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Stoichiometric relationships 1

Chapter outline

- Describe the three states of matter.
- Recall that atoms of different elements combine in fixed ratios to form compounds which have different properties from their component elements.
- Recall the difference between homogeneous and heterogeneous mixtures.
- Be able to construct and balance chemical equations.
- Recall the definitions for relative atomic mass, relative molecular/formula mass, empirical formula and molecular formula.
- Know how to solve calculations involving the composition of a compound, relative molecular/formula mass, empirical formula and molecular formula.
- Know that the mole is a fixed number of particles and refers to the amount of substance.
- Solve calculations involving the relationships between the number of particles, the amount of substance, the mass of a substance, the volume and concentration of a solution and the volume of a gas.
- Solve calculations using the ideal gas equation.
- Solve calculations involving theoretical and percentage yield including those involving a limiting reactant.

KEY TERMS AND FORMULAS

Avogadro's law: Equal volumes of ideal gases measured at the same temperature and pressure contain the same number of molecules.

Chemical properties: How a substance behaves in chemical reactions.

Empirical formula: The simplest whole-number ratio of the elements present in a compound.

Heterogeneous mixture: A mixture that does not have a uniform composition and consists of separate phases; it can be separated by mechanical means.

Homogeneous mixture: A mixture that has a uniform composition throughout the mixture and consists of only one phase.

Ideal gas: A theoretical model that approximates the behaviour of real gases.

Limiting reactant: The reactant that is used up first in a chemical reaction.

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Mole: The amount of substance that contains the same number of particles (atoms, ions, molecules, and so on) as there are carbon atoms in 12 g of carbon-12 (6.02×10^{23}).

Molecular formula: The total number of atoms of each element present in a molecule of the compound; the molecular formula is a multiple of the empirical formula.

Physical properties: Properties such as melting point, solubility and electrical conductivity, relating to the physical state of a substance and the physical changes it can undergo.

Titration: An analytical technique used where one solution is reacted with the exact stoichiometric amount of another solution.

Water of crystallisation: Water that is present in definite proportions in the crystals of hydrated salts (e.g. $CuSO_4{\cdot}5H_2O$).

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

where n = amount of substance in mol, m = mass in g and M_r = molar mass in $g\,mol^{-1}$

Number of particles = $n \times L$
where n = amount of substance in mol, L = Avogadro's constant (6.02×10^{23}) in mol^{-1} – this value is given in the data book

$$\%\ \text{yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

$$n = \frac{\text{volume at a given temperature and pressure}}{\text{molar gas volume at the same temperature and pressure}}$$

where n = amount of substance in mol

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

P_1 = pressure under conditions 1 / any unit of pressure
V_1 = volume under conditions 1 / any unit of volume
T_1 = temperature under conditions 1 / K
P_2 = pressure under conditions 2 / any unit of pressure (same unit as for P_1)
V_2 = volume under conditions 2 / any unit of volume (same unit as for V_1)
T_2 = temperature under conditions 2 / K

$$c = \frac{n}{V}$$

c = concentration/$mol\,dm^{-3}$
n = amount of substance/mol
V = volume/dm^3

$$\textbf{concentration in ppm} = \frac{\text{mass of solute} \times 10^6}{\text{mass of solution}}$$

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Exercise 1.1 – The particulate nature of matter

1 a Copy and complete Table 1.1, which describes the arrangement and movement of particles in solids, liquids and gases.

	Solids	Liquids	Gases
Diagram showing the arrangement of the particles			
Relative distance of the particles from one another			
Relative energy of the particles			
Relative speed of the particles			
Movement of the particles			
Relative force of attraction between the particles			

Table 1.1 Arrangement and movement of particles.

b Identify which of the descriptions of particles in Table 1.1 give rise to the fixed shape of solids and the lack of a fixed shape in liquids and gases.

c Which of the descriptions explain why, at a given temperature, the volume of a gas is not fixed but the volume of a liquid is?

2 Figure 1.1 shows the cooling curve for a substance.

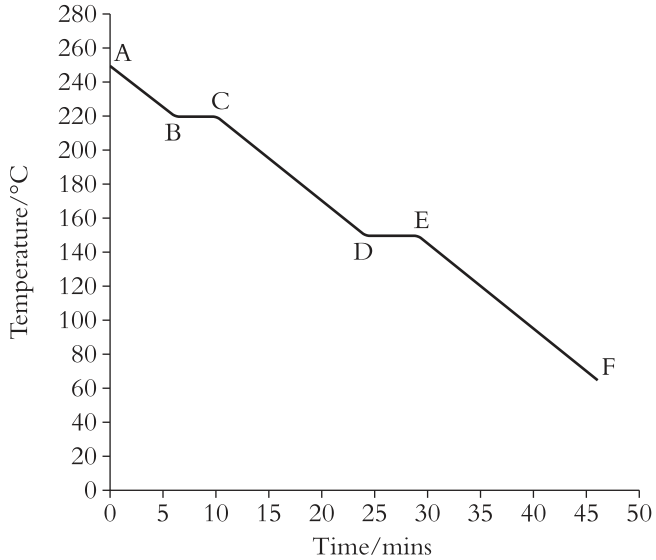

Figure 1.1 Cooling curve.

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Remember this is a cooling curve so the change goes from gas to liquid to solid. Melting points and boiling points are temperatures and should be read off the *y*-axis.

Label the diagram to show the following:

a the region where the substance is a solid
b the region where the substance is a liquid
c the region where the substance is a gas
d the region where the substance is freezing
e the region where the substance is condensing
f the melting point of the substance
g the boiling point of the substance.
h Explain, in terms of the movement and arrangement of the particles, why the temperature of the substance remains the same during a change of state.

3 Carbon dioxide and iodine are two examples of substances that undergo sublimation.
a State what is meant by the term 'sublimation'.
b State the term used to describe the reverse of sublimation.

Exercise 1.2 – Chemical change

1 a Identify whether the following substances are elements, mixtures or compounds:
- air
- water
- sodium chloride solution
- sodium chloride crystals
- iron
- chlorine gas
- carbon dioxide gas

b Describe what is meant by the terms 'homogeneous mixture' and 'heterogeneous mixture'.
c Explain why a mixture of the solids sodium chloride and sand is not a homogeneous mixture.

2 A balanced equation shows the ratio of the number of each type of particle in a chemical reaction. This is normally expressed as the simplest whole-number ratio and fractions are not normally used. (The major exception to this is in the Energetics topic – see Chapter 5).
Given the equation below for the combustion of butanol, answer the following questions

$$C_4H_9OH(l) + 6O_2(g) \rightarrow 4CO_2(g) + 5H_2O(l)$$

a Determine how many molecules of butanol would be required to produce 40 molecules of carbon dioxide.
b What number of oxygen molecules would be needed to react with 100 butanol molecules?
c State how the total number of molecules changes during this reaction.
d State how the total mass of all the atoms changes during this reaction.

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3 **a** Balance the following equations. Remember that only 'big' numbers can be used in front of a formula. 'Small' numbers are part of the formula of a substance and a formula must not be changed as this would change the identity of the substance.

i $Na + Cl_2 \rightarrow NaCl$
ii $N_2 + O_2 \rightarrow NO_2$
iii $Li(s) + H_2O(l) \rightarrow LiOH(aq) + H_2(g)$
iv $C_2H_6 + O_2 \rightarrow CO_2 + H_2O$
v $Fe_2O_3 + CO \rightarrow Fe + CO_2$
vi $Cu + HNO_3 \rightarrow Cu(NO_3)_2 + NO_2 + H_2O$

b Equation **iii** includes state symbols; what do the symbols (s), (l), (g) and (aq) indicate?

If an element occurs in several different substances (e.g. oxygen in questions **iii** and **iv**) then it is often best to tackle this element last.

Exercise 1.3 – The mole concept

Although balanced equations tell us the number of units of each species reacting in an equation, they do not tell us the masses that react because not all atoms have the same mass. The mass of an atom is so small that relative atomic masses are usually used.

1 **a** In the terms 'relative atomic mass', 'relative formula mass' and 'relative molecular mass', to what does the word 'relative' apply?

b Define the terms 'relative atomic mass', 'relative formula mass' and 'relative molecular mass'.

c Calculate the relative formula mass or relative molecular mass of the following substances:

i CO_2
ii H_2O_2
iii $NaNO_3$
iv $(NH_4)_2SO_4$
v C_2H_5OH
vi $(CH_3CH_2COO)_2Mg$
vi $H_3C_6H_5O_7$

Always use the values from the data book and be consistent with decimal places. The data book gives relative atomic masses to 2 d.p. so answers should be given to 2 d.p.

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2 Using the formulas $n = \frac{m}{M_r}$ and number of particles = $n \times L$, calculate the following:
  a the number of moles of sodium atoms in 10.0 g of sodium
  b the mass of 3 mol of $Cu(NO_3)_2$
  c the relative formula mass if 0.250 mol of the substance has a mass of 54.0 g
  d the number of water molecules in 2.50 mol of water
  e the number of oxygen molecules (O_2) in 64 g of oxygen
  f the number of oxygen atoms in 75.0 g of $CaCO_3$
  g the mass of one molecule of CO_2
  h the mass of hydrogen atoms in 0.040 mol of C_2H_6

Exercise 1.4 – Empirical and molecular formulas

Definitions are a very important part of Chemistry. You will come across a lot of them but learning them is an easy way to pick up marks and, more importantly, aids understanding.

1 Define the terms 'empirical formula' and 'molecular formula'.

2 Deduce the empirical formula from the following molecular formulas:
  a C_6H_6
  b H_2O_2
  c C_4H_{10}
  d $(COOH)_2$
  e NH_4NO_3
  f $CH_3CH_2CH_2COOH$

3 The formula of a substance can be used to find the percentage composition by mass of the substance using the formula

$$\% \text{ by mass of an element} = \frac{\text{number of atoms of the element} \times \text{atomic mass of the element}}{\text{relative formula mass}}$$

Calculate the % composition by mass of all the elements in $Mg_3(PO_4)_2$.

4 The composition of a compound can be found experimentally and can be used to determine the empirical formula of a compound. This is effectively the reverse of the calculation in Question 2 but it must be remembered that it always gives the empirical formula and not the molecular formula.
In an experiment, 0.36 g magnesium was reacted in an excess of oxygen and the oxide formed was found to have a mass of 0.60 g.
  a Calculate the mass of oxygen in the compound.
  b Deduce the empirical formula of the compound.

5 1.50 g of an organic compound containing only the elements carbon, hydrogen and oxygen with a relative molecular mass of 90.04 was combusted in excess oxygen. 1.47 g of CO_2 and 0.30 g of water were formed.
  a Calculate the empirical formula of the compound.
  b Deduce its molecular formula.

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Exercise 1.5 – Calculations involving moles and mass

Calculations from equations can be broken down into three main stages:

1 Calculate the number of moles of the substance/substances for which you have some other information such as the mass.
2 Use the ratio in the balanced equation to find the number of moles of the substance that you are being asked about.
3 Convert this number of moles into the units that the question demands (e.g. mass or volume or concentration).

1 **a** Calculate the mass of carbon dioxide produced when 1.25 g of copper(II) oxide is reacted with methane, CH_4.

$4CuO + CH_4 \rightarrow 4Cu + 2H_2O + CO_2$

b What mass of aluminium oxide is required to produce 1000 kg of aluminium?

$2Al_2O_3 \rightarrow 4Al + 3O_2$

c What mass of magnesium can be burnt in 2.50 g of carbon dioxide?

$2Mg + CO_2 \rightarrow 2MgO + C$

2 Not all reactions produce the quantity of product that is theoretically possible. The amount produced is often quoted as a percentage of the theoretical maximum that could have been obtained.

a Ethanol, C_2H_5OH, can be made from ethene, C_2H_4, using the following reaction:

$C_2H_4(g) + H_2O(g) \rightarrow C_2H_5OH(g)$

It was found that 10.0 g of ethene produced 14.6 g of ethanol. Calculate the percentage yield.

b Calcium oxide is made by the thermal decomposition of calcium carbonate which is found in limestone.

$CaCO_3 \rightarrow CaO + CO_2$

Calculate the percentage yield if 3.0 g of calcium carbonate produced 1.5 g of calcium oxide.

As with all calculations, use the same number of decimal places (for addition and subtraction) or significant figures (for multiplication and division) as are given in the original data and do not round any values until the last stage of the calculation.

3 Experiments are normally conducted using non-stoichiometric amounts. This means that the exact ratio of substances is not used; one or more of the reactants are in excess and so the amount of product that can be obtained is limited by the amount of the reactant that is not in excess. This is known as the limiting reactant.

a 2.50 g of sulfur was heated with 2.50 g of iron to form iron(II) sulfide.

$Fe(s) + S(s) \rightarrow FeS(s)$

Identify which is the limiting reactant and calculate the theoretical maximum yield of iron(II) sulfide that could be formed.

b Aluminium reacts with iodine according to the equation:

$2Al(s) + 3I_2(s) \rightarrow 2AlI_3(s)$

In an experiment 1.0 g of aluminium powder was mixed with 2.0 g of iodine. 2.0 g of aluminium iodide was produced. Identify the limiting reactant and calculate the percentage yield.

If the number of moles of each reactant is divided by the coefficient in the balanced equation, the smallest number indicates the limiting reactant.

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Exercise 1.6 – Calculations involving gas volumes

Watch out for the following when tackling questions involving gases:

- Units must be consistent. Remember 1000 cm^3 = 1 dm^3 and 1000 dm^3 = 1 m^3.
- When using $n = \dfrac{\text{volume at a given temperature and pressure}}{\text{molar gas volume at the same temperature and pressure}}$, any units can be used for the volume and for the molar gas volume as long as the same unit is used for both.
- When using $PV = nRT$, if the volume is in dm^3 then the pressure must be in kPa and vice versa.
- When using $PV = nRT$, if the pressure is in Pa then the volume must be in m^3 and vice versa.
- When using $PV = nRT$, the temperature must be in kelvin not °C.
- When using $\dfrac{P_1V_1}{T_1} = \dfrac{P_2V_2}{T_2}$, the temperature must be in kelvin not °C.
- When using $\dfrac{P_1V_1}{T_1} = \dfrac{P_2V_2}{T_2}$, any units for pressure and volume can be used as long as they are consistent on both sides of the equation.
- Think about the units!

For questions where the molar gas volume is known, for example at STP or where a value is given as in part **c**, then the expression $n = \dfrac{\textit{volume}}{\textit{molar gas volume}}$ can be used as long as the conditions are not changing. Watch out for units.

1 The molar gas volume under standard conditions of temperature and pressure is 22.7 $dm^3\,mol^{-1}$. This is given in the data book.

a Calculate the volume occupied by 0.75 mol of a gas at STP.

b How many moles of gas occupy 500 cm^3 at STP?

c At a given temperature and pressure, one mole of an ideal gas occupies 24.0 dm^3. Calculate the number of moles of gas that would occupy 150 cm^3 at the same temperature and pressure.

d What volume would 2.50 g of carbon dioxide occupy at STP?

2 The expression $\dfrac{P_1V_1}{T_1} = \dfrac{P_2V_2}{T_2}$ can be used when any of the temperature, pressure or volume of a gas are changed.

a 100 cm^3 of an ideal gas at a pressure of 100 kPa and a temperature of 330 K was cooled to a temperature of 250 K at the same pressure. What volume will the gas now occupy?

b A fixed mass of gas was sealed in a flask with a volume of 1 dm^3 at a pressure of 200 kPa and a temperature of 25 °C. The flask was heated to a temperature of 100 °C. What pressure will the gas now exert?

c At what temperature (in °C) would 200 cm^3 of an ideal gas at STP occupy 400 cm^3 at a pressure of 150 kPa?

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3 The expression $PV = nRT$ can be used to find either the volume, pressure, temperature or the number of moles of gas when the conditions do not change.

a Calculate the number of moles of gas that would occupy 400 cm^3 at a temperature of 298 K and a pressure of 1.5×10^5 Pa.

b Calculate the pressure exerted by 2.40 g of carbon dioxide with a volume of 1000 cm^3 at 25 °C.

c If 73.07 g of an ideal gas occupies 35.0 dm^3 at 50.0 °C and a pressure of 200 kPa, calculate the molar mass of the gas.

d A 0.25 g sample of a metal that contained copper was reacted with concentrated nitric acid and the nitrogen dioxide gas produced was collected in a gas syringe.

$Cu(s) + 4HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2H_2O(l) + 2NO_2(g)$

The gas occupied 87 cm^3 at a pressure of 100 kPa. The temperature of the gas was measured to be 298 K. Assuming that only the copper in the metal mixture reacted, what is the percentage of copper in the sample?

Exercise 1.7 – Calculations involving solutions

Concentration is most commonly measured in mol dm^{-3} but other units can also be used such as g dm^{-3} or ppm (part per million).

1 **a** Calculate the concentrations of the following solutions in mol dm^{-3}.

i 100 cm^3 of sodium chloride solution containing 0.75 g of NaCl

ii a solution of copper(II) sulfate ($CuSO_4$) with a concentration of 5.6 g dm^{-3}

iii a solution of volume 250 cm^3 containing 4.50 g of hydrated sodium carbonate ($Na_2CO_3{\cdot}10H_2O$)

b The analysis of a contaminated water supply found that a 150 g sample of the water contained 3.45 mg of Pb^{2+} ions. Calculate the concentration of lead ions in ppm.

c What is the concentration of aluminium ions when 3.5 g of aluminium sulfate ($Al_2(SO_4)_3$) is dissolved in water to give 200 cm^3 of solution?

A formula with '$\cdot x H_2O$' means that the solid is hydrated and contains water of crystallisation. The molar mass for the solid includes these water molecules.

2 **a** How many moles of solute are there in the following solutions?

i 25.00 cm^3 of 0.100 mol dm^{-3} NaOH(aq)

ii 50.0 cm^3 of 0.025 mol dm^{-3} $K_2Cr_2O_7$(aq)

b What mass of potassium bromide, KBr, is required to make 250 cm^3 of solution with a concentration of 0.250 mol dm^{-3}?

c What volume of calcium nitrate solution ($Ca(NO_3)_2$) with a concentration of 0.45 mol dm^{-3} will contain 0.25 mol of the solute?

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3 A small piece of calcium was added to a beaker of water and the volume of hydrogen gas produced was measured. The solution formed was then made up to a volume of 200 cm^3.

$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$

a What mass of calcium is required to produce 75 cm^3 of hydrogen gas at a temperature of 20 °C and a pressure of 101 kPa?

b What is the concentration of the calcium hydroxide solution formed?

c What volume of hydrochloric acid of concentration 0.500 $mol\,dm^{-3}$ would be required to react with the calcium hydroxide formed in this reaction?

$Ca(OH)_2(aq) + 2HCl(aq) \rightarrow CaCl_2(aq) + 2H_2O(l)$

The mass (and hence the molar mass) of oxalic acid excludes the water of crystallisation as this water is released when the solid is dissolved.

4 **a** A titration experiment was performed to find the concentration of a solution of sodium hydroxide. It was found that an average titre of 20.45 cm^3 of sulfuric acid with a concentration of 0.200 $mol\,dm^{-3}$ was required to neutralise 25.00 cm^3 of the sodium hydroxide solution.

i Give the equation for the reaction.

ii Calculate the number of moles of sulfuric acid used per titre.

iii Use the balanced equation to find the number of moles of sodium hydroxide per titre.

iv Calculate the concentration of the sodium hydroxide solution.

b 1.25 g of hydrated oxalic acid crystals, $H_2C_2O_4{\cdot}xH_2O$ were dissolved in water and made up to a volume of 250.0 cm^3. 25.00 cm^3 portions of this were then titrated against dilute sodium hydroxide of concentration 0.100 $mol\,dm^{-3}$. The average titre was found to be 19.85 cm^3.

i Calculate the number of moles of sodium hydroxide used per titre.

ii The balanced equation for the reaction is:

$H_2C_2O_4(aq) + 2NaOH(aq) \rightarrow C_2O_4Na_2(aq) + 2H_2O(l)$

Deduce the number of moles of oxalic acid in 25.00 cm^3 of solution.

iii Calculate the number of moles of oxalic acid in the original crystals.

iv Calculate the mass of oxalic acid present in the crystals.

v Calculate the mass of the water of crystallisation.

vi Calculate the number of moles of water of crystallisation.

vii Calculate the ratio of the number of moles of water to the number of moles of oxalic acid and hence determine the value for *x* in the formula of the hydrated oxalic acid.

c Succinic acid has a molar mass of 118.1 $g\,mol^{-1}$. In an experiment to find the number of hydrogen ions per succinic acid molecule, 1.01 g of succinic acid crystals were dissolved in water to make a solution of volume 250 cm^3. This was then titrated against 25.00 cm^3 portions of 0.100 $mol\,dm^{-3}$ sodium hydroxide until concordant results were obtained. It was found that an average titre of 24.40 cm^3 was required. Determine the ratio for the reaction between succinic acid and sodium hydroxide.

This question has not been structured so at first sight appears difficult. Start, as always, by calculating the number of moles of what you know. Set your work out clearly, showing the examiner each step.

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