# Cambridge OL

# Chemistry

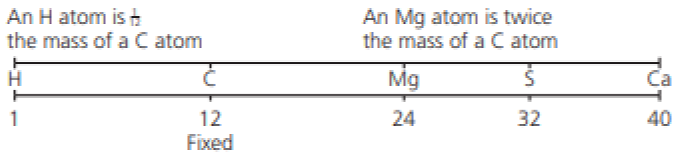
CODE: (5070)

# Chapter 03

# Stoichiometry – chemical calculations

# 4.1 Relative atomic mass

The relative atomic mass (Ar) scale is used to measure small quantities, with carbon atoms given a mass of 12.00, while other elements' atoms have a mass of 12.00.




▲ **Figure 4.1** The relative atomic masses of the elements H, C, Mg, S and Ca

**Key definition**
**Relative atomic mass, $A_r$,** is the average mass of the isotopes of an element compared to 1/12th of the mass of an atom of $^{12}C$.

## Reacting masses

Chemists often need to be able to show the relative masses of the atoms involved in a chemical process.

$$C + O_2 \rightarrow CO_2$$

The **relative molecular mass** of molecules like $O_2$ and $CO_2$ can be determined by calculating the sum of the relative atomic masses of all elements in the substance's molecular formula. For example, O2 has a relative molecular mass of 32, as each oxygen atom has a relative atomic mass of 16.

**Key definitions**
**Relative molecular mass**, $M_r$, is the sum of the relative atomic masses. **Relative formula mass**, $M_r$, is used for ionic compounds.

This number is $6.02 \times 10^{23}$ atoms, ions or molecules and is called **Avogadro's constant** after the famous Italian scientist Amedeo Avogadro (1776–1856). An amount of substance containing $6.02 \times 10^{23}$ particles is called a **mole** (often abbreviated to mol).

**Key definition**
The **mole**, symbol mol, is the unit of amount of substance. One mole contains $6.02 \times 10^{23}$ particles, e.g. atoms, ions, molecules. This number is called the Avogadro constant.

# 4.2 Calculating moles

The relative atomic mass scale compares the masses of other atoms with carbon atoms, ensuring that an element's relative mass contains $6.02 \times 10^{23}$ or 1 mole of its atoms.

## Moles and elements

The relative atomic mass (Ar) of elements like iron and aluminium can be calculated using the relationship mass (in grams) = number of moles × molar mass of the element. The molar mass of an element or compound is the mass of 1 mole, with units of g/mol. For example, a mole of magnesium contains $6.02 \times 10^{23}$ atoms.


**a** A mole of magnesium **b** A mole of carbon

▲ **Figure 4.2**

# 4.3 Moles and compounds

The molar mass of water molecules, consisting of 2 hydrogen and 1 oxygen atom, is 18 g. This compound's mass, denoted as its relative molecular mass (Mr), is expressed in units of g/mol. Understanding moles and compounds is crucial for comprehending their properties.

▲ **Figure 4.3** One mole of water ($H_2O$) (left) and 1 mole of ethanol ($C_2H_5OH$) (right) in separate measuring cylinders

If we know the mass of the compound, then we can calculate the number of moles of the compound using the relationship:

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

## Moles and gases

Gases exist as substances, and moles can be determined by measuring volume. One mole of gas occupies 24 dm3 at room temperature and pressure, known as the **molar gas volume.**

Therefore, it is relatively easy to convert volumes of gases into moles and moles of gases into volumes using the following relationship:

$$\text{number of moles of a gas} = \frac{\text{volume of the gas (in dm}^3\text{ at r.t.p.)}}{24\ \text{dm}^3}$$

or

$$\text{volume of a gas (in dm}^3\text{ at r.t.p.)} = \text{number of moles of gas} \times 24\ \text{dm}^3$$

Amedeo Avogadro proposed **Avogadro's Law**, which states that equal volumes of gases at the same temperature and pressure must contain the same number of molecules.

## Moles and solutions

Chemists frequently require the **concentration** of a solution, which can be measured in grams per cubic decimetre (g/dm3) or moles per cubic decimetre (mol/dm3).

**Key definition**

**Concentration** can be measured in $g/dm^3$ or $mol/dm^3$.

A simple method of calculating the concentration uses the relationship:

$$\text{concentration (in mol/dm}^3\text{)} = \frac{\text{number of moles}}{\text{volume (in dm}^3\text{)}}$$

Sometimes chemists need to know the mass of a substance that has to be dissolved to prepare a known volume of solution at a given concentration. A simple method of calculating the number of moles, and so the mass of substance needed, is by using the relationship:

$$\text{number of moles} = \text{concentration (in mol/dm}^3\text{)} \times \text{volume in solution (in dm}^3\text{)}$$

# 4.4 Calculating formulae

The formula for a compound indicates the number of moles of each element, like lead(II) bromide's $PbBr_2$. If the compound's formula is unknown, the masses of the elements can be determined experimentally, allowing for the calculation.

# 4.5 Moles and chemical equations

When we write a balanced chemical equation, we are indicating the numbers of moles of reactants and products involved in the chemical reaction. Consider the reaction between magnesium and oxygen.

magnesium + oxygen → magnesium oxide

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

Using the ideas of moles and masses we can use this information to calculate the quantities of the different chemicals involved.

| $2Mg(s)$ | + | $O_2(g)$ | → | $2MgO(s)$ |
|---|---|---|---|---|
| 2 moles | | 1 mole | | 2 moles |
| $2 \times 24$ | | $1 \times (16 \times 2)$ | | $2 \times (24 + 16)$ |
| = 48 g | | = 32 g | | = 80 g |

The **Law of Conservation** of **Mass**, first formulated by Antoine Lavoisier in 1774, states that the total mass of reactants equals the total mass of product in any chemical reaction.

## Gases

Many chemical processes involve gases. The volume of a gas is measured more easily than its mass. This example shows how chemists work out the volumes of gaseous reactants and products needed using Avogadro's Law and the idea of moles.

## Solutions

Chemists use solutions to perform reactions, calculating the volume of 1 mol/dm$^3$ $H_2SO_4$ solution needed to completely react with 6 g of magnesium (Ar: Mg = 24).

number of moles of magnesium

$$= \frac{\text{mass of magnesium}}{\text{molar mass of magnesium}} = \frac{6}{24}$$

$$= 0.25$$

| $Mg(s)$ | + | $H_2SO_4(aq)$ | → | $MgSO_4(aq)$ | + | $H_2(g)$ |
|---|---|---|---|---|---|---|
| 1 mole | | 1 mole | | 1 mole | | 1 mole |
| 0.25 mol | | 0.25 mol | | 0.25 mol | | 0.25 mol |

So 0.25 mol of $H_2SO_4(aq)$ is required. Using:

volume of $H_2SO_4(aq)$ (dm$^3$)

$$= \frac{\text{moles of } H_2SO_4}{\text{concentration of } H_2SO_4 \text{ (mol/dm}^3\text{)}} = \frac{0.25}{1}$$

$$= 0.25 \text{ dm}^3 \text{ or } 250 \text{ cm}^3$$

### Percentage yield

Chemical reactions are not 100% efficient, as demonstrated by the carbon-oxygen reaction, where the predicted product amount is rarely achieved.

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

The equation for the reaction states that 1 mole of carbon reacts with oxygen to give 1 mole of carbon dioxide gas

The theoretical yield of carbon dioxide from a 100% efficient reaction is 44g, but the actual yield may be less due to other reactions involving carbon and oxygen, resulting in carbon monoxide. The percentage yield is determined by the actual carbon dioxide produced.

$$2C(s) + O_2(g) \rightarrow 2CO(g)$$

### Percent composition

**Percent composition** is a measure of the percentage of a compound's mass by mass, calculated by dividing the mass of a specific element by its molar mass.

The molar mass of MgO is 24 + 16 = 40 g. Of this 24 g is magnesium and 16 g is oxygen.

$$\%\,Mg = \frac{24}{40} \times 100 = 60\%$$

$$\%\,O = \frac{16}{40} \times 100 = 40\%$$

### Percentage purity

Chapter 2 emphasizes the importance of substance **purity** in manufacturing medicines and food chemicals, as impurities can harm users or food consumers.

$$\textbf{percentage purity} = \frac{\text{mass of the pure product}}{\text{mass of the impure product obtained}} \times 100\%$$

**A limiting reactant** is a chemical reaction where one reactant is completely used up and does not remain at the end, indicating that the reaction is complete.

# Revision questions

1) Dilute hydrochloric acid reacts with sodium carbonate solution.

$$2HCl(aq) + Na_2CO_3(aq) \rightarrow 2NaCl(aq) + H_2O(l) + CO_2(g)$$

(a) Explain why effervescence is seen during the reaction.

(b) Dilute hydrochloric acid was titrated with sodium carbonate solution.

- 10.0 cm3 of 0.100 mol / dm3 hydrochloric acid were placed in a conical flask.
- A few drops of methyl orange indicator were added to the dilute hydrochloric acid.
- The mixture was titrated with sodium carbonate solution.
- 16.2 cm3 of sodium carbonate solution were required to react completely with the acid.

(i) What colour would the methyl orange indicator be in the hydrochloric acid?

(ii) Calculate how many moles of hydrochloric acid were used.

(iii) Use your answer to (b)(ii) and the equation for the reaction to calculate the number of moles of sodium carbonate that reacted

(iv) Use your answer to (b)(iii) to calculate the concentration of the sodium carbonate solution in mol / $dm^3$.

(c) In another experiment, 0.020 mol of sodium carbonate were reacted with excess hydrochloric acid.

Calculate the maximum volume (at r.t.p.) of carbon dioxide gas that could be made in this reaction.

2) (a) Alkanes and alkenes are examples of hydrocarbons.

(i) What is meant by the term hydrocarbon?
(ii) Give the general formula of straight-chain alkanes,

(b) A compound X contains carbon, hydrogen and oxygen only. X contains 54.54% of carbon by mass, 9.09% of hydrogen by mass and 36.37% of oxygen by mass.

(i) Calculate the empirical formula of compound X.

(ii) Compound X has a relative molecular mass of 88.

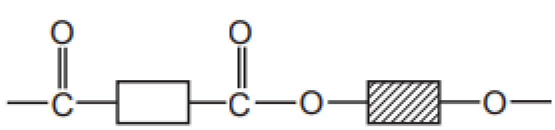
Deduce the molecular formula of compound X.

**(c)** An ester has the molecular formula $C_3H_6O_2$.

Name and give the structural formulae of **two** esters with the molecular formula $C_3H_6O_2$.

| | | |
|---|---|---|
| name of ester | | |
| structural formula | | |

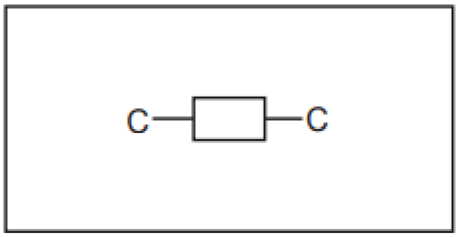
(d) Name the ester produced from the reaction of propanoic acid and methanol.
(e) A polyester is represented by the structure shown.

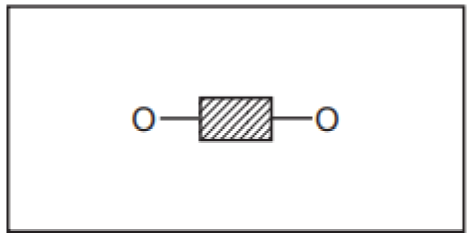


(i) What type of polymerisation is used for the production of polyesters?

(ii) Which simple molecule is removed when the polyester is formed?

(iii) Complete the diagrams below to show the structures of the monomers used to produce the polyester. Show all atoms and bonds.




3) (a) A compound, X, contains 55.85% carbon, 6.97% hydrogen and 37.18% oxygen.
(i) How does this prove that compound X contains only carbon, hydrogen and oxygen?
(ii) Use the above percentages to calculate the empirical formula of compound X.
(iii) The $M_r$ of X is 86. What is its molecular formula

(b) (i) Bromine water changes from brown to colourless when added to X. What does this tell you about the structure of X?
(ii) Magnesium powder reacts with an aqueous solution of X. Hydrogen is evolved. What does this tell you about the structure of X?
(iii) X contains two different functional groups. Draw a structural formula of X.

4) The Atacama Desert in Chile has deposits of the salt sodium nitrate. Very large amounts of this salt were exported to Europe for use as a fertiliser. After the introduction of the Haber process in 1913, this trade rapidly diminished.

(a) Explain why the introduction of the Haber process reduced the demand for sodium nitrate.

(ii) Suggest why surface deposits of sodium nitrate only occur in areas with very low rainfall such as desert areas.

(iii) The desert has smaller surface deposits of potassium nitrate. Suggest why potassium nitrate is a better fertiliser than the sodium salt.

(b) All nitrates decompose when heated. The extent to which a nitrate decomposes is determined by the metal in the salt.

(i) Sodium nitrate decomposes to form sodium nitrite, $NaNO_2$. Write the equation for decomposition of sodium nitrate.

(ii) Sodium nitrite is a reducing agent. What would be observed if an excess of sodium nitrite solution was added to a solution of acidified potassium manganate (VII)?

(iii) Copper (II) nitrate decomposes to form copper(II) oxide, nitrogen dioxide and oxygen. What is the relationship between the extent of decomposition and the reactivity of the metal in the nitrate?

5) The law of constant composition states that all pure samples of a compound contain the same elements in the same proportion by weight.

A typical experiment to test this law is to prepare the same compound by different methods and then show that the samples have the same composition.

Methods of making copper (II) oxide include:

• heating copper carbonate,
• heating copper hydroxide,
• heating copper nitrate,
• heating copper foil in air.

(a) Complete the following equations.

(i) $CuCO_3 \rightarrow$ ............ + ............

(ii) $Cu(OH)_2 \rightarrow$ ............ + ............

(iii) $2Cu(NO_3)_2 \rightarrow$ ............ + $4NO_2$ + ............

(b) Copper oxide can be reduced to copper by heating in hydrogen.
(i) What colour change would you observe during the reduction?
(ii) Explain why the copper must be allowed to cool in hydrogen before it is exposed to air.
(iii) Name another gas which can reduce copper(II) oxide to copper.
(iv) Name a solid which can reduce copper(II) oxide to copper.

6) (a) A compound X contains 82.76% of carbon by mass and 17.24% of hydrogen by mass.

(i) Calculate the empirical formula of compound X.

(ii) Compound X has a relative molecular mass of 58. Deduce the molecular formula of compound X.

(b) Alkenes are unsaturated hydrocarbons.

(i) State the general formula of alkenes.
(ii) State the empirical formula of alkenes.

(c) What is meant by the term unsaturated hydrocarbon?

(d) Describe a test that would distinguish between saturated and unsaturated hydrocarbons.

7) Quantities of chemicals, expressed in moles, can be used to find the formula of a compound, to establish an equation and to determine reacting masses.

(a) A compound contains 72% magnesium and 28% nitrogen. What is its empirical formula?

(b) A compound contains only aluminium and carbon. 0.03moles of this compound reacted with excess water to form 0.12moles of $Al(OH)_3$ and 0.09moles of $CH_4$.

Write a balanced equation for this reaction.

(c) 0.07moles of silicon reacts with 25g of bromine.

$$Si + 2Br_2 \longrightarrow SiBr_4$$

(i) Which one is the limiting reagent? Explain your choice.

(ii) How many moles of $SiBr_4$ are formed?

8) Soluble salts can be made using a base and an acid.

(a) Complete this method of preparing dry crystals of the soluble salt cobalt (II) chloride-6-water from the insoluble base cobalt (II) carbonate. step 1 Add an excess of cobalt (II) carbonate to hot dilute hydrochloric acid.

step 2 ……………………………..
step 3 ……………………………
step 4……………………………..

(b) (i) 5.95g of cobalt (II) carbonate were added to 40cm3 of hydrochloric acid, concentration 2.0mol/dm3 .

Calculate the maximum yield of cobalt (II) chloride-6-water and show that the cobalt(II) carbonate was in excess.

$$CoCO_3 + 2HCl \rightarrow CoCl_2 + CO_2 + H_2O$$

$$CoCl_2 + 6H_2O \rightarrow CoCl_2.6H_2O$$

**maximum yield:**

number of moles of $HCl$ used = ....................................................................

number of moles of $CoCl_2$ formed = ...............................................................

number of moles of $CoCl_2.6H_2O$ formed = .....................................................

mass of one mole of $CoCl_2.6H_2O$ = 238 g

maximum yield of $CoCl_2.6H_2O$ = ....................................................................g

**to show that cobalt(II) carbonate is in excess:**

number of moles of $HCl$ used = ..................................... (use your value from above)

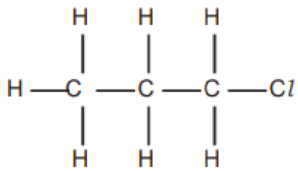
mass of one mole of $CoCO_3$ = 119 g

number of moles of $CoCO_3$ in 5.95 g of cobalt(II) carbonate = .............................. [5]

**(ii)** Explain how these calculations show that cobalt(II) carbonate is in excess.

...................................................................................................................... [1]

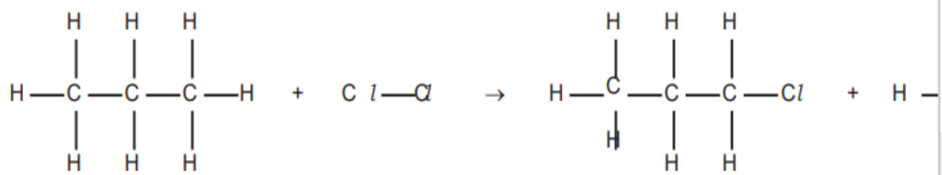
9) a) Propane reacts with chlorine to form a mixture of chloropropanes. This is a photochemical reaction.
(i) What is meant by the phrase photochemical reaction?
(ii) The products of this reaction include two isomers, one of which has the following structural formula.



Draw the structural formula of the other isomer.
(iii) Explain why these two different compounds are isomers.

(b) Bond breaking is an endothermic change and bond forming is an exothermic change.

Bond energy is the amount of energy in kJ / mol needed to break one mole of the specified bond.



Use the following bond energies to determine whether this reaction is exothermic or endothermic. You must show your reasoning.

| bond | bond energies in kJ/mol |
|---|---|
| C–C*l* | 338 |
| C–H | 412 |
| C*l*–C*l* | 242 |
| H–C*l* | 431 |
| C–C | 348 |

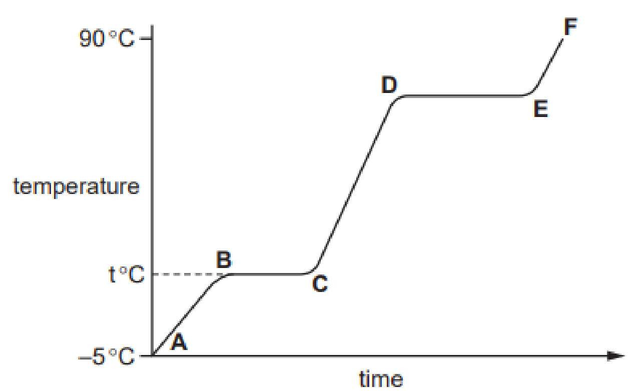
10) Compound X is a colourless liquid at room temperature.

(a) A sample of pure X was slowly heated from –5.0 °C, which is below its melting point, to 90 °C, which is above its boiling point. Its temperature is measured every minute, and the results are represented on the graph.

(i) Complete the equation for the equilibrium present in the region BC.

X(s) ⇌ ......................

(ii) What is the significance of temperature t°C?
(iii) What is the physical state of compound X in the region EF?
(iv) What would be the difference in the region BC if an impure sample of X had been used?




(b) Compound X is a hydrocarbon. It contains 85.7% of carbon. The mass of one mole of X is 84g.

(i) What is the percentage of hydrogen in the compound?

(ii) Calculate the empirical formula of X. Show your working.
(iii) What is the molecular formula of compound X?