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Steve Owen, Chris Martin
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1

STOICHIOMETRIC RELATIONSHIPS

Most chemical reactions involve two or more substances reacting with each other. Substances react with each other in certain ratios, and stoichiometry is the study of the ratios in which chemical substances combine.

This chapter covers the following topics:

- The particulate nature of matter and chemical change
- The mole concept
- Empirical and molecular formulas
- Reacting masses and volumes

1.1 The particulate nature of matter and chemical change

DEFINITIONS

COMPOUND a pure substance formed when two or more elements combine chemically.

The chemical and **physical properties** of a compound are very different to those of the elements from which it is formed. When elements combine to form compounds, they always combine in fixed ratios.

MIXTURE two or more substances mixed together.

The components of a mixture are not chemically bonded together and so retain their individual properties. The components of a mixture can be mixed together in any proportion.

HOMOGENEOUS MIXTURE a mixture that has the same (uniform) composition throughout. It consists of only one phase, such as a solution (e.g. NaCl(aq)) or a mixture of gases (e.g. the air).

These can be separated by physical means, e.g. distillation/crystallisation for NaCl(aq) and fractional distillation of liquid air.

HETEROGENEOUS MIXTURE a mixture that does not have uniform composition and consists of separate phases (regions of uniform composition). Examples include sand in water or a mixture of two or more solids (e.g. a mixture of iron and sulfur).

Heterogeneous mixtures can be separated by mechanical means, e.g. filtration or using a magnet (for a mixture of iron and sulfur).

1.2 States of matter (solids, liquids and gases)

Particle diagrams for the three states of **matter** and their interconversions are shown in Figure **1.1a**.

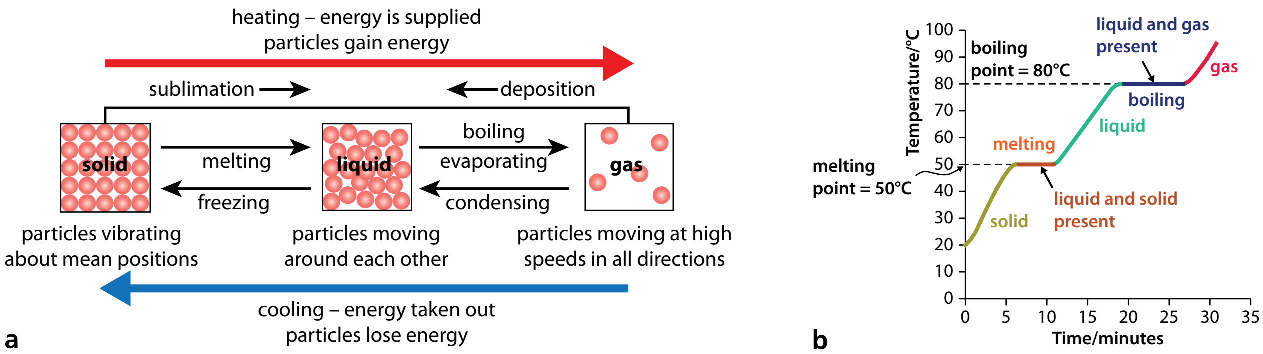

Figure 1.1

Figure **1.1b** shows how the temperature changes as a solid is heated until it becomes a gas. As a substance is heated the particles gain energy and move more quickly (vibrate more quickly for a solid). At the melting and boiling points all the energy being supplied is used to overcome forces between particles and the temperature does not rise again until the change of state is complete.

1.3 Balancing equations

When a chemical equation is required in answer to a question it must always be balanced – the number of atoms of each type (and total charge) must be the same on both sides.

How to write and balance equations:

- Never write a 'word equation' unless it is specifically requested – always write a symbol equation.
- Double check that all formulas are correct, then do not change them.
- Only large numbers (coefficients) can be inserted/changed – never change subscripts or superscripts.
- Balance compounds first and then elements last (they do not depend on anything else).
- Think about whether a full equation is required or an ionic equation (if writing an ionic equation you must also make sure that all charges balance).
- Think about the type of arrow required – a reversible arrow can sometimes be marked wrong – only use reversible arrows for questions on equilibrium and weak acids/bases unless you are absolutely sure that the reaction is reversible.
- Does the question request state symbols? If not, it is probably better to leave them out (better to leave them out than get them wrong). State symbols are essential in some equations, e.g. energetics questions (Chapter 5).

hint

Water is a liquid and not an aqueous solution – $H_2O(l)$.

State symbols:	(s) = solid	(l) = liquid	(g) = gas	(aq) = aqueous (dissolved in water)

hint

If asked to balance an equation in a multiple choice question and then work out the sum of the coefficients, do not forget the '1's: e.g. the sum of the coefficients in the equation $4Na + O_2 \rightarrow 2Na_2O$ is 7 because the coefficient of O_2 is 1.

TEST YOURSELF 1.1

State the sum of the coefficients when each of the following is balanced using the smallest possible integers.

1 $SF_4 + H_2O \rightarrow SO_2 + HF$
2 $Fe_2O_3 + CO \rightarrow Fe + CO_2$
3 $NH_3 + O_2 \rightarrow N_2 + H_2O$

1.4 Moles

DEFINITIONS

RELATIVE ATOMIC MASS (A_r) of an element is the average mass of the naturally occurring isotopes of the element relative to the mass of $\frac{1}{12}$ of an atom of carbon-12.

RELATIVE MOLECULAR MASS (M_r) of a compound is the mass of a molecule of the compound relative to the mass of $\frac{1}{12}$ of an atom of carbon-12.

The M_r is the sum of the relative atomic masses for the individual atoms making up a molecule. Therefore, the relative molecular mass of methane (CH_4) is 16.05.

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DEFINITIONS

A MOLE is the amount of substance that contains 6.02×10^{23} particles.
AVOGADRO'S CONSTANT has the value 6.02×10^{23} mol^{-1} and is sometimes given the symbol L or N_A.

For example, since the A_r of Si is 28.09 the molar mass of Si is 28.09 g mol^{-1}. This means that 28.09 g of Si contains 6.02×10^{23} Si atoms.

CO_2 – M_r 44.01 – molar mass 44.01 g mol^{-1} – 44.01 g of CO_2 contains 6.02×10^{23} **molecules**.

The number of moles present in a certain mass of substance can be worked out using the equation:

$$\text{Number of moles } (n) = \frac{\text{mass of substance}}{\text{molar mass}}$$

hint

Anything **relative** does not have units but molar mass does have units (g mol^{-1}) – the units of molar mass are not g.

TEST YOURSELF 1.2

1 Work out the number of moles in 2.00 g of methane (CH_4).
2 Work out the mass of 0.0100 mol calcium sulfate.

The mass of a molecule

The mass of 1 mole of water is 18.02 g. This contains 6.02×10^{23} molecules of water.

$$\text{mass of 1 molecule} = \frac{18.02}{6.02\times10^{23}} \text{ i.e. } 2.99\times10^{-23}\text{ g}$$

hint

Remember that the mass of a molecule is going to be a very **small** number!

TEST YOURSELF 1.3

1 Work out the mass of 1 molecule of propan-1-ol (C_3H_7OH).
2 Work out the mass of 1 molecule of CO_2.

The number of particles

If we multiply the number of moles of molecules by the number of a particular type of atom in a molecule (i.e. by the subscript of the atom), we get the number of moles of that type of atom. Thus, in 0.25 mol H_2SO_4 there are $4\times0.25 = 1.0$ mol oxygen atoms. If we now multiply the number of moles of atoms by 6.02×10^{23} we get the total number of atoms of that element present in the molecule – there are 6.02×10^{23} O atoms in 0.25 mol H_2SO_4.

Worked example 1.1

What is the total number of atoms present in 0.0100 mol of propane (C_3H_8)?
Avogadro's constant is 6.02×10^{23} mol^{-1}.

A 6.02×10^{21} **B** 5.47×10^{20} **C** 6.62×10^{22} **D** 1.02×10^{23}

Before starting the question: **read the question carefully** – do you need atoms, molecules, ions…?

It is handy to underline/highlight key words in the question.

There are 0.0100 mol propane molecules, therefore the number of **molecules** is $0.0100\times6.02\times10^{23}$. There are 11 atoms per molecule (3+8) and so the total number of atoms is going to be $0.0100\times6.02\times10^{23}\times11$. This calculation looks quite daunting without a calculator but it is made simpler by the fact that answers are given in the question – if we realise that 11×6 is 66, then answer C must be the correct one.

TEST YOURSELF 1.4

1 Work out the number of oxygen atoms in 0.200 mol HNO_3.
2 Work out the number of hydrogen atoms in 0.0461 g of ethanol (C_2H_5OH).

1.5 Empirical and molecular formulas

hint

The empirical formula must involve whole numbers, therefore $CH_{2.5}$ is not an empirical formula – double everything to get C_2H_5.

DEFINITION

EMPIRICAL FORMULA the simplest whole number ratio of the elements present in a compound.

Therefore, CH_2 is an empirical formula but C_2H_4 is not.

DEFINITION

MOLECULAR FORMULA the total number of atoms of each element present in a molecule of the compound.

The molecular formula is a whole number multiple of the empirical formula. For example, if the empirical formula of a compound is CH_2, the molecular formula is $(CH_2)_n$; i.e. C_2H_4 or C_3H_6 or C_4H_8 etc.

Worked example 1.2

hint

If the initial data is given in terms of percentage composition the calculation is done in exactly the same way – just start with the percentage composition of each element (instead of the mass as here) and divide by the A_r.

A compound has the following composition by mass: C 1.665 g; H 0.280 g; O 0.555 g.

a Calculate the empirical formula of the compound.

b If the relative molecular mass of the compound is 144.24, calculate the molecular formula.

a

	C	H	O
mass/g	1.665	0.280	0.555
divide by relative atomic mass to give number of moles	$\frac{1.665}{12.01}$	$\frac{0.280}{1.01}$	$\frac{0.555}{16.00}$
number of moles/mol	0.139	0.277	0.0347
divide by smallest to get ratio	$\frac{0.139}{0.0347}$	$\frac{0.277}{0.0347}$	$\frac{0.0347}{0.0347}$
ratio	4	8	1

Therefore, the empirical formula is C_4H_8O.

b The empirical formula mass is 72.12.

Divide the relative molecular mass by the empirical formula mass: $\frac{144.24}{72.12} = 2$

The empirical formula must therefore be multiplied by 2 to get the molecular formula – $C_8H_{16}O_2$.

hint

Take care when rounding numbers – 1.02 can usually be rounded to 1 but 1.33 cannot, e.g. if the final ratio comes out as Pb 1 : O 1.33 this cannot be rounded down to PbO but rather both numbers must be multiplied by 3 to give Pb_3O_4.

TEST YOURSELF 1.5

1. Work out which of the following are empirical formulas and which are molecular formulas:
   C_2H_2 CH_2 N_2H_3 C_3H_5 C_6H_6 NHO
2. A compound contains 93.88% P and 6.12% H. Calculate the empirical formula of the compound.
3. A compound, X, contains 24.7% K, 34.8% Mn and 40.5% O by mass. Work out the empirical formula of X.

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1.6 Using moles in calculations

There are three main steps to doing a moles calculation:

1 Work out the number of moles of anything you can.

2 Use the chemical (stoichiometric) equation to work out the number of moles of the quantity you require.

3 Convert moles to the required quantity – volume, mass etc.

There are three ways of working out the number of moles of a substance:

Masses are given

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

Volumes of gases are given

$$\text{Number of moles (mol)} = \frac{\text{volume of gas (dm}^3\text{)}}{\text{molar volume of the gas (dm}^3\ \text{mol}^{-1}\text{)}}$$

The molar volume of a gas at STP (273 K, 100 kPa) is 22.7 $dm^3\ mol^{-1}$.
STP is **standard temperature and pressure**.

Concentrations and volumes of solutions are given

$$\text{Number of moles (mol)} = \text{concentration (mol dm}^{-3}\text{)} \times \text{volume (dm}^3\text{)}$$

hint

When working out the number of moles of a gas having been given a volume you must be careful with the units – check whether the volume is given in dm^3 or cm^3. If the volume is given in cm^3 it must be divided by 1000 before working out the number of moles using the molar volume.

☆ Model answer 1.1

Consider the reaction: $CS_2(l) + 3Cl_2(g) \rightarrow CCl_4(l) + S_2Cl_2(l)$

What volume of Cl_2 (measured at STP) reacts exactly with 0.7615 g of CS_2?

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

Number of moles of CS_2:

$$\frac{0.7615}{76.15} = 0.01000 \text{ mol}$$

From the chemical equation 1 mol CS_2 reacts with 3 mol Cl_2.

Therefore, 0.01000 mol CS_2 reacts with 3×0.01000 mol Cl_2, i.e. 0.03000 mol Cl_2

Volume of $Cl_2 = 0.03000 \times 22.7 = 0.681\ dm^3$.

hint

Mass of substance must be in g but, if you are doing a multiple choice question where masses are given, for instance, in tonnes or kg, there is no need to convert so long as the answers are in the same units as the quantities in the question.

TEST YOURSELF 1.6

1 Calculate the volume of O_2 produced (measured at STP) when 5.00 g of $KClO_3$ decomposes according to the following equation:
$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$

2 Calculate the volume (in cm^3) of 0.250 mol dm^{-3} hydrochloric acid required to react exactly with 1.50 g $CaCO_3$ and the volume (in cm^3) of CO_2(g) produced at STP.
$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$

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Calculating the yield of a chemical reaction

DEFINITIONS

THEORETICAL YIELD the maximum possible amount of product formed.

EXPERIMENTAL YIELD the actual amount of product formed in the reaction.

PERCENTAGE YIELD $\frac{\text{experimental yield}}{\text{theoretical yield}} \times 100$

TEST YOURSELF 1.7

1 Calculate the percentage yield of ethyl ethanoate given that 10.0 g of ethanol reacts with excess ethanoic acid to produce 15.0 g of ethyl ethanoate.

$C_2H_5OH + CH_3COOH \rightarrow CH_3COOC_2H_5 + H_2O$

Ethanol + ethanoic acid → ethyl ethanoate + water

2 The yield of iron in the following reaction is 60.0%. Calculate how much iron is produced from 1.00 tonne of Fe_2O_3.

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

Limiting reactant

To do a moles question you only need to know the number of moles of one of the reactants. If you are given enough information to work out the number of moles of more than one reactant you must consider that one of these reactants will be the **limiting reactant**.

The amount of product formed is determined by the amount of the limiting reactant. The other reactants (not the limiting reactant) are present in excess – there is more than enough to react.

To find the limiting reactant, divide the number of moles of each reactant by its coefficient in the chemical equation and the smallest number indicates the limiting reactant.

Worked example 1.3

Consider the reaction between magnesium and hydrochloric acid:

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

Calculate the volume of hydrogen gas produced (at STP) when 0.100 g Mg reacts with 50.0 cm^3 of 0.100 mol dm^{-3} hydrochloric acid.

We have been given enough information to work out the number of moles of both magnesium and hydrochloric acid and must therefore consider that one of the reactants is limiting:

$$\text{number of moles of Mg} = \frac{0.100}{24.31} = 4.11 \times 10^{-3}\ \text{mol}$$

$$\text{number of moles of HCl} = \frac{50.0}{1000} \times 0.100 = 5.00 \times 10^{-3}\ \text{mol}$$

The coefficient of magnesium in the equation is 1 and that of HCl is 2.

$\frac{4.11 \times 10^{-3}}{1} = 4.11 \times 10^{-3}$ and $\frac{5.00 \times 10^{-3}}{2} = 2.50 \times 10^{-3}$, therefore Mg is in excess (larger number) and HCl is the limiting reactant (smaller number).

For the rest of the question we must work with the limiting reactant.

From the chemical equation, 5.00×10^{-3} mol HCl produces $\frac{5.00 \times 10^{-3}}{2}$ mol H_2, i.e. 2.50×10^{-3} mol

Volume of $H_2 = 2.50 \times 10^{-3} \times 22.7 = 0.0568\ \text{dm}^3$.

TEST YOURSELF 1.8

1 Work out the limiting reactant when 0.1 mol Sb_4O_6 reacts with 0.5 mol H_2SO_4.
$Sb_4O_6 + 6H_2SO_4 \rightarrow 2Sb_2(SO_4)_3 + 6H_2O$

2 What is the limiting reactant when 2.7 mol O_2 reacts with 2.7 mol SO_2?
$2SO_2 + O_2 \rightarrow 2SO_3$

Ideal gases

An **ideal gas** is used to model the behaviour of real gases. Two assumptions made when defining an ideal gas are that the molecules themselves have no volume and that no forces exist between them (except when they collide).

Gases deviate most from ideal behaviour at high pressure and low temperature (when a gas is most like a liquid). Under these conditions the particles will be close together – the forces between molecules and the volumes occupied by the molecules will be significant.

Nature of Science. Scientists use simplified models to describe more complex systems and make predictions – ideal gases can be described fairly simply using mathematical equations, which allows predictions to be made about the properties of real gases. Agreement/disagreement between models and observations/measurements can lead to refinement of models. Quite complex models are commonly used in describing climate change.

Calculations involving gases

Avogadro's Law: equal volumes of ideal gases measured at the same temperature and pressure contain the same number of particles. Another way of saying this is that the number of moles of an ideal gas is proportional to its volume (as long as temperature and pressure remain constant).

If you are given a volume of gas and the answer requires a volume of gas you do not have to convert to moles. So, for instance, from the equation

$H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$

1 mole of H_2 reacts with 1 mole of Cl_2 to give 2 moles of HCl, or

1 volume of H_2 reacts with 1 volume of Cl_2 to give 2 volumes of HCl, e.g.

50 cm^3 of H_2 reacts with 50 cm^3 of Cl_2 to give 100 cm^3 of HCl.

TEST YOURSELF 1.9

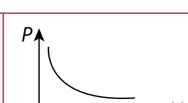
1 Work out the volume of oxygen required to react exactly with 10 cm^3 of methane in the following reaction:
$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$

2 What volume of SO_3 is produced when 20 cm^3 of SO_2 reacts with 5 cm^3 of O_2?
$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$

Nature of Science. Scientific laws develop from observations and careful measurements. A scientific law is a description of a phenomenon – a theory is a proposed explanation of that phenomenon. Avogadro's Law and the other gas laws can be explained in terms of the kinetic theory.

Relationship between pressure (P), volume (V) and temperature (T) for an ideal gas

The gas laws can be summarised as:

$P \propto \frac{1}{V}$	if the pressure of a gas is doubled at constant temperature then the volume will be halved and vice versa	

hint

A graph of PV (y-axis) against P is a horizontal straight line.

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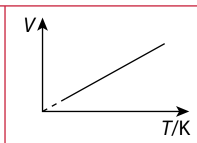
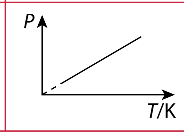
hint

Remember that V and P are only proportional to temperature in K – the relationship does **not** work for temperatures in °C. For instance, if the volume of an ideal gas at 25 °C is 500 cm^3 the volume it will occupy at 50 °C will be about 560 cm^3 – i.e. not double.

hint

Any units may be used for P and V as long as they are consistent between both sides of the equation, but temperature must be in K.

$V \propto T$	if the temperature (in K) is doubled the volume will double	
$P \propto T$	if the temperature (in K) is doubled the pressure will double	

The overall gas law equation

If you are given pressure, volume and temperature for an ideal gas you can use this equation to work out how changing some of these affects the other quantities:

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

TEST YOURSELF 1.10

Work out the final volume when 200 cm^3 of hydrogen gas is heated from 27 °C to 327 °C at constant pressure.

The ideal gas equation

$PV = nRT$

R (8.31 J K^{-1} mol^{-1}) is the **gas constant** (ideal gas constant),

n = number of moles

A consistent set of units must be used with this equation:

Pressure in Pa	or	Pressure in kPa
Volume in m^3		Volume in dm^3
Temperature in K		

Worked example 1.4

0.120 g of an ideal gas was introduced into a gas syringe. The volume occupied by the gas at a pressure of 1.02×10^5 Pa and temperature 20 °C was 49.3 cm^3. Calculate the molar mass of the gas.

Since we know the mass of the gas we can work out the molar mass if we can calculate how many moles are present. $PV = nRT$ can be used to work out the number of moles.

It is important when using $PV = nRT$, that the units are correct, so the first step should always be to write out all quantities with a consistent set of units:

$P = 1.02 \times 10^5$ Pa ✓ units fine

$V = 49.3$ cm^3 ✗ units not consistent – must be changed to m^3

1 m^3 is $100 \times 100 \times 100$ cm^3 i.e. 1×10^6 cm^3, 1 m^3 ⇔ 1×10^6 cm^3

To convert cm^3 to m^3 we go from 1×10^6 to 1, therefore we must divide by 1×10^6

$V = 49.3/(1 \times 10^6) = 4.93 \times 10^{-5}$ m^3 $n = ?$ $R = 8.31$ J $K^{-1}mol^{-1}$

$T = 20$ °C ✗ units not consistent – must be changed to K: i.e. 293 K

$PV = nRT$

$1.02 \times 10^5 \times 4.93 \times 10^{-5} = n \times 8.31 \times 293$

Re-arranging the equation gives $n = 2.07 \times 10^{-3}$ mol

This number of moles has a mass of 0.120 g

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

$$\text{i.e. molar mass} = \frac{0.120}{2.07 \times 10^{-3}}$$

i.e. 58.1 g mol^{-1}

TEST YOURSELF 1.11

1 Calculate the volume occupied by 0.100 mol of an ideal gas at 25 °C and 1.2×10^5 Pa.

2 Calculate the molar mass of an ideal gas given that 0.586 g of the gas occupies a volume of 282 cm^3 at a pressure of 1.02×10^5 Pa and a temperature of –18 °C.

1.7 Solutions

DEFINITION

A STANDARD SOLUTION a solution of known concentration.

The equation for working out concentrations in mol dm^{-3} is:

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

If the concentration is expressed in g dm^{-3}, the relationship is:

$$\text{concentration (g dm}^{-3}) = \frac{\text{mass (g)}}{\text{volume (dm}^3)}$$

Square brackets indicate the concentration of a substance e.g. [HCl] = 2.0 mol dm^{-3}

Worked example 1.5

A solution of sodium thiosulfate was made up by dissolving 2.48 g of $Na_2S_2O_3{\cdot}5H_2O$ in water and making up to a total volume of 100.0 cm^3. What is the concentration, in mol dm^{-3}, of sodium ions in the solution?

A 0.100 B 0.0100 C 0.200 D 0.314

The key thing to notice here is that the concentration of **sodium ions** is required and not that of sodium thiosulfate.

$Na_2S_2O_3{\cdot}5H_2O$ has a molar mass of 248 g mol^{-1} (do not worry about the decimal places in the A_r values when doing multiple choice questions).

$$\frac{2.48}{248} = 0.0100 \text{ mol}$$

This number is the same as Answer B – do not just pounce on the first number you get that looks like one of the answers – you must realise that the concentration is required, not the number of moles – therefore dismiss answer B.

To work out a concentration in mol dm^{-3} the number of moles must be divided by the volume in dm^3 (0.1).

0.0100/0.1 = 0.100 mol dm^{-3}

This looks like answer A but, again, you must not be tempted – this is the concentration of sodium thiosulfate and not the concentration of sodium ions. There are two sodium ions per formula unit and so this concentration must be multiplied by 2 to give answer C as the correct answer.

Answer D is incorrect – it would be obtained if the water of crystallisation is not included in the molar mass of sodium thiosulfate.

hint

If you rip the front page off Paper 1 you can keep the periodic table in front of you during the whole examination.

Titrations

Titration is a technique for finding out the volumes of solutions that react exactly with each other.

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Worked example 1.6

A student carried out an experiment to determine the concentration of a solution of sodium hydroxide using a standard solution of ethanedioic acid. She made up a standard solution of ethanedioc acid by weighing out 1.600 g of $(COOH)_2{\cdot}2H_2O$, dissolving it in distilled water then making it up to a total volume of 250.0 cm^3 with distilled water in a volumetric flask.

The student then carried out a series of titrations to determine the concentration of the unknown solution of sodium hydroxide. She measured out 25.00 cm^3 of the ethanedioic acid solution using a pipette and transferred it to a conical flask. She then added two drops of phenolphthalein indicator. A burette was filled with the sodium hydroxide solution and this was titrated against the ethanedioic acid solution. The end point was indicated by the indicator changing colour to pink. The procedure was repeated three times and the results are shown in the table.

	Trial 1	Trial 2	Trial 3
initial burette reading /cm^3 +/– 0.03	0.00	1.50	2.20
final burette reading /cm^3 +/– 0.03	29.40	30.10	30.80

The equation for the reaction is: $2NaOH + (COOH)_2 \rightarrow Na_2C_2O_4 + 2H_2O$

a Calculate the concentration of the solution of ethanedioic acid. [2]

b Determine the concentration of the sodium hydroxide solution. [4]

a In moles questions such as this the first stage is usually to determine the number of moles of anything you can. In this case we have enough information to determine the number of moles of ethanedioic acid.

M_r of $(COOH)_2{\cdot}2H_2O = 126.08$

$$\text{Number of moles} = \frac{\text{mass}}{\text{molar mass}} = \frac{1.600}{126.08} = 0.01269 \text{ mol}$$

$$\text{Concentration (mol dm}^{-3}) = \frac{\text{number of moles (mol)}}{\text{volume (dm}^3)} = \frac{0.01269}{0.250} = 0.05076 \text{ mol dm}^{-3}$$

b We need to know the volume of sodium hydroxide used in the titrations (titre) so we need to subtract the initial burette readings from the final ones (ignore uncertainties unless asked specifically about them).

	Trial 1	Trial 2	Trial 3
titre/cm^3	29.40	28.60	28.60

When experimental data is given for titrations you should always check whether the first titration was just a rough one – sometimes it is labelled as such, but not always – look to see if one titration involves a volume significantly different to the others. Here, the first titration appears to be a rough one and should be ignored in subsequent calculations.

Again, the first stage is to work out the number of moles of whatever you can – you know the volume and concentration of the ethanedioic acid solution used in each experiment so the moles of this can be worked out:

Number of moles = concentration × volume in dm^3

$$\text{Number of moles} = 0.05076 \times \frac{25}{1000} \quad \text{i.e. } 1.269 \times 10^{-3} \text{ mol}$$

The chemical equation must be used to work out the number of moles of NaOH.

$2NaOH + (COOH)_2 \rightarrow Na_2C_2O_4 + 2H_2O$

Moles of NaOH = $2 \times 1.269 \times 10^{-3} = 2.538 \times 10^{-3}$ mol

$$\text{Concentration of NaOH} = \frac{\text{number of moles (mol)}}{\text{volume (dm}^3)} = \frac{2.538 \times 10^{-3}}{0.02860} = 0.08874 \text{ mol dm}^{-3}$$