

YOU WILL EXAMINE:

- the use of mass spectrometry in identifying isotopes of elements
- how to calculate relative atomic masses
- the mole concept
- Avogadro's constant and its relationship to the mole concept
- how to calculate molar masses
- how to determine percentage composition and empirical formulas.

It must . . . be admitted that very simple relations . . . exist between the volumes of gaseous substances and the numbers of simple or compound molecules which form them.

Amedeo Avogadro

In an oil refinery, there are billions of reactions between individual atoms. How can these reactions be controlled without knowing how many atoms and molecules are reacting? The unique unit for measuring amounts of chemicals is called the mole. This enormous number allows billions of atoms to be counted in a practical way.

ELEMENTS		
	Hydrogen	1
	Azote	5
	Carbon	5
	Oxygen	7
	Phosphorus	9
	Sulphur	13
	Magnesia	20
	Lime	24
	Soda	28
	Potash	42
	Strontian	46
	Barites	68
	Iron	50
	Zinc	56
	Copper	56
	Lead	90
	Silver	190
	Gold	190
	Platina	190
	Mercury	167

John Dalton devised his own set of symbols for the known elements. Note the relative weights of the elements that he calculated based on hydrogen having a mass of 1.

Carbon-12 is used as the standard for comparing relative isotopic masses.

A mass spectrometer deflects the path of the lightest element the most. Similarly, it deflects the most highly charged particle the furthest.

A mass spectrometer. Different types of atoms of the same element can be separated on the basis of their mass:charge ratio.

Measuring atoms

Atoms are extremely tiny. Even though the air is full of oxygen and nitrogen molecules, you cannot see them. We know a lot about atoms and molecules, and this knowledge is invaluable when explaining the properties of substances. But, how do we measure atoms? The scale of atomic size means that chemists rarely deal with atoms one at a time. What is needed is a convenient way to compare and measure masses and amounts of particles in elements and compounds. The mole concept, which is central to most chemical calculations, helps us to work with the vast numbers of atoms that are present in different types of substances.

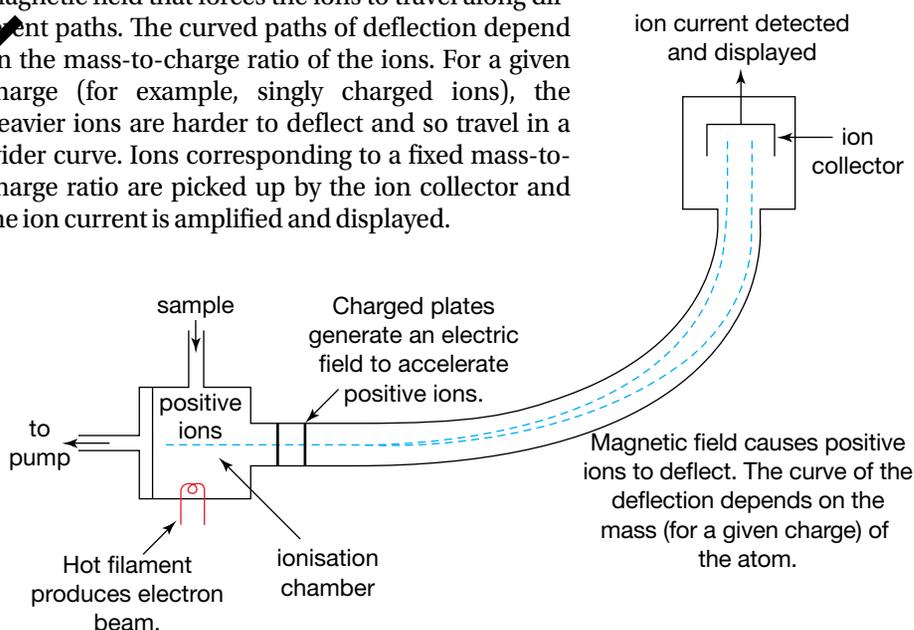
Atoms are extremely small but they still have mass. The problem is to find a way of measuring that mass. In 1803, English chemist John Dalton described matter as made up of particles that were solid, indivisible and having weight. He prepared a table of atomic weights by experimentally determining relative weights of elements in compounds; he based his figures on hydrogen having a mass of 1. We now use the term 'mass' instead of 'weight'.

This reference standard was later changed from hydrogen to oxygen, but this led to a disagreement between chemists, who used natural oxygen as the standard, and physicists, who used the oxygen-16 isotope only. Having two slightly different lists of atomic masses caused many problems. An agreement was reached in 1961 to change the standard to carbon-12, and this settled the dispute; this change had the added advantage that carbon-12 can be measured very accurately because it is stable and abundant. The development of an instrument called the mass spectrometer allowed scientists to accurately compare the masses of all atoms.

Measuring masses of atoms

In 1913, Joseph John Thomson, an English physicist, discovered that some elements can have isotopes. Francis Aston, another English scientist, developed Thomson's equipment into a **mass spectrometer**. Francis Aston was later awarded the Nobel Prize for identifying the 212 naturally occurring isotopes by comparing the relative masses of atoms.

The figure below shows the main features of a mass spectrometer. A sample of the element to be analysed is injected as a gas into the ionisation chamber, where the atoms are ionised by bombardment with electrons produced by the hot filament. The positive ions formed are accelerated through an electric field and deflected in a magnetic field that forces the ions to travel along different paths. The curved paths of deflection depend on the mass-to-charge ratio of the ions. For a given charge (for example, singly charged ions), the heavier ions are harder to deflect and so travel in a wider curve. Ions corresponding to a fixed mass-to-charge ratio are picked up by the ion collector and the ion current is amplified and displayed.



A mass spectrum shows information about the relative isotopic mass and percentage abundance of each isotope.

study on

Unit 1

AOS 1

Topic 3

Concept 1

Relative isotopic mass and relative atomic mass
Summary screen and practice questions

study on

Unit 1

AOS 1

Topic 3

Concept 2

Using mass spectrometry to find relative atomic mass
Summary screen and practice questions

The relative abundance of an isotope measures how much of that isotope is present compared with the other isotopes in an element.

A = mass number
 Z = atomic number

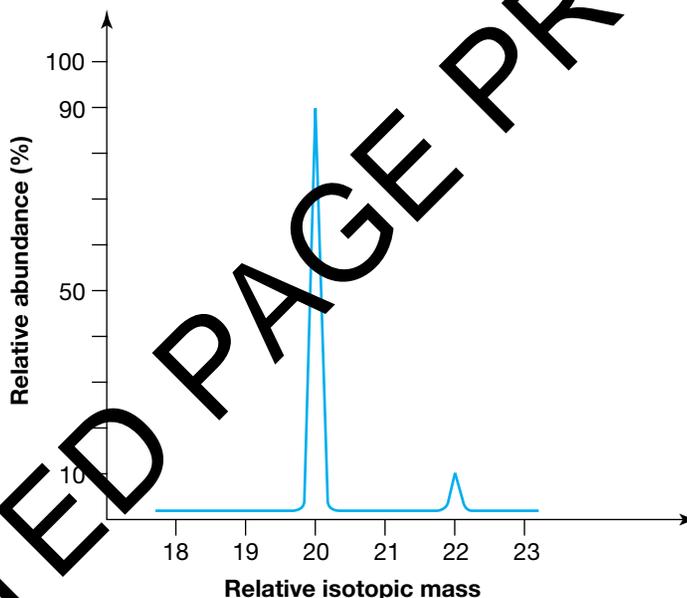
$$A_r = \frac{\sum \text{RIM} \times \text{abundance}}{100}$$

The mass spectrometer provides us with information about:

- the number of isotopes in a given sample of an element
- the relative isotopic mass of each isotope
- the percentage abundance of the isotopes.

The **relative isotopic mass** is the mass of a single isotope and is determined by comparing the mass of ions of the isotope to the value of a standard, $^{12}_6\text{C}$ (which has been assigned a mass of 12 exactly).

The figure below shows the mass spectrum for neon. The two peaks in the trace represent the two isotopes of neon. Both isotopes of neon have an atomic number of 10 (that is, the number of protons is 10), but one isotope has a relative isotopic mass of 20 and the other has a relative isotopic mass of 22. The atoms of neon-20 have 10 protons and 10 neutrons, and atoms of neon-22 have 10 protons and 12 neutrons. The number of protons added to the number of neutrons in an atom is called its **mass number**.



A mass spectrum for the element neon. The element neon is made up of two isotopes of different proportions. 90% of the element neon is made up of the isotope with relative mass of 20, while the remaining 10% is the isotope with relative mass of 22.

Relative atomic mass

The structure of an atom is identified using the convention ^A_ZE , where E is the symbol of the element, A is the mass number (number of protons + neutrons), and Z is the atomic number (number of protons). For example, the two isotopes of neon are represented as $^{20}_{10}\text{Ne}$ (atomic number, $Z = 10$; mass number, $A = 20$) and $^{22}_{10}\text{Ne}$ ($Z = 10$; $A = 22$).

Most elements consist of a mixture of isotopes. The **relative atomic mass** (A_r) of an element represents the average mass of one atom, taking into consideration the number of isotopes of the element, their relative isotopic mass (RIM) and their relative abundance. Using data from the mass spectrum, one could calculate the relative atomic mass for an element by using the following method of calculation:

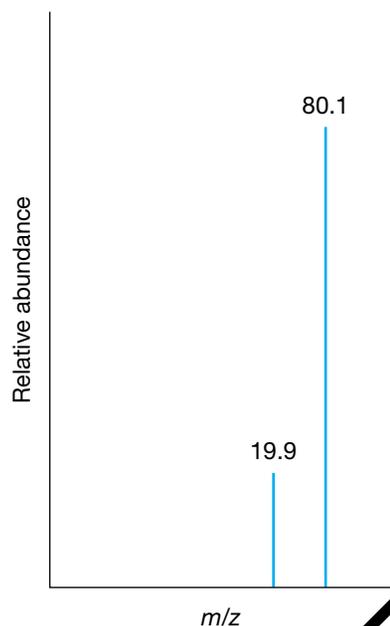
$$A_r = \frac{(\text{RIM of first isotope} \times \text{abundance}) + (\text{RIM of second isotope} \times \text{abundance}) + \dots}{100}$$

Because A_r is a relative mass, a unit is not required; however, the unit 'unified atomic mass' (symbol u) can be used. This replaces the old unit 'amu' (atomic mass unit).

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Interactivity

Relative isotopic mass
int-6353



Relative abundance of boron isotopes

For example, the two isotopes of neon, $^{20}_{10}\text{Ne}$ and $^{22}_{10}\text{Ne}$, have relative isotopic masses of 20 and 22 respectively. Their relative abundances are 90.0% and 10.0% respectively (see the graph on the previous page).

$$A_r(\text{Ne}) = \frac{(20 \times 90) + (22 \times 10)}{100} = 20.2$$

The relative molecular mass (M_r) of a molecule is the sum of the relative atomic masses, as shown in the periodic table, of elements in the formula. For example:

$$M_r(\text{NH}_3) = A_r(\text{N}) + (3 \times A_r(\text{H})) = 14.0 + (3 \times 1.0) = 17.0$$

Revision questions

- Why is carbon used as a standard to determine relative atomic masses?
- Lithium consists of two isotopes. One isotope, ^6_3Li , has a relative isotopic mass of 6.01 and an abundance of 7.42%. The other isotope, ^7_3Li , has a relative isotopic mass of 7.01 and an abundance of 92.58%.
 - Is the relative atomic mass closer to 6 or 7? Explain your answer.
 - Calculate the relative atomic mass of lithium.
- Copper is widely used for electrical wiring. It has two isotopes: $^{63}_{29}\text{Cu}$ and $^{65}_{29}\text{Cu}$. The lighter isotope has an abundance of 69.2%. Calculate the relative atomic mass of copper. (Note that, if accurate values of the relative isotopic masses are not provided, use mass numbers.)
- Three isotopes of magnesium and their relative abundances are $^{24}_{12}\text{Mg}$ (78.8%), $^{25}_{12}\text{Mg}$ (10.2%) and $^{26}_{12}\text{Mg}$ (11.0%).
 - Sketch on a graph the mass spectrum for magnesium.
 - Calculate the relative atomic mass of magnesium.
- Gallium has two isotopes. One isotope is $^{69}_{31}\text{Ga}$ and has a relative abundance of 60.50%. The relative atomic mass of gallium is 69.70. Find the relative isotopic mass of the other isotope.
- The green colour of fireworks can be produced using the element boron. Boron's two isotopes are $^{10}_5\text{B}$ and $^{11}_5\text{B}$. Use the mass spectrum on the left to calculate the relative atomic mass of boron.
- Calculate M_r of each of the following compounds.
 - CO_2
 - NaCl
 - H_2O_2
 - H_2SO_4
 - $\text{C}_6\text{H}_{12}\text{O}_6$

Counting atoms

In one gram of sugar there are about 1 760 000 000 000 000 000 000 or 1.76×10^{21} sugar molecules. If we count in dozens, that corresponds to 140 000 000 000 000 000 000 dozen molecules. Clearly, in chemistry, we need a measurement that can manage these large numbers of atoms and molecules. The mole concept helps us count atoms and is fundamental to chemical calculations.

The mole concept

The term **mole** (symbol mol) represents a number. It is a unit of measurement, just as the term 'dozen' represents 12 and 'kilo' represents 1000. Since there

Chemists use a unit called the mole to count atoms, ions and molecules. The mole contains a very large number of particles. This number is 6.02×10^{23} .

To find the mass of one mole of an element, simply add 'g' to the relative atomic mass of that element.

study on

Unit 1

The mole

AOS 1

Summary screen and practice questions

Topic 3

Concept 3

One mole each of mercury, zinc, silicon, aluminium, sulfur and bromine. This means that the quantity of each of the substances shown contains 6.02×10^{23} particles.

are billions of atoms in any substance, the mole must represent a very large number. Thus, a mole of hydrogen atoms means a certain large number of hydrogen atoms.

A mole is defined as the amount of substance that contains as many particles (atoms, ions or molecules) as there are atoms in exactly 12 g of the $^{12}_6\text{C}$ isotope.

The number of carbon atoms in 12 g of $^{12}_6\text{C}$ isotope has been experimentally estimated to be 602 000 000 000 000 000 000 000 atoms, or 6.02×10^{23} atoms. This is often called **Avogadro's number** (N_A). We can therefore say:

The number of atoms in one mole of an element is 6.02×10^{23} .

You may have noticed that the definition of the mole, like the definition of relative isotopic mass (see chapter 2), is based on the $^{12}_6\text{C}$ isotope. There is a good reason for this. The relative isotopic scale allows us to deal with atoms 'one at a time'. The mole concept allows us to deal with much larger groups 'one mole at a time'. The common reference to $^{12}_6\text{C}$ means that, to find out the mass of one mole (molar mass) of any element, all you have to do is add 'g' to the relative atomic mass.



Molar mass (M)

The molar mass (M) of an element is defined as the mass of 1 mol of the element: that is, 6.02×10^{23} atoms of the element. The unit is grams per mole (g mol^{-1}). For example:

Molar mass of carbon atoms = mass of 1 mol of C atoms = 12.0 g mol^{-1}

\therefore 12.0 g of carbon contains 6.02×10^{23} atoms of carbon.

Molar mass of oxygen atoms = mass of 1 mol of O atoms = 16.0 g mol^{-1}

\therefore 16.0 g of oxygen contains 6.02×10^{23} atoms of oxygen.

Molar mass of oxygen *molecules* = $2 \times$ mass of 1 mol of O atoms
= $2 \times 16.0 \text{ g mol}^{-1} = 32.0 \text{ g mol}^{-1}$

\therefore 32.0 g of oxygen contains 6.02×10^{23} molecules of oxygen.

32.0 g of oxygen contains 1.20×10^{24} atoms of oxygen.

Compounds and molar mass

The molar mass of a compound is defined as the mass of 1 mol of the compound expressed in grams per mole (g mol^{-1}).

study on

Unit 1

Molar mass

AOS 1

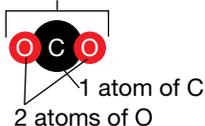
Summary screen and practice questions

Topic 3

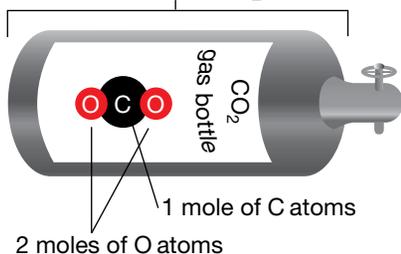
Concept 4

The molar mass of a compound is the mass of one mole of that compound. It is found by adding together all the relative atomic masses for the atoms in its formula and adding 'g mol⁻¹' as the unit.

1 molecule of CO₂



1 mole of CO₂



Molecular compounds

For compounds consisting of molecules, the molar mass (M) is numerically equal to the relative molecular mass (M_r) expressed in grams per mole (g mol⁻¹). For example:

- Molar mass of water molecules = mass of 1.00 mol of H₂O molecules = 18.0 g mol⁻¹
∴ 18.0 g of water contains 6.02×10^{23} molecules of water.
- Molar mass of chlorine gas = mass of 1.00 mol of Cl₂ molecules = 71.0 g mol⁻¹
∴ 71.0 g of chlorine gas contains 6.02×10^{23} molecules of chlorine.

Ionic compounds

The M_r of an ionic compound is found by adding the A_r of each atom in the formula of the compound. For example:

$$\begin{aligned} M_r(\text{CuSO}_4) &= A_r(\text{Cu}) + A_r(\text{S}) + (4 \times A_r(\text{O})) \\ &= 63.5 + 32.1 + (4 \times 16.0) \\ &= 159.6 \end{aligned}$$

The molar mass (M) of CuSO₄ is 159.6 g mol⁻¹.

$$M(\text{CuSO}_4) = 159.6 \text{ g mol}^{-1}$$

Whether a substance is made up of atoms, molecules or ions, the same principle applies: one mole of any substance always contains 6.02×10^{23} particles. This number is always 6.02×10^{23} .

Revision question

8. Calculate the molar masses of the following.
- | | |
|----------------------------|----------------------------|
| (a) ozone, O ₃ | (c) calcium oxide, CaO |
| (b) iodine, I ₂ | (d) hydrochloric acid, HCl |

Changing moles to numbers of particles

We have described a relationship between the number of particles and the mass of the substance. This means that, when we measure the mass of a substance equal to its molar mass (e.g. 18.0 g of H₂O), we are also indirectly counting out 6.02×10^{23} particles of that substance (e.g. water molecules). If the molecule or formula unit contains more than one kind of atom, we can also calculate the number of different atoms present in one mole of the substance. For example:

1.00 mol of carbon dioxide molecules contains:

- 1.00 mol of CO₂ molecules
- 6.02×10^{23} CO₂ molecules
- 1.00 mol of C atoms
- 6.02×10^{23} C atoms
- 2.00 mol of O atoms
- 1.20×10^{24} O atoms.

The molar mass of CO₂ is 44.0 g mol⁻¹.

These observations suggest the following:

$$\text{mass} \begin{array}{c} \xrightarrow{\text{divide by molar mass}} \\ \xleftarrow{\text{multiply by molar mass}} \end{array} \text{moles}$$

or

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

where n = number of moles, m = mass, M = molar mass

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Interactivity

The mole concept
int-6354

Weblink

What is mole day?

The formula $n = \frac{m}{M}$, where n is the number of moles, m is the mass of the sample and M is the molar mass, can be used to calculate the number of moles of an element or a compound.

The formula

$$n = \frac{\text{number of particles}}{6.02 \times 10^{23}}$$

can be used to calculate the number of moles.

and

$$\text{number of particles (} N \text{)} \begin{array}{l} \xleftarrow{\text{divide by } 6.02 \times 10^{23}} \\ \xrightarrow{\text{multiply by } 6.02 \times 10^{23}} \end{array} \text{ moles}$$

or

$$n = \frac{\text{number of particles}}{6.02 \times 10^{23}} = N/N_A$$

Sample problem 5.1

A gas balloon contains 5.5 mol of helium atoms. How many helium atoms are present?

Solution:

$$\begin{aligned} \text{number of particles} &= n(\text{He}) \times 6.02 \times 10^{23} \\ &= 5.5 \times 6.02 \times 10^{23} \\ &= 3.3 \times 10^{24} \text{ atoms of helium} \end{aligned}$$

Revision questions

- Write a brief explanation of the mole concept, and explain the importance of Avogadro's number.
- Calculate the number of chlorine atoms or chloride ions in:
 - 2.3 mol of chlorine atoms
 - 15.8 mol of chlorine molecules, Cl_2
 - 3.5 mol of sodium chloride, NaCl
 - 0.50 mol of magnesium chloride, MgCl_2 .
- Determine how many:
 - mol of ethanoic acid molecules
 - molecules of the acid
 - mol of oxygen atoms
 - atoms of oxygenare present in 16.2 g of ethanoic acid, CH_3COOH .

Sample problem 5.2

Pure liquid ammonia is called anhydrous ammonia and is used extensively as a fertiliser as it has a high nitrogen content. If we have 4.6×10^{28} molecules of ammonia, NH_3 , in a fertiliser plant store, how many moles does this represent?

Solution:

$$\begin{aligned} \text{number of moles (} \text{NH}_3 \text{)} &= \frac{\text{number of } \text{NH}_3 \text{ molecules}}{6.02 \times 10^{23}} \\ &= \frac{4.6 \times 10^{28}}{6.02 \times 10^{23}} \\ &= 7.6 \times 10^4 \end{aligned}$$

Therefore, 7.6×10^4 mol of ammonia molecules is stored.

Revision question

- Calculate the number of moles of each of the following particles present in 5.2×10^{24} molecules of methane, CH_4 .
 - methane molecules
 - carbon atoms
 - hydrogen atoms

Changing mass to moles and moles to mass

Since one mole of a substance refers to both a *mass* of the substance and a number of particles of that substance (moles), we can calculate one from the other. There are many different ways to refer to the composition of a substance.

Sample problem 5.3

Ethanol, C_2H_5OH , is found in alcoholic beverages. If one such beverage contains 4.6 g of ethanol, how many moles does this represent?

Solution:

$$\begin{aligned} n(\text{ethanol}) &= \frac{\text{mass}(\text{ethanol})}{M(\text{ethanol})} \\ &= \frac{4.6}{46.0} \\ &= 0.10 \text{ mol} \end{aligned}$$

Revision question

13. Calculate the amount, in moles, of:
- 46 g of water, H_2O
 - 2.4 g of carbon dioxide, CO_2
 - 67 g of chlorine gas, Cl_2
 - 2.0 g of sodium chloride, $NaCl$
 - 128 g of copper(II) sulfate pentahydrate, $CuSO_4 \cdot 5H_2O$
 - 38 kg of iron(III) oxide, Fe_2O_3 .

Sample problem 5.4

Pure ethanoic acid, CH_3COOH , can be used to make vinegar when dissolved in water. If 3.5 moles of ethanoic acid was used, what mass was weighed out?

Solution:

$$\begin{aligned} \text{Since } n &= \frac{m}{M}, m = n \times M. \\ \text{mass}(\text{ethanoic acid}) &= n(\text{ethanoic acid}) \times M(\text{ethanoic acid}) \\ &= 3.5 \times 60.0 \\ &= 210 \text{ g} \end{aligned}$$

Therefore, 2.1×10^2 g of ethanoic acid was weighed out.

Revision questions

14. Calculate the mass of:
- 0.41 mol of carbon monoxide, CO
 - 12.0 mol of sulfur dioxide, SO_2
 - 3.84 mol of sucrose, $C_{12}H_{22}O_{11}$
 - 58.2 mol of iron, Fe
 - 0.0051 mol of silver chloride, $AgCl$
 - 2.53 mol of magnesium phosphate, $Mg_3(PO_4)_2$.
15. Which of the following substances has the greatest mass?
- 2.5 mol of hydrogen gas, H_2
 - 0.2 mol of zinc, Zn
 - 11.56 g of calcium chloride, $CaCl_2$

Sample problem 5.5

What is the mass of 3.01×10^{23} molecules of hydrochloric acid, HCl?

Solution:

$$n(\text{HCl}) = \frac{3.01 \times 10^{23}}{6.02 \times 10^{23}}$$

$$= 0.500 \text{ mol}$$

$$\text{Since } n = \frac{m}{M}, m = n \times M$$

$$\therefore \text{mass}(\text{HCl}) = n \times M$$

$$= 0.500 \times 36.5$$

$$= 18.3 \text{ g}$$

Revision questions

16. What is the mass of each of the following?
- 5.25×10^{24} molecules of glucose, $\text{C}_6\text{H}_{12}\text{O}_6$
 - 1.83×10^{21} molecules of nitrogen dioxide, NO_2
 - 3.56×10^{14} molecules of carbon dioxide, CO_2
 - 4.13×10^{28} molecules of carbon disulfide, CS_2
 - 3.62×10^{24} molecules of dinitrogen tetroxide, N_2O_4
17. Which of the following substances has the greatest mass?
- 200 g of magnesium
 - 5.00 mol of sulfur
 - 1.2×10^{24} atoms of helium
 - 3.5×10^{22} molecules of alanine, $\text{C}_3\text{H}_7\text{O}_2\text{N}$
18. Use a spreadsheet program to produce a spreadsheet that converts amounts of substances to moles. Set up your spreadsheet according to the following template, and then complete the table. Use a formula and then the fill function.

Conversion of amounts of substances to moles

	A	B	C	D	E	F	G
	Substance	Amount (g)	Molar mass (M)	Number of atoms in the molecule	Number of moles (n) of substance	Number of molecules	Total number of atoms
1	water, H_2O	3.2					
2	methane, CH_4	2.7					
3	ammonia, NH_3	0.056					
4	acetic acid, CH_3COOH				27.3		
5	benzene, C_6H_6				0.56		
6	octane, C_8H_{18}				2.34		
7	ethanol, $\text{CH}_3\text{CH}_2\text{OH}$					6.0×10^{24}	
8	ozone, O_3					1.27×10^{21}	
9	sulfuric acid, H_2SO_4						3.0×10^{26}
10	carbon dioxide, CO_2						7.5×10^{25}

Percentage composition

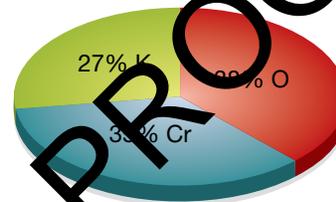
The percentage that each element contributes to the total mass of a compound is called the percentage composition.

The chemical formula of a compound tells us about the relative numbers of atoms in its constituent elements. However, the composition of a compound is often expressed in terms of the percentage that each element contributes to its mass. This is called **percentage composition**. For instance, although potassium chromate, K_2CrO_4 , and potassium dichromate, $\text{K}_2\text{Cr}_2\text{O}_7$, have the same number of potassium atoms, the percentage that potassium contributes to the mass of a unit of the chromate is higher than its contribution to a unit of the dichromate.

The percentages of K, Cr and O in (a) potassium chromate, K_2CrO_4 and (b) potassium dichromate, $\text{K}_2\text{Cr}_2\text{O}_7$



(a) potassium chromate, K_2CrO_4



(b) potassium dichromate, $\text{K}_2\text{Cr}_2\text{O}_7$

The following formula can be used to calculate percentage composition:

$$\% \text{ of } x \text{ in compound containing } = \frac{\text{mass of } x \text{ in 1 mol of compound}}{\text{molar mass of compound}} \times \frac{100}{1}$$

Sample problem 6

Stannous (tin) fluoride, SnF_2 , is an active ingredient in some toothpastes. Find the percentage composition by mass of each element in the compound.

Solution: One formula unit of SnF_2 contains one Sn atom and two F atoms.

$$\begin{aligned} M_r(\text{SnF}_2) &= 118.7 + (2 \times 19.0) \\ &= 156.7 \end{aligned}$$

Since Sn contributes 118.7 units toward this mass:

$$\begin{aligned} \% \text{ mass Sn} &= \frac{118.7}{156.7} \times \frac{100}{1} \\ &= 75.7\% \end{aligned}$$

F_2 contributes 2×19.0 units, so we get:

$$\begin{aligned} \% \text{ mass F} &= \frac{2 \times 19.0}{156.7} \times \frac{100}{1} \\ &= 24.3\% \end{aligned}$$

Revision question

19. Calculate the percentage composition of each element in the following compounds.
- (a) propane, C_3H_8
 - (b) sodium hydrogen sulfate, NaHSO_4
 - (c) calcium acetate, $\text{Ca}(\text{CH}_3\text{COO})_2$
 - (d) hydrogen cyanide, HCN

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Interactivity

Percentage composition
int-6355

Study on

Unit 1

Percentage composition by mass

AOS 1

Topic 3

Summary screen and practice questions

Concept 5

Calculating percentage composition of hydrated compounds

Ionic compounds often incorporate water into their lattice structure to form hydrated compounds. For many hydrated ionic compounds, this number of water molecules is very specific, and is called the 'water of crystallisation.' The actual number of water molecules in the formula is called the 'degree of hydration.'

Some ionic compounds crystallise from an aqueous solution to form a **hydrated** ionic compound. In these compounds, water molecules are included in the crystal lattice structure. This water is called **water of crystallisation**. Hydrated copper(II) sulfate, for example, appears as a blue crystalline solid and has the formula $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$. This means that, for each Cu^{2+} ion and SO_4^{2-} ion in the crystal lattice structure, five water molecules are also included. Heating the copper(II) sulfate crystals removes the water molecules to leave a white powder known as **anhydrous** copper(II) sulfate. The formula of anhydrous copper(II) sulfate is CuSO_4 . The mass of the water of crystallisation, and its percentage contribution to the mass of the ionic compound, can be calculated when the masses of both the hydrated and anhydrous compound are known or if the degree of hydration is shown in the formula.

Hydrated copper(II) sulfate,
 $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$



Many structures have water molecules included in their composition. When the structures are dehydrated, the water is released. Dehydrated copper sulfate is white in colour. When water is added, it turns the characteristic blue colour of hydrated copper sulfate.

Sample problem 5.7

Calculate the percentage of water by mass in $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$.

Solution: **STEP 1**

Find the molar mass, making sure to include the water present in the formula.

$$\begin{aligned}M(\text{CuSO}_4 \cdot 5\text{H}_2\text{O}) &= 63.5 + 32.1 + (4 \times 16.0) + 5([2 \times 1.0] + 16.0) \\ &= 249.6 \text{ g mol}^{-1}\end{aligned}$$

STEP 2

Calculate the mass of the water.

$$\begin{aligned}M &= 5([2 \times 1.0] + 16.0) \\ &= 90.0 \text{ g mol}^{-1}\end{aligned}$$

STEP 3

Calculate the percentage mass of water.

$$\begin{aligned}\% \text{ mass water} &= \frac{90.0}{249.6} \times \frac{100}{1} \\ &= 36.07\% = 36.1\%\end{aligned}$$

Revision question

20. Calculate the percentage of water in each of the following compounds.
- nickel(II) sulfate hexahydrate, $\text{NiSO}_4 \cdot 6\text{H}_2\text{O}$
 - sodium carbonate decahydrate, $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$
 - magnesium chloride hexahydrate, $\text{MgCl}_2 \cdot 6\text{H}_2\text{O}$

Empirical formulas

The empirical formula of a compound is the simplest whole number ratio of atoms or ions that are present in that compound. Empirical formulas can be calculated from percentage compositions or molecular formulas.

The **empirical formula** of a compound gives the simplest whole number ratio of the atoms or ions present in the compound and can be found only by experiment.

To determine the empirical formula of a compound, an experimentally determined *ratio of elements by mass* must be converted to a *ratio of elements by numbers*. This is done by calculating the number of moles of each element.

The steps involved in finding an empirical formula are:

- Write down the *symbols* of the elements present.
- Assume that the *mass* of the sample is 100 g and all percentages become grams.
- Convert masses to *moles*.
- Find the simplest whole number *ratio* of the atoms by dividing all numbers of moles by the smallest number of moles.
- If necessary, multiply by a *factor* to convert all numbers to whole numbers.

Sample problem 5.8

A compound of sulfur contains 2.4% hydrogen, 39.0% sulfur and 58.6% oxygen. Find the empirical formula of the compound.

Solution: Follow the steps outlined above.

STEP 1	symbols	H	S	O
STEP 2	masses	2.4 g	39.0 g	58.6 g

study on

Unit 1

Empirical formula of an ionic compound

Topic 3

Summary screen and practice questions

Concept 6

	$\frac{2.4}{1.0}$	$\frac{39.0}{32.1}$	$\frac{58.6}{16.0}$
STEP 3 moles	= 2.4	= 1.2	= 3.7
	$\frac{2.4}{1.2}$	$\frac{1.2}{1.2}$	$\frac{3.7}{1.2}$
STEP 4 ratio/divide by smallest	= 2.0	= 1.0	= 3.1

The empirical formula of the compound is H_2SO_3 .

Empirical formulas can also be calculated when experiments reveal the actual mass of each element that is present in a sample of a compound. In this situation, the masses involved are simply written in at step 1 in the sequence of steps on the previous page.

Revision questions

- Aspirin is a drug used extensively for pain relief. Chemical analysis of an aspirin tablet determined that it was composed of 57.7% carbon, 37.5% oxygen and 4.8% hydrogen. Calculate the empirical formula of aspirin.
- Hydroquinone is a liquid that is used as a photographic developer. Determine the empirical formula of hydroquinone given that it contains 65.4% carbon and 29.1% oxygen with the remainder being hydrogen.

Sample problem 5

Washing soda crystals may be used to bleach linen. When crystallised from water, washing soda (sodium carbonate, Na_2CO_3) forms crystals of a hydrated ionic compound. When 5.00 g of washing soda crystals were dried in a desiccator, 1.85 g of sodium carbonate remained. Calculate the empirical formula of the hydrated compound.

Solution:

STEP 1

Calculate the mass of water present in the hydrated compound. Since 1.85 g of sodium carbonate was obtained on dehydration of the crystals, the amount of water is 3.15 g, or the difference in mass between the hydrated and dehydrated compounds ($5.00 - 1.85$) = 3.15 g.

STEP 2

The mole ratio of sodium carbonate to water is:

$$\begin{aligned} & \frac{1.85}{106.0} : \frac{3.15}{18.0} \\ & = 0.0175 : 0.175 \\ & = 1 : 10 \end{aligned}$$

So the empirical formula of the washing soda crystals is $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$.

Revision questions

- A 1.124 g sample of CaSO_4 crystals was heated to drive off the water of crystallisation. When completely dry, a residue of 0.889 g was obtained. Determine the empirical formula of the hydrated compound.
- A 0.942 g sample of MgSO_4 crystals was heated to drive off the water of crystallisation. When completely dry, a residue of 0.461 g was obtained. Determine the empirical formula of the hydrated compound.

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Experiment 5.1 Empirical formula of magnesium oxide
doc-15998

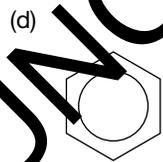
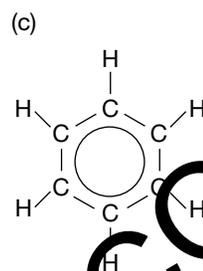
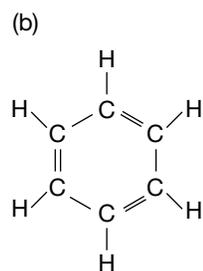
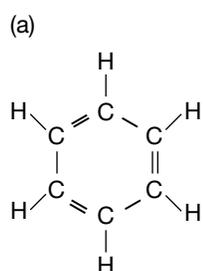


Sodium carbonate can be used to bleach linen.

The molecular formula of a compound is the actual number of atoms that are present in a molecule of that substance. It can be equal to the empirical formula, or it can be a whole number multiple of the empirical formula.

Note that only molecular compounds can have a molecular formula.

Four structural representations of a benzene molecule. A benzene molecule is the basis of a group of chemicals called aromatics.



Molecular formulas

The **molecular formula** of a compound represents the actual composition of a compound that is made up of molecules. A molecular formula is either the same as its experimentally determined empirical formula, or is a whole-number multiple of it. Table 5.1 shows the empirical and molecular formulas of some common hydrocarbons. Although their empirical formulas may be the same, compounds with different molecular formulas may have very different properties. For example, acetylene is a gas that is used in welder's torches, whereas benzene is a highly flammable liquid that is a recognised carcinogen. Although both compounds have the same empirical formula, CH , the molecular formula of acetylene is C_2H_2 , and the molecular formula of benzene is C_6H_6 .

The molecular formula of a compound may be determined from its empirical formula only if its molar mass is also known.

$$n \times (\text{empirical formula}) = \text{molecular formula}$$

where n represents a whole number.

TABLE 5.1 Comparison of empirical and molecular formulas

Name of molecule	Empirical formula (simplest ratio of atoms in molecule)	Molecular formula (actual number of atoms in molecule)
ethyne (acetylene)	CH	C_2H_2
benzene	CH	C_6H_6
formaldehyde	CH_2O	CH_2O
ethanoic acid (acetic acid)	CH_2O	$\text{C}_2\text{H}_4\text{O}_2$
glucose	CH_2O	$\text{C}_6\text{H}_{12}\text{O}_6$

Sample problem 5.10

Benzene has the empirical formula CH and its molar mass is 78 g mol^{-1} . Find the molecular formula.

Solution

STEP 1 If the molar mass is 78.0 g mol^{-1} , then its relative molecular mass (the sum of the relative atomic masses according to the molecular formula) is also 78.0. If the empirical formula is written as C_xH_y , then the molecular formula is $(\text{C}_x\text{H}_y)_n$, where x , y and n are whole numbers. The value of n may be determined by comparing the relative molecular mass and the empirical formula mass.

STEP 2 The empirical formula is CH .

$$\begin{aligned} \text{empirical formula mass} &= A_r(\text{C}) + A_r(\text{H}) \\ &= 12.0 + 1.0 \\ &= 13.0 \end{aligned}$$

STEP 3

$$\frac{\text{molecular mass}}{\text{empirical formula mass}} = \frac{78.0}{13.0}$$

ratio = 6

STEP 4 \therefore molecular formula = $6 \times$ empirical formula

$$\begin{aligned} &= 6 \times \text{CH} \\ &= \text{C}_6\text{H}_6 \end{aligned}$$

Revision questions

- During the decay of animal tissues, a noxious compound called putrescine may be produced. Putrescine has the empirical formula C_2H_6N and a relative molecular mass of 88.0. What is the molecular formula of putrescine?
- Calculate the molecular mass of each compound listed in table 5.1. What is the mathematical relationship between the empirical and molecular formulas of a substance?
- Nicotine, the main active chemical in tobacco, has the empirical formula C_5H_7N and a relative molecular mass of 162.0. Determine the molecular formula of nicotine.

Sample problem 5.11

The common insect repellent sold commercially as 'moss balls' is the organic compound naphthalene. It is a hydrocarbon containing 93.7% carbon and 6.3% hydrogen and has a molar mass of 128.0 g mol^{-1} . Find the empirical and molecular formulas of naphthalene.

Solution:

A molecular formula is calculated in the same way as an empirical formula. However, it needs one additional piece of information — the molar mass. This allows a comparison to be made with the empirical formula to determine the whole number multiple.

Calculate the empirical formula using the five steps listed on page 102; then calculate the empirical formula mass and use it to find the molecular formula.

STEP 1 H C

STEP 2 6.3 g 93.7 g

STEP 3 $\frac{6.3}{1.0}$ $\frac{93.7}{12.0}$
 $= 6.3$ $= 7.81$

STEP 4 $\frac{6.3}{6.3}$ $\frac{7.81}{6.3}$
 $= 1.0$ $= 1.24$

STEP 5 1.0×4 1.24×4
 $= 4.0$ $= 5.0$

The empirical formula of naphthalene is C_5H_4 .

STEP 6 empirical formula mass = $A_r(C) + A_r(H)$
 $= (5 \times 12.0) + (4 \times 1.0)$
 $= 64.0$

STEP 7 $\frac{\text{molecular mass}}{\text{empirical formula mass}} = \frac{128.0}{64.0}$
 $\text{ratio} = 2$

STEP 8 molecular formula = $2 \times$ empirical formula
 $= 2 \times C_5H_4$
 $= C_{10}H_8$

Revision questions

- Caffeine is a stimulant that is found naturally in coffee, tea and chocolate. Analysis of caffeine shows that it contains 49.5% carbon, 28.9% nitrogen, 16.5% oxygen and 5.1% hydrogen by mass. Determine the molecular formula of caffeine given that its molar mass is 194.2 g mol^{-1} .
- The compound methyl butanoate smells like apples. Its percentage composition is 58.8% C, 9.8% H and 31.4% O. If its molar mass is 102.0 g mol^{-1} , what is its molecular formula?

Chapter review

Summary

- The relative atomic mass (A_r) of elements shown in the periodic table is calculated using the weighted mean of the relative isotopic masses:

$$A_r = \frac{(\text{RIM of first isotope} \times \text{abundance}) + (\text{RIM of second isotope} \times \text{abundance}) + \dots}{100}$$

where A_r is the relative atomic mass and RIM is the relative isotopic mass.

- Relative isotopic masses are determined using a mass spectrometer and are masses compared with the carbon-12 isotope.
- The carbon-12 isotope is defined as having a mass of 12 g exactly. It is used as a standard as it is a stable and abundant isotope, which means that its mass can be determined accurately.
- The mass spectrometer is an instrument that can separate isotopes of an element based on their mass-to-charge ratio.
- A mass spectrum of an element shows the number of isotopes present in an element, their relative isotopic masses and their proportions.
- The mole concept is an important chemical idea, enabling us to 'count' particles present in a quantity of a substance.
 - The unit that measures the amount of a substance is the mole (mol).
 - A mole of any substance is composed of Avogadro's number (6.02×10^{23}) of representative particles.
 - The molar mass (M) of an element or compound is the mass of 1 mole of the substance (numerically equal to the relative molecular mass (M_r) and is expressed in g mol⁻¹).
 - The number of particles in a sample of substance may be determined by the formula:

$$\text{number of particles} = \text{number of moles of substance} \times \text{Avogadro's number}$$

- The percentage composition of a compound is the percentage by mass of each element in the compound:

$$\% \text{ of } x \text{ in compound containing } x = \frac{\text{mass of } x \text{ in compound}}{\text{molar mass of compound}} \times \frac{100}{1}$$

Hydrated ionic compounds contain water molecules as part of their lattice structure. This water is called water of crystallisation. When the mass of both the hydrated and anhydrous (dried) compounds is known, the percentage composition by mass of water in the hydrated compound can be determined.

- Formulas for compounds take two forms:
 - An empirical formula is the simplest whole number ratio of atoms of the elements in the compound and is determined experimentally.
 - A molecular formula is the same as, or some simple multiple of, an empirical formula and specifies the exact number of each type of atom in a molecule of the compound.

Multiple choice questions

- The particles in a mass spectrometer are deflected according to their:

A abundance only	C charge only
B mass only	D mass and charge.
- Rubidium has two stable isotopes. Rubidium-85 has an abundance of 72% and rubidium-87 has an abundance of 28%. The mass spectrum of rubidium has:

A 1 peak: a smaller peak at RIM 85 and a larger peak at RIM 87
B 2 peaks: a smaller peak at RIM 87 and a larger peak at RIM 85
C 1 peak at about RIM 85.5
D 1 peak at about RIM 86.
- When the element fluorine, F_2 , is passed through the mass spectrometer the particles detected are:

A F_2 ions	C F^- and F^+ ions
B F ions	D F^+ ions.
- A certain element has two isotopes of relative isotopic mass 203 and 205. Given that its A_r value is 204.4, which of the following is the most likely percentage abundance of the heavier isotope?

A 7%
B 30%
C 50%
D 70%
- The difference between the relative molecular mass (M_r) of carbon dioxide and the molar mass of carbon dioxide is that the relative molecular mass and the molar mass of carbon dioxide are, respectively:

A 44.0 g and 1 mol
B 44.0 and 44.0 g mol ⁻¹
C 44.0 g and 44.0 g mol ⁻¹
D 44.0 and 44.0 g.
- The number of moles of oxygen atoms in 143 g of sodium carbonate decahydrate, $Na_2CO_3 \cdot 10H_2O$, is:

A 0.50
B 1.5
C 2.0
D 6.5.

7. The *smallest* number of molecules would be contained in:
- 1.0 g of N_2
 - 1.0 g of O_2
 - 1.0 g of NO
 - 1.0 g of NO_2 .
8. The *largest* number of oxygen atoms would be found in:
- 300 g of water, H_2O
 - 3.2 mol of hydrated copper(II) sulfate, $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$
 - 3 kg of a fat having the molecular formula $\text{C}_{57}\text{H}_{110}\text{O}_6$
 - 7.35×10^{24} molecules of nitrogen dioxide, NO_2 .
9. Polymers may be made from the three monomers acrylic acid, $\text{C}_3\text{H}_4\text{O}_2$, methyl acrylate, $\text{C}_4\text{H}_6\text{O}_2$, and acrylonitrile, $\text{C}_3\text{H}_3\text{N}$. The order of increasing percentage carbon content of these monomers is:
- acrylic acid, methyl acrylate, acrylonitrile
 - acrylic acid, acrylonitrile, methyl acrylate
 - acrylonitrile, acrylic acid, methyl acrylate
 - acrylonitrile, methyl acrylate, acrylic acid.
10. Which of the following lists of substances shows only empirical formulas?
- H_2O , CH_3COOH , HCl , $\text{Na}_2\text{Cr}_2\text{O}_7$
 - H_2SO_4 , H_2O , $\text{Al}(\text{NO}_3)_2$, $\text{Al}_2(\text{SO}_4)_3$
 - HCl , $\text{C}_6\text{H}_{12}\text{O}_6$, NH_3 , $\text{Al}(\text{NO}_3)_2$
 - $\text{Al}_2(\text{SO}_4)_3$, O_2 , HCl , H_2SO_4
11. A substance, on analysis, was found to contain 27.3% carbon and 72.7% oxygen. Which of the following is the empirical formula of the substance?
- CO_2
 - CO_3
 - C_2O_4
 - C_3O_7
12. A hydrocarbon that forms the major component of liquefied petroleum gas was determined experimentally to contain 16.3% hydrogen by mass. If the hydrocarbon has a relative molecular mass of 36, its molecular formula may be represented as
- C_3H_6
 - C_6H_{14}
 - C_7H_{12}
 - C_7H_{16} .

Review questions

Relative atomic mass

- Silicon-containing ores have three isotopes: 92% silicon-28, 5% silicon-29 and 3% silicon-30.
 - Draw the mass spectrum for silicon.
 - What is the atomic number of silicon?
 - What are the relative isotopic masses of the three silicon isotopes?
 - What is the relative atomic mass of silicon?
2. Silver is an unreactive metal that is used to produce jewellery. Its relative atomic mass is 107.9, and it consists of two naturally occurring isotopes: silver-107 and silver-109. Calculate the abundance of the lighter isotope.

Mole calculations

- In 2.0 moles of $(\text{NH}_4)_3\text{PO}_4$ (fertiliser) there are
 - _____ moles of nitrogen atoms
 - _____ moles of hydrogen atoms
 - _____ moles of oxygen atoms
 - _____ moles of phosphorus atoms
 - _____ (total number) nitrogen atoms
 - _____ grams of phosphate ions
 - _____ grams of nitrogen atoms.
- Copy and complete the table below.

Name	Formula	Molar mass (M)	Mole (n)	Mass (m)
sodium hydroxide	NaOH			3.41 g
	CCl_4		1.40 mol	
sodium carbonate		106 g mol^{-1}	1.00 mol	
	KCl		0.25 mol	
ammonium phosphate				8.46 g

- Find the number of moles *and* the number of atoms present in the following samples.
 - 14.6 g acetylene (ethyne), C_2H_2
 - 0.48 g propane, C_3H_8
 - 485 g ethanol, $\text{C}_2\text{H}_5\text{OH}$
 - 8.6 g carbon dioxide
 - 67 g iodine, I_2
- Calculate the mass of each of the following.
 - 160 mol Fe
 - 0.075 mol silicon dioxide
 - 4.23 mol NO_2
- Sodium fluoride is thought to reduce tooth decay, especially in children. It is therefore added to some brands of toothpaste. If a tube of toothpaste contains 0.013 g of sodium fluoride,
 - how many moles of sodium fluoride does this represent
 - how many fluoride ions does this represent?
- To prevent a gum disease called scurvy, the minimum daily requirement of vitamin C, $\text{C}_6\text{H}_8\text{O}_6$, is 60 mg.
 - How many moles of vitamin C is this?
 - How many molecules is this?

- (c) If 10 g of spinach is found to contain 1.2×10^{-5} g of vitamin C, how much spinach must be eaten to attain the minimum daily requirement?
9. A typical vitamin C tablet contains 500.0 mg of ascorbic acid. (The molecular formula of vitamin C is $C_6H_8O_6$.)
- How many moles of vitamin C does a typical vitamin C tablet contain?
 - How many molecules of vitamin C does a typical vitamin C tablet contain?
10. Too much cholesterol, $C_{27}H_{46}O$, is associated with heart disease, although the body produces its own cholesterol in the liver. A 250 g sample of cholesterol was removed from the arteries of a patient.
- How many moles of cholesterol are present in this sample?
 - How many moles of carbon atoms does this represent?
 - How many grams of carbon atoms are present in the sample?
 - How many grams of oxygen atoms are present in the sample?
11. Diamond is a naturally occurring form of pure carbon. The mass of a diamond is measured in a unit called a carat, where 1.00 carat = 0.200 g. How many atoms of carbon are in:
- a 1.00-carat diamond
 - a 3.15-carat diamond?

Percentage composition

12. Calculate the percentage composition of all the elements in:
- ethanoic acid, CH_3COOH
 - hydrated magnesium chloride, $MgCl_2 \cdot 6H_2O$
 - $Fe_2(SO_4)_3$.
13. Photocells use a semiconducting material that produces an electric current on exposure to light. Compounds of cadmium such as CdS, CdSe and CdTe are used in many common photocells. Calculate the percentage, by mass, of cadmium in:
- CdS
 - CdSe
 - CdTe.
14. A number of different compounds may be formed when oxygen and nitrogen react together. Calculate the percentage, by mass, of nitrogen and oxygen in each of the following oxides of nitrogen.
- NO , a colourless gas that is formed in internal combustion engines
 - NO_2 , a brown gas that is mainly responsible for the brown colour of the photochemical smog that hangs over many industrialised cities
 - N_2O_4 , a colourless liquid that is used as a fuel in space shuttles

(d) N_2O , a colourless gas (commonly called laughing gas) that is used as a dental anaesthetic

15. The synthetic narcotic methadone is used for treatment of heroin addiction and has the molecular formula $C_{21}H_{27}NO$. Calculate:
- the molar mass of methadone
 - its percentage composition.
16. Nitrogen is essential for plant growth. Ammonium nitrate, NH_4NO_3 , and urea, CON_2H_4 , can be used as fertilisers since each contains a significant proportion of nitrogen.
- Calculate the percentage, by mass, of nitrogen in each of the fertilisers.
 - Would you expect all compounds with a high nitrogen content to be suitable for use as a fertiliser? Justify your response.

Empirical formula

17. Coffee contains the stimulant caffeine. Analysis shows it consists of 49.48% carbon, 5.19% hydrogen, 28.85% nitrogen and 16.48% oxygen by mass. Calculate the empirical formula of caffeine.
18. The amino acid cysteine contains the elements carbon, hydrogen, nitrogen, oxygen and sulfur. Analysis of a 1.210 g sample of cysteine shows it to contain 0.0704 g of hydrogen, 1.80×10^{22} carbon atoms, 0.0100 mol of nitrogen and equal masses of sulfur and oxygen. Determine the empirical formula of cysteine.
19. Methanol is an alternative fuel to petrol that has been used in experimental cars. Determine the empirical formula of methanol, given that it is composed of 49.9% oxygen, 37.5% carbon and 12.6% hydrogen.
20. Borax is a naturally occurring compound that is used in the manufacture of optical glasses. Analysis of borax shows that it is made up of three elements: sodium (22.8%), boron (21.5%) and oxygen. Determine the empirical formula of borax.

Molecular formula

21. A compound of Na, S and O contains 17.04% Na and 47.41% S. The M_r of the compound is 270.0. Calculate the empirical formula and the molecular formula.
22. The taste of sour milk is due to lactic acid. The percentage composition of lactic acid by mass is 40.00% carbon, 6.71% hydrogen and 53.29% oxygen, and the molar mass is 90.0 g mol^{-1} . Find the empirical formula and molecular formula of lactic acid.
23. A compound contains 12.8% carbon and 2.13% hydrogen, the rest being bromine. The relative molecular mass of the compound is 188.0. Calculate the empirical formula and the molecular formula of the compound.

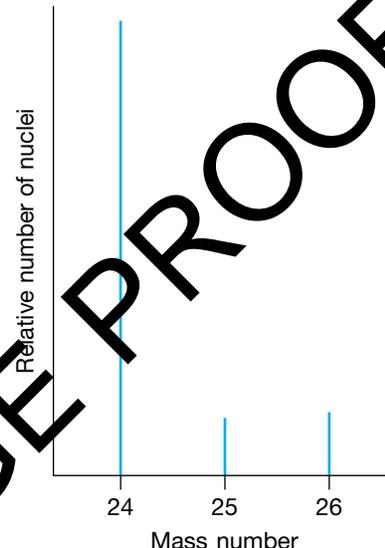
Exam practice questions

In a chemistry examination you will be required to answer a number of short and extended response questions.

Multiple choice question

1. The mass spectrum of magnesium is shown on the right. Which of the following statements is true?
- A Three isotopes are shown with masses in the ratio 24 : 25 : 26 and relative abundances of about 8%, 1% and 1%.
 - B The three isotopes have the same mass number but different atomic numbers.
 - C Isotopes 25 and 26 have relative abundances of about 10%.
 - D Isotope 24 has a relative abundance of about 10%.

1 mark



Extended response questions

2. Octane, C₈H₁₈, is a major component of petrol.

- (a) What is the M_r of octane?
- (b) What is the mass of 3.20 mol of octane?
- (c) How many molecules are there in 2.5 g of octane?
- (d) How many individual atoms are there in 5.0 g of octane?
- (e) What is the percentage of hydrogen in octane?
- (f) What mass of carbon would be present in 150 g of octane?

6 marks

3. The odour of rancid butter is caused by butyric acid, which contains 54.5% carbon and 9.1% hydrogen, the rest being oxygen. Given that the M_r of butyric acid is 88.0, calculate the empirical and molecular formulas of the acid.

4 marks

4. An oxide of copper is heated in a stream of hydrogen until only the copper remains, according to the equation:



The data for the experiment is given in the table below. Calculate the empirical formula of the oxide of copper.

Item	Mass (g)
crucible	27.002
crucible plus contents before heating	27.128
crucible plus contents after heating	27.114

3 marks

5. 4.6 g of anhydrous zinc sulfate with an M_r of 161.5 g mol⁻¹ was obtained by driving the water from 8.2 g of the crystalline hydrated salt with empirical formula ZnSO₄·xH₂O. Calculate the value of x.

3 marks

6. Insects of a particular species can identify their mate by using special chemicals called pheromones that transmit chemical messages. The pheromone that serves as a sex attractant for gypsy moths is called disparlure and contains the elements C, H and O. Analysis of disparlure shows that 0.282 g contains 16.00 × 10⁻³ g of O atoms and 0.228 g of C atoms. The M_r of disparlure is 282 g mol⁻¹. Determine the molecular formula of disparlure.

3 marks

7. Use examples to explain the difference between the terms 'mass number', 'relative atomic mass' and 'molar mass'.

6 marks

study on

Unit 1

Quantifying chemistry

AOS 1

Topic 3



Sit topic test