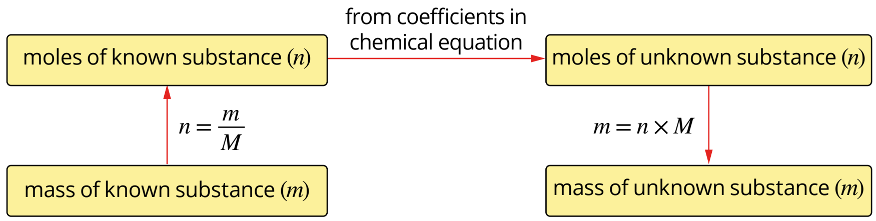


FIGURE 7.4.2 A flow chart for mass–mass stoichiometric calculations is helpful when trying to solve these problems.

Worked example 7.4.2

SOLVING MASS–MASS STOICHIOMETRIC PROBLEMS

When iron(III) oxide is reacted with carbon, iron metal is formed, along with carbon dioxide (Figure 7.4.3). Calculate the mass of iron that can be formed from 5.20 kg of iron(III) oxide.

Thinking	Working
Write a balanced equation for the reaction.	$2Fe_2O_3(s) + 3C(s) \rightarrow 4Fe(s) + 3CO_2(g)$
Calculate the number of mol of the known substance (the iron(III) oxide). Convert the mass to grams: $n = \frac{m}{M}$	$n(Fe_2O_3) = \frac{5.20 \times 10^3}{159.70}$ $= 32.6$ mol
Calculate the mole ratio: mole ratio $= \frac{\text{coefficient of unknown chemical}}{\text{coefficient of known chemical}}$	mole ratio $= \frac{\text{coefficient of Fe}}{\text{coefficient of } Fe_2O_3}$ $= \frac{4}{2}$
Calculate the number of mol of the unknown substance: $n(\text{unknown}) = n(\text{known}) \times (\text{mole ratio})$	$n(Fe) = 32.6 \times \frac{4}{2}$ $= 65.1$ mol
Calculate the mass of the unknown substance: $m = n(\text{unknown}) \times M$	$m(Fe) = 65.1 \times 55.85$ $= 3.64 \times 10^3$ g $= 3.64$ kg

Worked example: Try yourself 7.4.2

SOLVING MASS–MASS STOICHIOMETRIC PROBLEMS

PA 2.3 PA 2.3

A reaction between solutions of sodium sulfate and barium nitrate produces solid barium sulfate with a mass of 2.440 g. Aqueous sodium nitrate is also formed. Calculate the mass of sodium sulfate required to produce this amount of barium sulfate. The equation for this reaction is:

$$Na_2SO_4(aq) + Ba(NO_3)_2(aq) \rightarrow BaSO_4(s) + 2NaNO_3(aq)$$

7.4 Review

SUMMARY

- A balanced equation shows the ratio of the amount, in moles, of reactants used and products formed in the reaction.
- Given the quantity of one of the reactants or products of a chemical reaction, such as in a precipitation reaction, the quantity of all other reactants and products can be predicted by working through the following steps.
  1 Write a balanced equation for the reaction.
  2 Calculate the amount, in moles, of the given substance.
  3 Use the mole ratios of reactants and products in the balanced equation to calculate the amount, in moles, of the required substance.
  4 Use the appropriate formula to determine the quantity needed of the required substance. Usually, this would be the mass, using the formula $m = n \times M$.

KEY QUESTIONS

1 A tree undergoing photosynthesis uses 120 g of carbon dioxide to make glucose ($C_6H_{12}O_6$). The equation for the reaction is:

$$6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$$

What mass of glucose is formed?

2 The mass of aluminium nitrate in a solution is determined by adding sodium carbonate solution to form aluminium carbonate and sodium nitrate. The equation for the reaction occurring is:

$$2Al(NO_3)_3(aq) + 3Na_2CO_3(aq) \rightarrow Al_2(CO_3)_3(s) + 6NaNO_3(aq)$$

In a particular reaction, 4.68 g of aluminium carbonate is obtained.

a Calculate the moles of aluminium carbonate produced.
b Determine the required mole ratio for the reaction.
c Calculate the mass of aluminium nitrate that reacted.

3 A solar panel is made from 45.9 g of silicon. What mass of silicon tetrachloride is needed to make this mass of silicon? The equation for the reaction is:

$$SiCl_4(g) + 2H_2(g) \rightarrow Si(s) + 4HCl(g)$$

4 The reaction between mercury(II) acetate and sodium iodide is represented by the following equation:

$$Hg(CH_3COO)_2(aq) + 2NaI(aq) \rightarrow HgI_2(s) + 2NaCH_3COO(aq)$$

Solid mercury(II) iodide of mass 4.82 g is formed when sodium iodide is added to a solution of mercury(II) acetate:

$M(Hg(CH_3COO)_2) = 318.69\ g\ mol^{-1}$

$M(HgI_2) = 454.4\ g\ mol^{-1}$

Calculate the mass of mercury(II) acetate that reacted to produce this solid.

7.5 Calculations based on the amounts of two reactants

In the previous section, you calculated the amount of product that could be formed given the amount of one reactant. You were also able to calculate the amount of a second reactant that would be required, given the amount of the first. In chemistry, once you know the mass of a reactant that is completely consumed in a chemical reaction, you can use that information to determine the amount of any other component based on the chemical equation.

In this section, you will learn how to perform stoichiometric calculations in which two reactants are involved, but the reactant that is completely consumed may not be obvious straight away. In this style of question, you are given sufficient information to allow you to calculate the amounts of both reactants present. But before you can calculate the amount of product that is formed, you will need to determine which reactant is completely consumed. You will then know which reactant will be the limiting factor in the reaction.

STOICHIOMETRY PROBLEMS INVOLVING EXCESS REACTANTS

When two reactants are mixed to create a chemical reaction, it is possible that they will be combined in just the right mole ratio, as indicated in the equation, for each to react completely. However, it is more likely that they are not present in exactly the right mole ratio. In that case, one of the reactants will be used up before the other, and some of the other reactant will be left over at the end of the reaction.

To illustrate this idea, consider a problem that does not involve chemicals. Suppose that you have been given some skateboard decks and wheels, and you want to make as many complete skateboards as you can. As you can see in Figure 7.5.1, a complete skateboard is made up of one deck and four wheels.

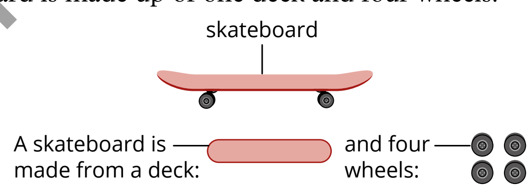


FIGURE 7.5.1 In order to make a complete skateboard, you must always use one deck and four wheels. This sets up the basic formula for a skateboard.

Now consider the situation shown in Figure 7.5.2. If you were given two decks and ten wheels, how many complete skateboards could you make from these materials?

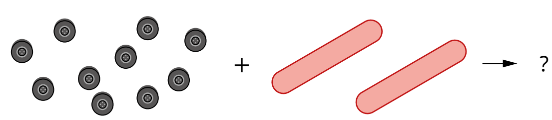


FIGURE 7.5.2 When provided with ten wheels and two skateboard decks, how many complete skateboards can be made?

The answer is that you could make two complete skateboards and there would be two wheels left over (Figure 7.5.3).

FIGURE 7.5.3 When supplied with two decks and ten wheels, the maximum number of skateboards that can be made is two. There will be two wheels that are not used.

In this example, you can say that the two wheels left at the end were in excess. Also, given two decks and ten wheels, the number of complete skateboards you could make was limited by the number of decks available.

A similar situation arises in chemical reactions when the quantities of the reactants supplied are not in the exact same ratio as that shown in the equation for the reaction.

When that happens, the reactant that is:

- completely consumed is the **limiting reactant**
- not completely consumed is the **excess reactant**.

Figure 7.5.4 shows three different scenarios for the reaction in which hydrogen gas and oxygen gas combine to form water, according to the equation:

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

Each of the diagrams provides examples to illustrate the concepts of limiting and excess reactants.

> In a chemical reaction, the limiting reactant is completely used up.

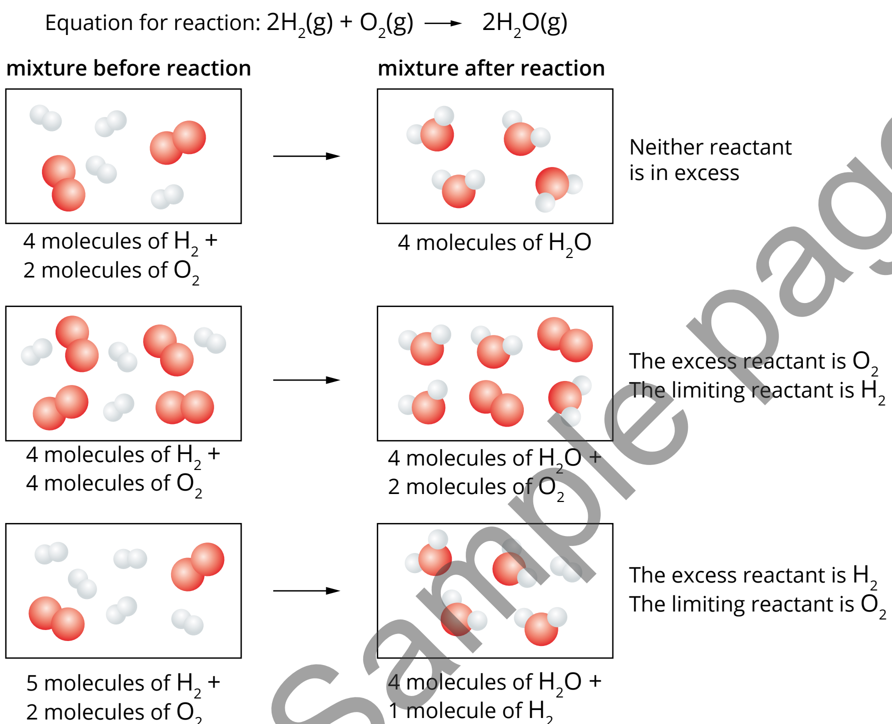


FIGURE 7.5.4 Different scenarios showing the concept of a limiting reactant for the reaction $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$.

Note that in each of the examples shown in Figure 7.5.4, the amount of product formed in these types of reactions:

- is determined by the amount of the limiting reactant present in the reaction mixture
- cannot be determined from the amount of excess reactant.

In the skateboard example, it was the number of decks, not the number of wheels, that determined how many complete skateboards could be made.

> You must always use the amount of the limiting reactant to determine the amount of product that will be formed.

Steps in solving stoichiometry problems involving excess reactants

When attempting to solve a limiting reactant problem in which you are required to work out the amount of product, there are three main steps:

1. Calculate the number of moles of each reactant.
2. Determine which reactant is in excess, and therefore which is the limiting reactant.
3. Use the amount of the limiting reactant to work out the amount of product formed.

Worked example 7.5.1

SOLVING MASS–MASS STOICHIOMETRY PROBLEMS WITH ONE REACTANT IN EXCESS

A solution containing 20.0 g of dissolved sodium hydroxide is added to a solution containing 25.0 g aluminium nitrate to form solid aluminium hydroxide and aqueous sodium nitrate. The equation for this reaction is:

$$3NaOH(aq) + Al(NO_3)_3(aq) \rightarrow Al(OH)_3(s) + 3NaNO_3(aq)$$

a Which reactant is the limiting reactant?	
Thinking	**Working**
Calculate the number of mol of each of the reactants in the equation using: $n = \frac{m}{M}$	Use the equation $n = \frac{m}{M}$. For NaOH: $n(NaOH) = \frac{20.0}{40.00} = 0.500$ mol For $Al(NO_3)_3$: $n(Al(NO_3)_3) = \frac{25.0}{213.01} = 0.117$ mol
Use the coefficients of the equation to find the limiting reactant.	The equation shows that 3 mol of NaOH react with 1 mol of $Al(NO_3)_3$. So to react all of the $Al(NO_3)$ you will require: $\frac{3}{1} \times n(Al(NO_3)_3$ of NaOH $\frac{3}{1} \times 0.117 = 0.352$ mol As there is 0.500 mol available, the NaOH is in excess. So, $Al(NO_3)_3$ is the limiting reactant. (It will be completely consumed.)

b What mass of aluminium hydroxide will be formed?	
Thinking	**Working**
Find the mole ratio of the unknown substance to the limiting reactant from the equation coefficients: mole ratio $= \frac{\text{coefficient of unknown chemical}}{\text{coefficient of known chemical}}$	From the equation coefficients: $\frac{\text{coefficient of } Al(OH)_3}{\text{coefficient of } Al(NO_3)_3} = \frac{1}{1}$
Calculate the number of mol of the unknown substance using the number of mol of the limiting reactant: n(unknown) = n(limiting reactant) × (mole ratio)	$n(Al(OH)_3) = n(Al(NO_3)_3) \times \frac{1}{1}$ $= 0.117 \times \frac{1}{1}$ $= 0.117$ mol
Calculate the mass of the unknown substance using: m(unknown) = n(unknown) × M	Molar mass of $Al(OH)_3 = 78.00\ g\ mol^{-1}$ $m(Al(OH)_3) = 0.117 \times 78.00$ $= 9.15$ g

Worked example: Try yourself 7.5.1

SOLVING MASS–MASS STOICHIOMETRY PROBLEMS WITH ONE REACTANT IN EXCESS

A space shuttle is loaded with 1.06×10^8 g of liquefied H_2 and 6.29×10^8 g of liquefied O_2, which are burnt as fuels during take-off. The reaction is:

$$2H_2(l) + O_2(l) \rightarrow 2H_2O(g)$$

WS 2.4

a Which reactant is the limiting reactant?

b What mass of water will be formed?

7.5 Review

SUMMARY

- If quantities of more than one reactant are given, the amount, in moles, of each reactant needs to be calculated.
- The limiting reactant needs to be determined; this will be the reactant that is consumed completely.
- The limiting reactant is used to predict the amount of product formed and the amount of the other reactant in excess.

KEY QUESTIONS

1 Aluminium reacts with oxygen gas to form aluminium oxide. The balanced equation for the reaction is:

$$4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$$

In a particular reaction, 40 g of aluminium reacts with 35 g of oxygen gas. List the following steps in the order in which the calculations should be completed to determine the mass of aluminium oxide that will form.

A Use mole ratios to determine which reactant is limiting.
B Calculate the number of mol of aluminium and oxygen.
C Calculate the mass of aluminium oxide that forms.
D Refer to the balanced equation.
E Calculate the number of mol of aluminium oxide that forms.

2 In three separate experiments, different amounts of nitrogen and hydrogen reacted to form ammonia, according to the equation:

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

This table shows the amounts of reactants and products in each experiment. Complete the table to indicate the amount of each product remaining at the end of the reaction.

Nitrogen molecules available	Hydrogen molecules available	Ammonia molecules produced	Nitrogen molecules in excess	Hydrogen molecules in excess
2	10			
879	477			
9 mol	6 mol			

3 Sodium metal can react with chlorine gas to form sodium chloride. 25.0 g of sodium is reacted with 50.0 g of chlorine gas.

a Write the balanced equation for the reaction between sodium and chlorine.
b Calculate the mass of sodium chloride that will form in the reaction.

4 Potassium iodide and lead(II) nitrate solutions react together to form solid lead(II) iodide and aqueous potassium nitrate:

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

In each of the following cases, carry out the calculations to determine the quantities required.

a If 1.0 mol of potassium iodide reacts with 1.0 mol of lead(II) nitrate, determine which reactant is in excess and by how many mol.
b If 0.50 mol of potassium iodide reacts with 2.0 mol of lead(II) nitrate, determine which reactant is in excess and by how many mol.
c If 1.00 g of lead(II) nitrate reacts with 1.50 g of potassium iodide, determine which reactant is in excess and the mass of lead(II) iodide that forms.

Chapter review

07

KEY TERMS

amount of substance
Avogadro's constant
empirical formula
excess reactant
limiting reactant
molar mass
mole
mole ratio
molecular formula
percentage composition
relative formula mass
relative isotopic mass
stoichiometric calculation
stoichiometry

REVIEW QUESTIONS

1 For each of the following numbers of molecules, calculate the amount of substance, in mol.
 a 4.50×10^{23} molecules of water (H_2O)
 b 9.00×10^{24} molecules of methane (CH_4)
 c 2.3×10^{28} molecules of chlorine (Cl_2)
 d 1 molecule of sucrose ($C_{12}H_{22}O_{11}$)

2 For each of the following amounts of molecular substances, calculate the:
 i number of molecules
 ii total number of atoms
 iii the number of significant figures to which the number of mol can be given.
 a 1.45 mol of ammonia (NH_3)
 b 0.576 mol of hydrogen sulfide (H_2S)
 c 0.0153 mol of hydrogen nitrate (HNO_3)
 d 2.5 mol of sucrose ($C_{12}H_{22}O_{11}$)

3 How would the molar mass (M) of a compound differ from its relative molecular mass (M_r)?

4 What is the molar mass (M) of each of the following?
 a iron (Fe)
 b sulfuric acid (H_2SO_4)
 c sodium oxide (Na_2O)
 d zinc nitrate ($Zn(NO_3)_2$)
 e glycine (H_2NCH_2COOH)
 f aluminium sulfate ($Al_2(SO_4)_3$)
 g hydrated iron(III) chloride ($FeCl_3{\bullet}6H_2O$)

5 a If 6.022×10^{23} atoms of calcium have a mass of 40.08 g, what is the mass of one calcium atom?
 b If 1 mol of water molecules has a mass of 18.02 g, what is the mass of one water molecule?
 c What is the mass of one molecule of carbon dioxide?

6 For each of the following molecular substances, calculate the:
 i amount of substance in moles
 ii number of molecules
 iii total number of atoms.
 a 4.2 g of phosphorus (P_4)
 b 75.0 g of sulfur (S_8)
 c 0.32 g of hydrogen chloride (HCl)
 d 2.2×10^{-2} g of glucose ($C_6H_{12}O_6$).

7 For each of the following ionic substances, calculate the amount of:
 i substance, in moles
 ii each ion, in moles.
 a 5.85 g of NaCl
 b 45.0 g of $CaCl_2$
 c 1.68 g of $Fe_2(SO_4)_3$

8 Calculate the molar mass of a substance if:
 a 2.0 mol of the substance has a mass of 80 g
 b 0.10 mol of the substance has a mass of 9.8 g
 c 1.7 mol of the substance has a mass of 74.8 g
 d 3.50 mol of the substance has a mass of 371 g.

9 Which of the following metal samples has the greatest mass?
 A 100 g copper
 B 4.0 mol of iron atoms
 C 1.2×10^{24} atoms of silver

10 A new antibiotic has been isolated and only 2.0 mg is available. The molar mass is found to be 12.5 kg mol^{-1}.
 a Express the molar mass in g mol^{-1}.
 b Calculate the amount of antibiotic, in mol.
 c How many molecules of antibiotic have been isolated?

11 Calculate the percentage by mass of each element in:
 a Al_2O_3
 b $Cu(OH)_2$
 c $MgCl_2{\bullet}6H_2O$
 d $Fe_2(SO_4)_3$
 e perchloric acid ($HClO_4$).

12 A compound used as a solvent for dyes has the following composition by mass: 32.0% carbon, 6.7% hydrogen, 18.7% nitrogen and 42.6% oxygen. Find the empirical formula of the compound.

13 A clear liquid extracted from fermented lemons was found to consist of carbon, hydrogen and oxygen. Analysis showed it to be 52.2% carbon and 34.8% oxygen by mass.

a Find the empirical formula of the substance.

b If 2.17 mol of the compound has a mass of 100 g, find the molecular formula of the compound.

14 A hydrocarbon is a compound that contains carbon and hydrogen only. Determine the empirical formula of a hydrocarbon that is used as a specialty fuel and contains 90.0% carbon.

15 Find the relative atomic mass of nickel if 3.370 g nickel was obtained by reduction of 4.286 g of the oxide (NiO). You may need to refer to the periodic table at the end of this book.

16 4.150 g tungsten was burned in chlorine and 8.950 g tungsten chloride (WCl_6) was formed. Find the relative atomic mass of tungsten. You may need to refer to the periodic table at the end of this book.

17 Determine the molecular formulae of compounds with the following compositions and relative molecular masses:

a 82.75% carbon, 17.25% hydrogen; $M_r = 58.12$

b 43.66% phosphorus, 56.34% oxygen; $M_r = 283.88$

c 40.0% carbon, 6.7% hydrogen, 53.3% oxygen; $M_r = 180.16$

d 0.164 g hydrogen, 5.25 g sulfur, 9.18 g oxygen; $M_r = 178.16$

18 Using suitable examples, clearly distinguish between:

a relative isotopic mass

b relative atomic mass

c relative molecular mass

d relative formula mass

e molar mass.

19 Caffeine contains 49.48% carbon, 5.15% hydrogen, 28.87% nitrogen; the rest is oxygen.

a Determine the empirical formula of caffeine.

b If 0.200 mol of caffeine has a mass of 38.8 g, what is the molar mass of a caffeine molecule?

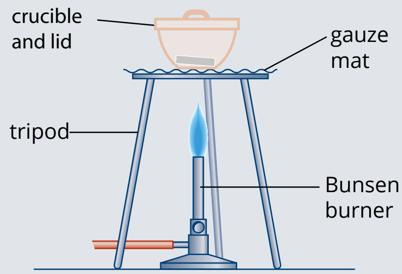
c Determine the molecular formula of caffeine.

d How many moles of caffeine molecules are in 1.00 g caffeine?

e How many molecules of caffeine are in 1.00 g caffeine?

f How many atoms all together are there in 1.00 g caffeine?

20 The empirical formula of a metal oxide can be found by experimentation, as shown in the figure below.



The mass of the oxygen that reacts with the mass of the metal must be determined. Steps A–F form the experimental method.

A Ignite a burner and heat the metal.

B Allow the crucible to cool, then weigh it.

C Continue the reaction until no further change occurs.

D Clean a piece of metal with emery paper to remove any oxide layer.

E Place the metal in a clean, weighed crucible and cover with a lid.

F Weigh the metal and record its mass.

a Place the steps in the correct order by letter.

b Wan and Eric collected the following data:

mass of the metal = 0.542 g

mass of the empty crucible = 20.310 g

mass of the crucible and metal oxide = 21.068 g

They found from this data that the metal oxide has a 1:1 formula, i.e. MO, where M = metal. Copy and complete the figure below with the data provided.

	Metal	Oxygen
mass (g)		
relative atomic mass		16.00
moles		
ratio		

c What metal was used in the experiment?

21 For each amount given, calculate the amounts, in mol, of the other reactants and products required for a complete reaction according to the following equation:

$3Ca(NO_3)_2(aq) + 2Na_3PO_4(aq) \rightarrow Ca_3(PO_4)_2(s) + 6NaNO_3(aq)$

$Ca(NO_3)_2$	Na_3PO_4	$Ca_3(PO_4)_2$	$NaNO_3$
27 mol			
	0.48 mol		
		0.18 mol	
			2.4 mol

22 When solutions of iron(II) sulfate and potassium hydroxide are mixed, the reaction that occurs can be represented as:

$FeSO_4(aq) + 2KOH(aq) \rightarrow Fe(OH)_2(s) + K_2SO_4(aq)$

Complete the following expressions based on this equation. The first one has been done for you.

$n(KOH) = \frac{2}{1} \times n(Fe(OH)_2)$

$n(FeSO_4) = \frac{(\ )}{(\ )} \times n(KOH)$

$n(KOH) = \frac{(\ )}{(\ )} \times n(K_2SO_4)$

$n(Fe(OH)_2) = \frac{(\ )}{(\ )} \times n(FeSO_4)$

23 A barbecue burns 50.2 g of propane (C_3H_8) to cook a meal with this reaction:

$$C_3H_8(l) + 10O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$$

What mass of CO_2 is produced?

24 To make a mobile phone battery, molten lithium chloride is used to make lithium metal and chlorine gas in a reaction that is represented as:

$$2LiCl(l) \rightarrow 2Li(l) + Cl_2(g)$$

What mass of LiCl is needed to produce 8.22 g of Li for a battery?

25 In three separate experiments, different amounts of carbon and oxygen were reacted together to form carbon dioxide according to the equation:

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

The table lists the amounts of reactants and products in each of the three experiments. Complete the table to indicate the amounts of the products formed and of the remaining reactants at the end of the reaction.

Carbon atoms available	Oxygen molecules available	Carbon dioxide molecules produced	Carbon atoms in excess	Oxygen molecules in excess
8	20			
1000	3000			
9 mol	6 mol			

26 Sodium can react with oxygen gas to form sodium oxide. The equation for this reaction is:

$$4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$$

3.0 mol of sodium is reacted with 0.8 mol of oxygen gas.

a Determine which reactant is in excess.

b How many moles of sodium oxide is produced in the reaction?

27 16.0 g of hydrogen sulfide is mixed with 20.0 g of sulfur dioxide, and they react according to this equation:

$$2H_2S(g) + SO_2(g) \rightarrow 2H_2O(l) + 3S(s)$$

a Calculate the mass of sulfur produced.

b Calculate the mass of reactant left after the reaction.

28 4.40 g of P_4O_6 and 3.00 g of I_2 are mixed and allowed to react according to this equation:

$$5P_4O_6(s) + 8I_2(g) \rightarrow 4P_2I_4(s) + 3P_4O_{10}(s)$$

a Which reactant is in excess and by how much, in g?

b What mass of P_2I_4 forms?

c What mass of P_4O_{10} forms?

d What is the total mass of all the products? (Hint: Compare this with the mass of the reactants.)

29 Reflect on the Inquiry activity on page 222. Now that you understand the mole, explain any patterns you saw in your results.