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Chapter 1: Moles and equations

Learning outcomes

You should be able to:

- define and use the terms:
  - **relative atomic mass**, **isotopic mass** and **formula mass** based on the ^{12}C scale
  - **empirical formula** and **molecular formula**
  - the **mole** in terms of the Avogadro constant
- analyse and use mass spectra to calculate the relative atomic mass of an element
- calculate empirical and molecular formulae using combustion data or composition by mass
- write and construct balanced equations
- perform calculations, including use of the mole concept involving:
  - reacting masses (from formulae and equations)
  - volumes of gases (e.g. in the burning of hydrocarbons)
  - volumes and concentrations of solutions
- deduce stoichiometric relationships from calculations involving reacting masses, volumes of gases and volumes and concentrations of solutions.

Introduction

For thousands of years, people have heated rocks and distilled plant juices to extract materials. Over the past two centuries, chemists have learnt more and more about how to get materials from rocks, from the air and the sea, and from plants. They have also found out the right conditions to allow these materials to react together to make new substances, such as dyes, plastics and medicines. When we make a new substance it is important to mix the reactants in the correct proportions to ensure that none is wasted. In order to do this we need to know about the relative masses of atoms and molecules and how these are used in chemical calculations.

Figure 1.1 A titration is a method used to find the amount of a particular substance in a solution.

Masses of atoms and molecules

Relative atomic mass, A_r

Atoms of different elements have different masses. When we perform chemical calculations, we need to know how heavy one atom is compared with another. The mass of a single atom is so small that it is impossible to weigh it directly. To overcome this problem, we have to weigh a lot of atoms. We then compare this mass with the mass of the same number of 'standard' atoms. Scientists have chosen to use the isotope carbon-12 as the standard. This has been given a mass of exactly 12 units. The mass of other atoms is found by comparing their mass with the mass of carbon-12 atoms. This is called the **relative atomic mass**, A_r.

> The relative atomic mass is the weighted average mass of naturally occurring atoms of an element on a scale where an atom of carbon-12 has a mass of exactly 12 units.

From this it follows that:

$$A_r\,[\text{element } Y] = \frac{\text{average mass of one atom of element } Y \times 12}{\text{mass of one atom of carbon-12}}$$

We use the average mass of the atom of a particular element because most elements are mixtures of isotopes. For example, the exact A_r of hydrogen is 1.0079. This is very close to 1 and most periodic tables give the A_r of hydrogen as 1.0. However, some elements in the Periodic Table have values that are not whole numbers. For example, the A_r for chlorine is 35.5. This is because chlorine has two isotopes. In a sample of chlorine, chlorine-35 makes up about three-quarters of the chlorine atoms and chlorine-37 makes up about a quarter.

Relative isotopic mass

Isotopes are atoms that have the same number of protons but different numbers of neutrons (see page 28). We represent the **nucleon number** (the total number of neutrons plus protons in an atom) by a number written at the top left-hand corner of the atom's symbol, e.g. ^{20}Ne, or by a number written after the atom's name or symbol, e.g. neon-20 or Ne-20.

We use the term **relative isotopic mass** for the mass of a particular isotope of an element on a scale where an atom of carbon-12 has a mass of exactly 12 units. For example, the relative isotopic mass of carbon-13 is 13.00. If we know both the natural abundance of every isotope of an element and their isotopic masses, we can calculate

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the relative atomic mass of the element very accurately. To find the necessary data we use an instrument called a mass spectrometer (see box on mass spectrometry).

Relative molecular mass, M_r

The **relative molecular mass** of a **compound** (M_r) is the relative mass of one molecule of the compound on a scale where the carbon-12 isotope has a mass of exactly 12 units. We find the relative molecular mass by adding up the relative atomic masses of all the atoms present in the molecule.

For example, for methane:

formula	CH_4
atoms present	$1 \times C; 4 \times H$
add A_r values	$(1 \times A_r[C]) + (4 \times A_r[H])$
M_r of methane	$= (1 \times 12.0) + (4 \times 1.0)$
	$= 16.0$

Relative formula mass

For compounds containing ions we use the term **relative formula mass**. This is calculated in the same way as for relative molecular mass. It is also given the same symbol, M_r. For example, for magnesium hydroxide:

formula	$Mg(OH)_2$
ions present	$1 \times Mg^{2+}; 2 \times (OH^-)$
add A_r values	$(1 \times A_r[Mg]) + (2 \times (A_r[O] + A_r[H]))$
M_r of magnesium hydroxide	$= (1 \times 24.3) + (2 \times (16.0 + 1.0))$
	$= 58.3$

QUESTION

1 Use the Periodic Table on page 473 to calculate the relative formula masses of the following:

a calcium chloride, $CaCl_2$

b copper(II) sulfate, $CuSO_4$

c ammonium sulfate, $(NH_4)_2SO_4$

d magnesium nitrate-6-water, $Mg(NO_3)_2.6H_2O$

Hint: for part **d** you need to calculate the mass of water separately and then add it to the M_r of $Mg(NO_3)_2$.

Accurate relative atomic masses

MASS SPECTROMETRY

A **mass spectrometer** (Figure 1.2) can be used to measure the mass of each isotope present in an element. It also compares how much of each isotope is present – the relative abundance (isotopic abundance). A simplified diagram of a mass spectrometer is shown in Figure 1.3. You will not be expected to know the details of how a mass spectrometer works, but it is useful to understand how the results are obtained.

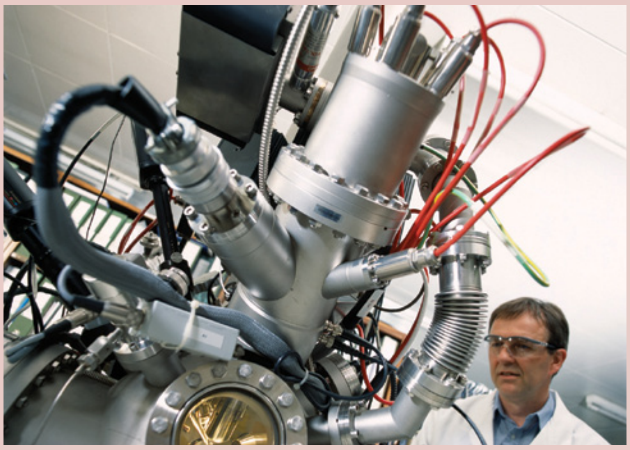

Figure 1.2 A mass spectrometer is a large and complex instrument.

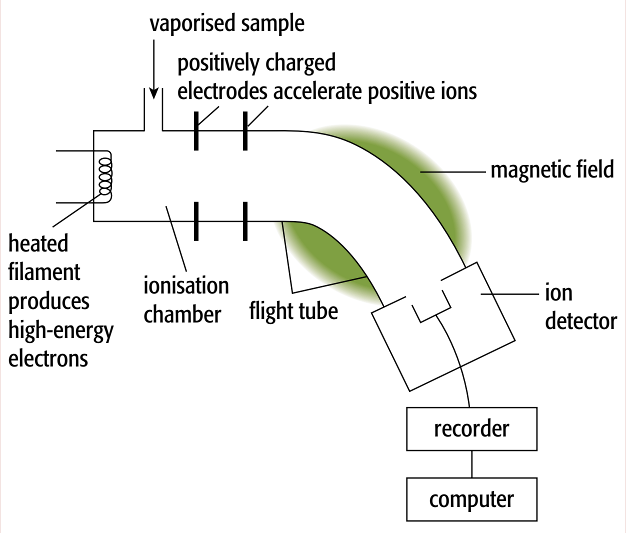

Figure 1.3 Simplified diagram of a mass spectrometer.

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MASS SPECTROMETRY (CONTINUED)

The atoms of the element in the vaporised sample are converted into ions. The stream of ions is brought to a detector after being deflected (bent) by a strong magnetic field. As the magnetic field is increased, the ions of heavier and heavier isotopes are brought to the detector. The detector is connected to a computer, which displays the mass spectrum.

The mass spectrum produced shows the relative abundance (isotopic abundance) on the vertical axis and the mass to ion charge ratio (*m*/*e*) on the horizontal axis. Figure 1.4 shows a typical mass spectrum for a sample of lead. Table 1.1 shows how the data is interpreted.

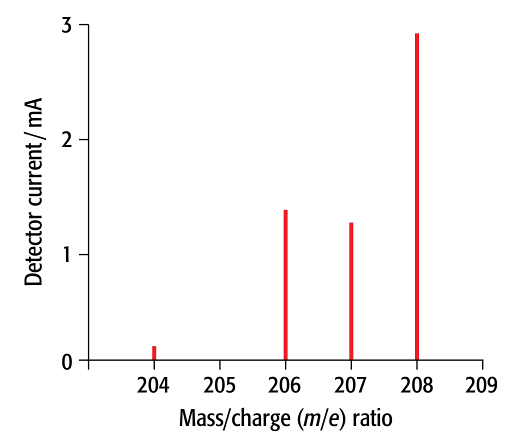

Figure 1.4 The mass spectrum of a sample of lead.

For singly positively charged ions the *m*/*e* values give the nucleon number of the isotopes detected. In the case of lead, Table 1.1 shows that 52% of the lead is the isotope with an isotopic mass of 208. The rest is lead-204 (2%), lead-206 (24%) and lead-207 (22%).

Isotopic mass	Relative abundance / %
204	2
206	24
207	22
208	52
total	100

Table 1.1 The data from Figure 1.4.

Determination of A_r from mass spectra

We can use the data obtained from a mass spectrometer to calculate the relative atomic mass of an element very accurately. To calculate the relative atomic mass we follow this method:

- multiply each isotopic mass by its percentage abundance
- add the figures together
- divide by 100.

We can use this method to calculate the relative atomic mass of neon from its mass spectrum, shown in Figure 1.5.

The mass spectrum of neon has three peaks:

^{20}Ne (90.9%), ^{21}Ne (0.3%) and ^{22}Ne (8.8%).

A_r of neon

$$= \frac{(20 \times 90.9) + (21.0 \times 0.3) + (22 \times 8.8)}{100} = 20.2$$

Note that this answer is given to 3 significant figures, which is consistent with the data given.

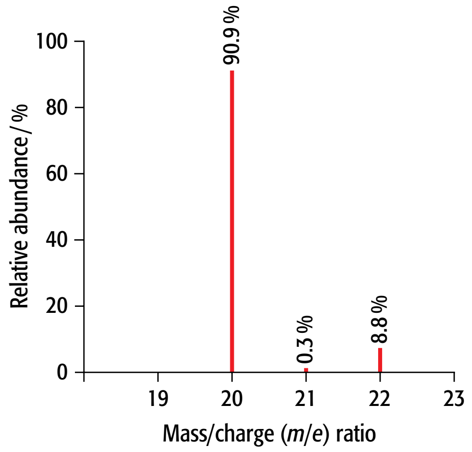

Figure 1.5 The mass spectrum of neon, Ne.

A high-resolution mass spectrometer can give very accurate relative isotopic masses. For example ^{16}O = 15.995 and ^{32}S = 31.972. Because of this, chemists can distinguish between molecules such as SO_2 and S_2, which appear to have the same relative molecular mass.

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QUESTION

2 Look at the mass spectrum of germanium, Ge.

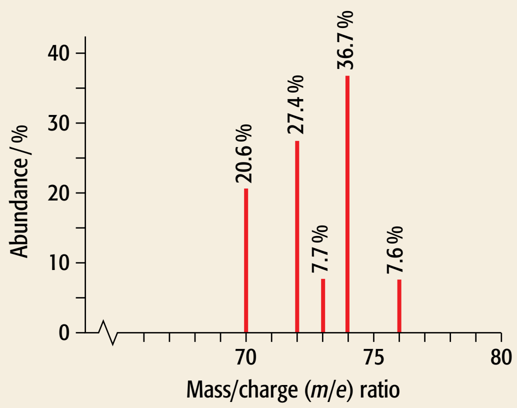

Figure 1.6 The mass spectrum of germanium.

a Write the isotopic formula for the heaviest isotope of germanium.

b Use the % abundance of each isotope to calculate the relative atomic mass of germanium.

Amount of substance

The mole and the Avogadro constant

The formula of a compound shows us the number of atoms of each element present in one formula unit or one molecule of the compound. In water we know that two atoms of hydrogen (A_r = 1.0) combine with one atom of oxygen (A_r = 16.0). So the ratio of mass of hydrogen atoms to oxygen atoms in a water molecule is 2 : 16. No matter how many molecules of water we have, this ratio will always be the same. But the mass of even 1000 atoms is far too small to be weighed. We have to scale up much more than this to get an amount of substance that is easy to weigh.

The relative atomic mass or relative molecular mass of a substance in grams is called a **mole** of the substance. So a mole of sodium (A_r = 23.0) weighs 23.0 g. The abbreviation for a mole is mol. We define the mole in terms of the standard carbon-12 isotope (see page 28).

One mole of a substance is the amount of that substance that has the same number of specific particles (atoms, molecules or ions) as there are atoms in exactly 12 g of the carbon-12 isotope.

We often refer to the mass of a mole of substance as its **molar mass** (abbreviation *M*). The units of molar mass are g mol^{-1}.

The number of atoms in a mole of atoms is very large: 6.02×10^{23} atoms. This number is called the **Avogadro constant** (or Avogadro number). The symbol for the Avogadro constant is *L* (the symbol N_A may also be used). The Avogadro constant applies to atoms, molecules, ions and electrons. So in 1 mole of sodium there are 6.02×10^{23} sodium atoms and in 1 mole of sodium chloride (NaCl) there are 6.02×10^{23} sodium ions and 6.02×10^{23} chloride ions.

It is important to make clear what type of particles we are referring to. If we just state 'moles of chlorine', it is not clear whether we are thinking about chlorine atoms or chlorine molecules. A mole of chlorine molecules, Cl_2, contains 6.02×10^{23} chlorine molecules but twice as many chlorine atoms, as there are two chlorine atoms in every chlorine molecule.

Figure 1.7 Amedeo Avogadro (1776–1856) was an Italian scientist who first deduced that equal volumes of gases contain equal numbers of molecules. Although the Avogadro constant is named after him, it was left to other scientists to calculate the number of particles in a mole.

Moles and mass

The Système International (SI) base unit for mass is the kilogram. But this is a rather large mass to use for general laboratory work in chemistry. So chemists prefer to use the relative molecular mass or formula mass in grams (1000 g = 1 kg). You can find the number of moles of a substance by using the mass of substance and the relative atomic mass (A_r) or relative molecular mass (M_r).

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

WORKED EXAMPLE

1 How many moles of sodium chloride are present in 117.0 g of sodium chloride, NaCl?

(A_r values: Na = 23.0, Cl = 35.5)

molar mass of NaCl = 23.0 + 35.5

= 58.5 g mol^{-1}

number of moles $= \frac{\text{mass}}{\text{molar mass}}$

$= \frac{117.0}{58.5}$

= 2.0 mol

Figure 1.8 From left to right, one mole of each of copper, bromine, carbon, mercury and lead.

QUESTION

3 **a** Use these A_r values (Fe = 55.8, N = 14.0, O = 16.0, S = 32.1) to calculate the amount of substance in moles in each of the following:

i 10.7 g of sulfur atoms

ii 64.2 g of sulfur molecules (S_8)

iii 60.45 g of anhydrous iron(III) nitrate, $Fe(NO_3)_3$.

b Use the value of the Avogadro constant (6.02×10^{23} mol^{-1}) to calculate the total number of atoms in 7.10 g of chlorine atoms. (A_r value: Cl = 35.5)

To find the mass of a substance present in a given number of moles, you need to rearrange the equation

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

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

WORKED EXAMPLE

2 What mass of sodium hydroxide, NaOH, is present in 0.25 mol of sodium hydroxide?

(A_r values: H = 1.0, Na = 23.0, O = 16.0)

molar mass of NaOH = 23.0 + 16.0 + 1.0

= 40.0 g mol^{-1}

mass = number of moles × molar mass

= 0.25 × 40.0 g

= 10.0 g NaOH

QUESTION

4 Use these A_r values: C = 12.0, Fe = 55.8, H = 1.0, O = 16.0, Na = 23.0.

Calculate the mass of the following:

a 0.20 moles of carbon dioxide, CO_2

b 0.050 moles of sodium carbonate, Na_2CO_3

c 5.00 moles of iron(II) hydroxide, $Fe(OH)_2$

Mole calculations

Reacting masses

When reacting chemicals together we may need to know what mass of each reactant to use so that they react exactly and there is no waste. To calculate this we need to know the chemical equation. This shows us the ratio of moles of the reactants and products – the **stoichiometry** of the equation. The balanced equation shows this stoichiometry. For example, in the reaction

$$Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2$$

1 mole of iron(III) oxide reacts with 3 moles of carbon monoxide to form 2 moles of iron and 3 moles of carbon dioxide. The stoichiometry of the equation is 1 : 3 : 2 : 3. The large numbers that are included in the equation (3, 2 and 3) are called stoichiometric numbers.

In order to find the mass of products formed in a chemical reaction we use:

- the mass of the reactants
- the molar mass of the reactants
- the balanced equation.

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Figure 1.9 Iron reacting with sulfur to produce iron sulfide. We can calculate exactly how much iron is needed to react with sulfur and the mass of the products formed by knowing the molar mass of each reactant and the balanced chemical equation.

WORKED EXAMPLE

3 Magnesium burns in oxygen to form magnesium oxide.

$2Mg + O_2 \longrightarrow 2MgO$

We can calculate the mass of oxygen needed to react with 1 mole of magnesium. We can calculate the mass of magnesium oxide formed.

Step 1 Write the balanced equation.

Step 2 Multiply each formula mass in g by the relevant stoichiometric number in the equation.

$2Mg$	+	O_2	$\longrightarrow$	$2MgO$
2 × 24.3 g		1 × 32.0 g		2 × (24.3 g + 16.0 g)
48.6 g		32.0 g		80.6 g

From this calculation we can deduce that:

- 32.0 g of oxygen are needed to react exactly with 48.6 g of magnesium
- 80.6 g of magnesium oxide are formed.

If we burn 12.15 g of magnesium (0.5 mol) we get 20.15 g of magnesium oxide. This is because the stoichiometry of the reaction shows us that for every mole of magnesium burnt we get the same number of moles of magnesium oxide.

In this type of calculation we do not always need to know the molar mass of each of the reactants. If one or more of the reactants is in excess, we need only know the mass in grams and the molar mass of the reactant that is not in excess (the limiting reactant).

WORKED EXAMPLE

4 Iron(III) oxide reacts with carbon monoxide to form iron and carbon dioxide.

$Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2$

Calculate the maximum mass of iron produced when 798 g of iron(III) oxide is reduced by excess carbon monoxide.

(A_r values: Fe = 55.8, O = 16.0)

Step 1 $Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2$

Step 2 1 mole iron(III) oxide $\longrightarrow$ 2 moles iron

(2 × 55.8) + (3 × 16.0) $\longrightarrow$ 2 × 55.8

159.6 g Fe_2O_3 $\longrightarrow$ 111.6 g Fe

Step 3 798 g $\quad \dfrac{111.6}{159.6} \times 798$

= 558 g Fe

You can see that in step **3**, we have simply used ratios to calculate the amount of iron produced from 798 g of iron(III) oxide.

QUESTION

5 a Sodium reacts with excess oxygen to form sodium peroxide, Na_2O_2.

$2Na + O_2 \longrightarrow Na_2O_2$

Calculate the maximum mass of sodium peroxide formed when 4.60 g of sodium is burnt in excess oxygen.

(A_r values: Na = 23.0, O = 16.0)

b Tin(IV) oxide is reduced to tin by carbon. Carbon monoxide is also formed.

$SnO_2 + 2C \longrightarrow Sn + 2CO$

Calculate the mass of carbon that exactly reacts with 14.0 g of tin(IV) oxide. Give your answer to 3 significant figures.

(A_r values: C = 12.0, O = 16.0, Sn = 118.7)

The stoichiometry of a reaction

We can find the stoichiometry of a reaction if we know the amounts of each reactant that exactly react together and the amounts of each product formed.

For example, if we react 4.0 g of hydrogen with 32.0 g of oxygen we get 36.0 g of water. (A_r values: H = 1.0, O = 16.0)

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$$\begin{array}{ccccc} \text{hydrogen } (H_2) & + & \text{oxygen } (O_2) & \longrightarrow & \text{water } (H_2O) \\ \dfrac{4.0}{2 \times 1.0} & & \dfrac{32.0}{2 \times 16.0} & & \dfrac{36.0}{(2 \times 1.0) + 16.0} \\ = 2\text{ mol} & & = 1\text{ mol} & & = 2\text{ mol} \end{array}$$

This ratio is the ratio of stoichiometric numbers in the equation. So the equation is:

$$2H_2 + O_2 \longrightarrow 2H_2O$$

We can still deduce the stoichiometry of this reaction even if we do not know the mass of oxygen that reacted. The ratio of hydrogen to water is 1 : 1. But there is only one atom of oxygen in a molecule of water – half the amount in an oxygen molecule. So the mole ratio of oxygen to water in the equation must be 1 : 2.

QUESTION

6 56.2 g of silicon, Si, reacts exactly with 284.0 g of chlorine, Cl_2, to form 340.2 g of silicon(IV) chloride, $SiCl_4$. Use this information to calculate the stoichiometry of the reaction.

(A_r values: Cl = 35.5, Si = 28.1)

Significant figures

When we perform chemical calculations it is important that we give the answer to the number of significant figures that fits with the data provided. The examples show the number 526.84 rounded up to varying numbers of significant figures.

rounded to 4 significant figures = 526.8
rounded to 3 significant figures = 527
rounded to 2 significant figures = 530

When you are writing an answer to a calculation, the answer should be to the same number of significant figures as the least number of significant figures in the data.

WORKED EXAMPLE

5 How many moles of calcium oxide are there in 2.9 g of calcium oxide?

(A_r values: Ca = 40.1, O = 16.0)

If you divide 2.9 by 56.1, your calculator shows 0.051 693 .... The least number of significant figures in the data, however, is 2 (the mass is 2.9 g). So your answer should be expressed to 2 significant figures, as 0.052 mol.

WORKED EXAMPLE (CONTINUED)

Note 1 Zeros before a number are not significant figures. For example, 0.004 is only to 1 significant figure.

Note 2 After the decimal point, zeros after a number are significant figures. 0.0040 has 2 significant figures and 0.004 00 has 3 significant figures.

Note 3 If you are performing a calculation with several steps, do not round up in between steps. Round up at the end.

Percentage composition by mass

We can use the formula of a compound and relative atomic masses to calculate the percentage by mass of a particular element in a compound.

% by mass

$$= \frac{\text{atomic mass} \times \text{number of moles of particular element in a compound}}{\text{molar mass of compound}} \times 100$$

WORKED EXAMPLE

6 Calculate the percentage by mass of iron in iron(III) oxide, Fe_2O_3.

(A_r values: Fe = 55.8, O = 16.0)

$$\text{\% mass of iron} = \frac{2 \times 55.8}{(2 \times 55.8) + (3 \times 16.0)} \times 100$$
$$= 69.9\,\%$$

Figure 1.10 This iron ore is impure Fe_2O_3. We can calculate the mass of iron that can be obtained from Fe_2O_3 by using molar masses.

QUESTION

7 Calculate the percentage by mass of carbon in ethanol, C_2H_5OH.
(A_r values: C = 12.0, H = 1.0, O = 16.0)

Empirical formulae

The **empirical formula** of a compound is the simplest whole number ratio of the elements present in one molecule or formula unit of the compound. The **molecular formula** of a compound shows the total number of atoms of each element present in a molecule.

Table 1.2 shows the empirical and molecular formulae for a number of compounds.

- The formula for an ionic compound is always its empirical formula.
- The empirical formula and molecular formula for simple inorganic molecules are often the same.
- Organic molecules often have different empirical and molecular formulae.

Compound	Empirical formula	Molecular formula
water	H_2O	H_2O
hydrogen peroxide	HO	H_2O_2
sulfur dioxide	SO_2	SO_2
butane	C_2H_5	C_4H_{10}
cyclohexane	CH_2	C_6H_{12}

Table 1.2 Some empirical and molecular formulae.

QUESTION

8 Write the empirical formula for:
a hydrazine, N_2H_4
b octane, C_8H_{18}
c benzene, C_6H_6
d ammonia, NH_3

The empirical formula can be found by determining the mass of each element present in a sample of the compound. For some compounds this can be done by combustion. An organic compound must be very pure in order to calculate its empirical formula. Chemists often use gas chromatography to purify compounds before carrying out formula analysis.

WORKED EXAMPLES

7 Deduce the formula of magnesium oxide.

This can be found as follows:

- burn a known mass of magnesium (0.486 g) in excess oxygen
- record the mass of magnesium oxide formed (0.806 g)
- calculate the mass of oxygen that has combined with the magnesium (0.806 – 0.486 g) = 0.320 g
- calculate the mole ratio of magnesium to oxygen
- (A_r values: Mg = 24.3, O = 16.0)

$$\text{moles of Mg} = \frac{0.486\,\text{g}}{24.3\,\text{g mol}^{-1}} = 0.0200\,\text{mol}$$

$$\text{moles of oxygen} = \frac{0.320\,\text{g}}{16.0\,\text{g mol}^{-1}} = 0.0200\,\text{mol}$$

The simplest ratio of magnesium : oxygen is 1 : 1. So the empirical formula of magnesium oxide is MgO.

8 When 1.55 g of phosphorus is completely combusted 3.55 g of an oxide of phosphorus is produced. Deduce the empirical formula of this oxide of phosphorus.

(A_r values: O = 16.0, P = 31.0)

		P	**O**
Step 1	note the mass of each element	1.55 g	3.55 – 1.55 = 2.00 g
Step 2	divide by atomic masses	$\frac{1.55\,\text{g}}{31.0\,\text{g mol}^{-1}}$ = 0.05 mol	$\frac{2.00\,\text{g}}{16.0\,\text{g mol}^{-11}}$ = 0.125 mol
Step 3	divide by the lowest figure	$\frac{0.05}{0.05} = 1$	$\frac{0.125}{0.05} = 2.5$
Step 4	if needed, obtain the lowest whole number ratio to get empirical formula	P_2O_5	

An empirical formula can also be deduced from data that give the percentage composition by mass of the elements in a compound.

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WORKED EXAMPLE

9 A compound of carbon and hydrogen contains 85.7% carbon and 14.3% hydrogen by mass. Deduce the empirical formula of this hydrocarbon.

(A_r values: C = 12.0, O = 16.0)

	C	H
Step 1 note the % by mass	85.7	14.3
Step 2 divide by A_r values	$\frac{85.7}{12.0} = 7.142$	$\frac{14.3}{1.0} = 14.3$
Step 3 divide by the lowest figure	$\frac{7.142}{7.142} = 1$	$\frac{14.3}{7.142} = 2$

Empirical formula is CH_2.

QUESTION

9 The composition by mass of a hydrocarbon is 10% hydrogen and 90% carbon. Deduce the empirical formula of this hydrocarbon.

(A_r values: C = 12.0, H = 1.0)

Molecular formulae

The molecular formula shows the actual number of each of the different atoms present in a molecule. The molecular formula is more useful than the empirical formula. We use the molecular formula to write balanced equations and to calculate molar masses. The molecular formula is always a multiple of the empirical formula. For example, the molecular formula of ethane, C_2H_6, is two times the empirical formula, CH_3.

In order to deduce the molecular formula we need to know:

- the relative formula mass of the compound
- the empirical formula.

WORKED EXAMPLE

10 A compound has the empirical formula CH_2Br. Its relative molecular mass is 187.8. Deduce the molecular formula of this compound.

(A_r values: Br = 79.9, C = 12.0, H = 1.0)

Step 1 find the empirical formula mass:
12.0 + (2 × 1.0) + 79.9 = 93.9

WORKED EXAMPLE (CONTINUED)

Step 2 divide the relative molecular mass by the empirical formula mass: $\frac{187.8}{93.9} = 2$

Step 3 multiply the number of atoms in the empirical formula by the number in step **2**:
2 × CH_2Br, so molecular formula is $C_2H_4Br_2$.

QUESTION

10 The empirical formulae and molar masses of three compounds, A, B and C, are shown in the table below. Calculate the molecular formula of each of these compounds.

(A_r values: C = 12.0, Cl = 35.5, H = 1.0)

Compound	Empirical formula	M_r
A	C_3H_5	82
B	CCl_3	237
C	CH_2	112

Chemical formulae and chemical equations

Deducing the formula

The electronic structure of the individual elements in a compound determines the formula of a compound (see page 33). The formula of an ionic compound is determined by the charges on each of the ions present. The number of positive charges is balanced by the number of negative charges so that the total charge on the compound is zero. We can work out the formula for a compound if we know the charges on the ions. Figure 1.11 shows the charges on some simple ions related to the position of the elements in the Periodic Table. The form of the Periodic Table that we shall be using has 18 groups because the transition elements are numbered as Groups 3 to 12. So, aluminium is in Group 13 and chlorine is in Group 17.

For simple metal ions in Groups 1 and 2, the value of the positive charge is the same as the group number. For a simple metal ion in Group 13, the value of the positive charge is 3+. For a simple non-metal ion in Groups 15 to 17, the value of the negative charge is 18 minus the group

