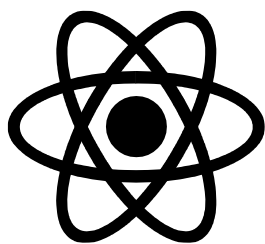


# **Bridging Course for Chemistry**



# Chemistry at Durham Johnston

## Introduction

Well done on making the choice to study Chemistry with us at Durham Johnston. This booklet will give you a brief introduction to the AS and A-level courses together with some questions to get you started on your studies.

At Durham Johnston we follow the OCR Chemistry A specification. You will have 5 lessons of chemistry per week with 2 teachers.

It is possible to study Chemistry as either an AS or full A-level: the AS level includes the teaching of units 1-4, with the full A-level adding units 5 and 6 in the 2<sup>nd</sup> year.

The modules are:

<b>Unit</b>	<b>Content</b>
1	Development of practical skills in chemistry
2	Foundations in chemistry
3	Periodic Table and energy
4	Core organic chemistry
5	Physical chemistry and transition elements
6	Organic chemistry and analysis

## Exams

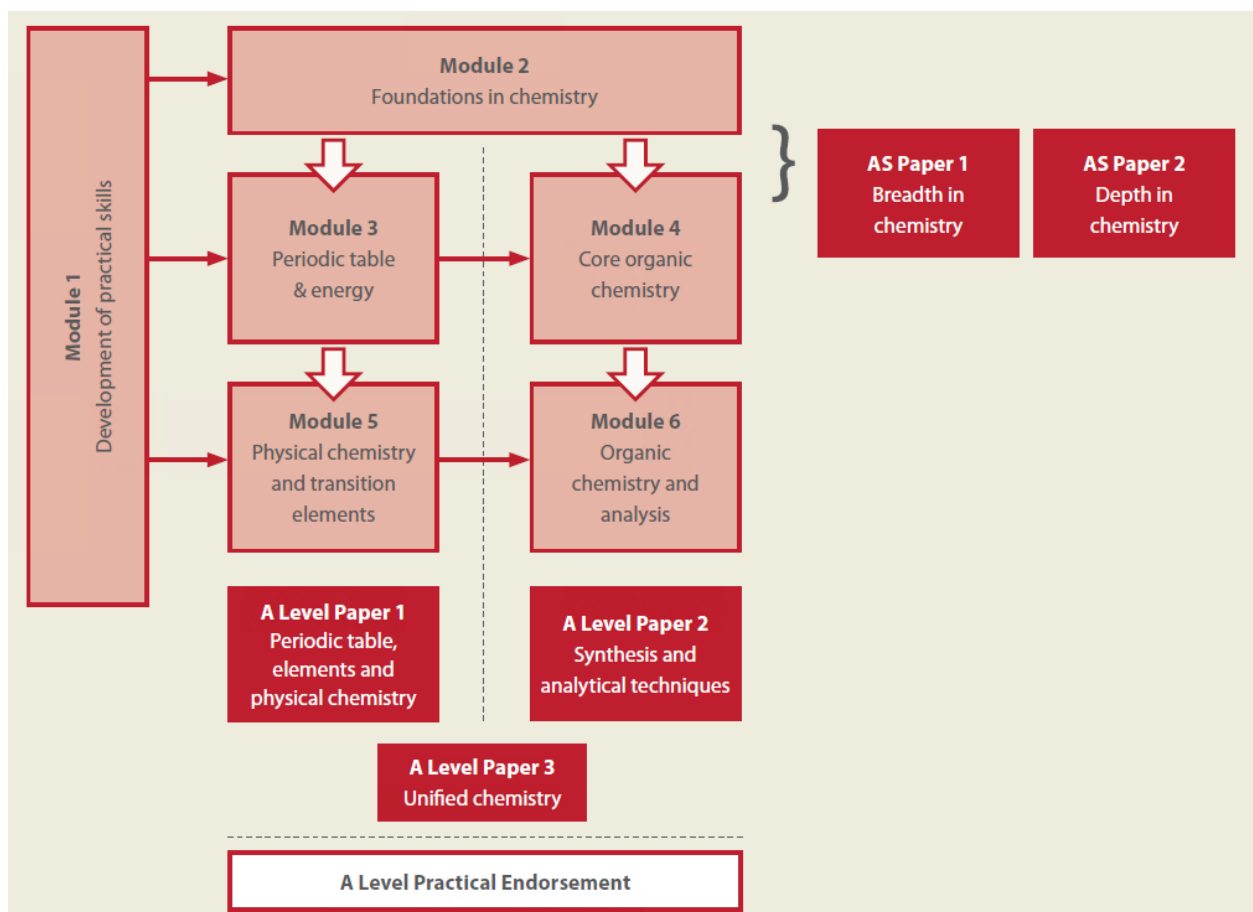
At AS level there are 2 exams with both papers assessing the content of modules 1-4.

Component	Duration	Marks	Weighting
01 (Breadth in chemistry)	1h 30	70	50%
02 (Depth in chemistry)	1h 30	70	50%

At A-level there are 3 exams, split into discrete module content as follows:

Component	Duration	Marks	Weighting
01-Periodic table, elements and physical chemistry	2h 15	100	37%
02 – Synthesis and analytical techniques	2h 15	100	37%
03 – Unified chemistry	1h 30	70	26%

The diagram below shows how the content fits together for both the AS and A-level chemistry qualifications.

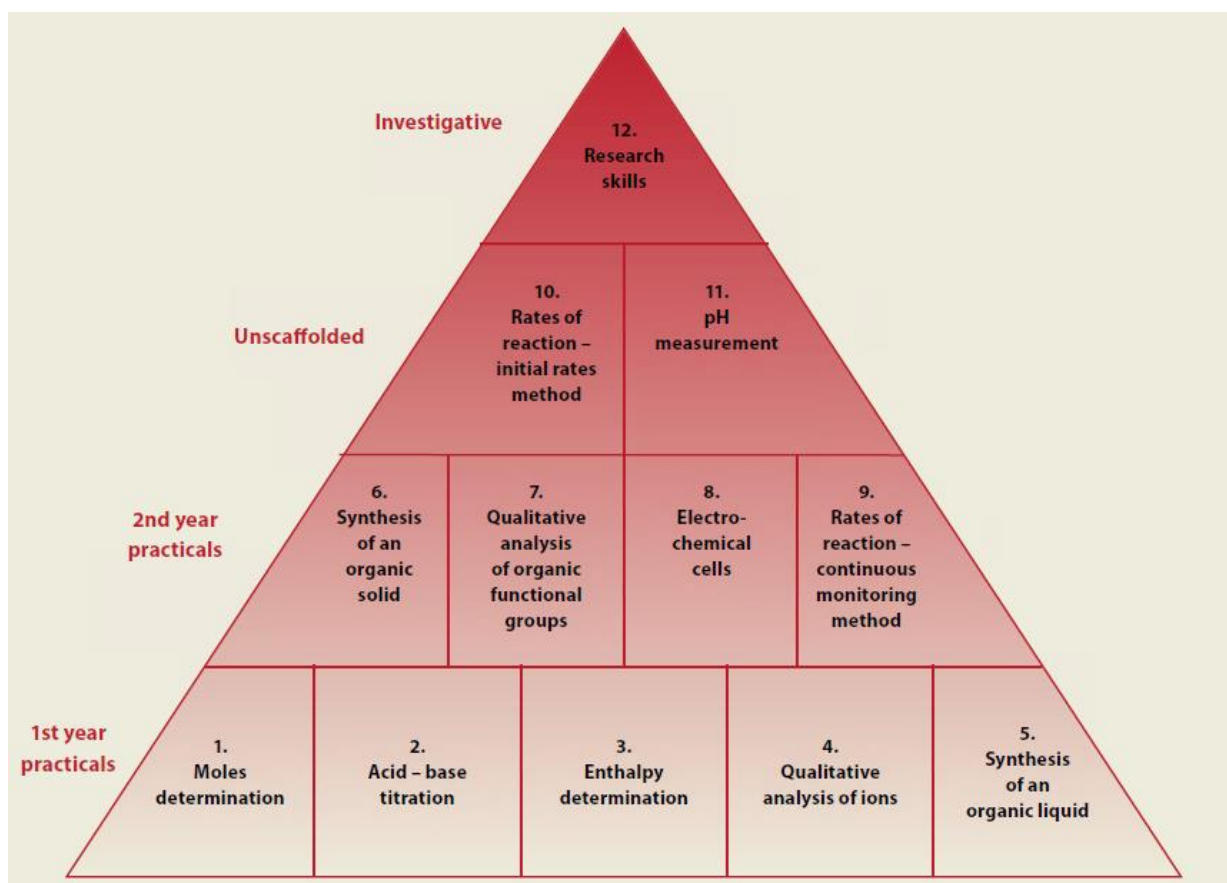


## Practical

Chemistry is a practical subject and practical skills are taught throughout the course. There are no assessed practicals as part of the qualification, the practical content is assessed within the written exam papers, and the practical

endorsement part of the course will be reported as a pass alongside the examination result at A level if students have been successful across the 12 skill areas.

The pyramid below shows how the 12 skill areas are delivered through the course.



### Want some more information?

The website for OCR can be found at: [www.ocr.org.uk](http://www.ocr.org.uk)

Links to the chemistry specifications are below, the A-level specification includes the content for the AS specification too (it's units 1-4) only.

The AS chemistry specification can be found at:

<https://www.ocr.org.uk/qualifications/as-and-a-level/chemistry-a-h032-h432-from-2015/#as-level>

The A-level chemistry specification can be found at:

<https://www.ocr.org.uk/qualifications/as-and-a-level/chemistry-a-h032-h432-from-2015/>

CGP are currently offering their 'Headstart to A-level Chemistry' as a **free** kindle download through the Amazon store. It says it's free until May 29<sup>th</sup>, but we don't know if this will be extended or not. It would be sensible to download it as soon as possible and to have a read through it. Most of the content is a revision of GCSE material so is good starting point, there are some new ideas and you might want to make some notes to get as well prepared as possible. **Important:** If you have not done separate chemistry GCSE then Sections 7 on organic chemistry (page 23 onwards), Section 12 on bond energy (page 42) are particularly important to go through as these outline content which you will not have covered before. There are little questions at the bottom of each section and the answers are at the back so you can check. Remember to write anything down that you don't understand so that you can ask us.

Finally, Chemistry is not an easy subject at A-level so you will need to work hard. However, we are all here to support you and we offer regular drop-in sessions as well as more bespoke support when it is needed.

The following pages outline some GCSE content which you should review and prepare answers for when we begin the course. You should complete your work on A4 paper.

If you have any questions about the content, then please write them down and **do remember to ask us when we return.**

Good luck – we look forward to working with you soon.

Dr Coleman (April 2020)

# Physical and Inorganic Chemistry

## Atomic Structure

### Sub-atomic particles

Scientists believe that atoms are made up of three main sub-atomic particles, the **proton**, the **neutron** and the **electron**. Their properties are summarised in Table 1.

The information in the **Periodic Table** enables us to find the number of protons, neutrons, and electrons in any atom. Since the electrons have practically no mass, the mass of the atom is simply the sum of the masses of the protons and neutrons.

Table 1: Sub-atomic particles

	<b>Proton</b>	<b>Neutron</b>	<b>Electron</b>
<b>Approximate mass in <math>\mu^*</math></b>	1	1	1/1840 (approximately 0)
<b>Electric charge</b>	+1	0	-1

\*1  $\mu$  is approximately  $1.6 \times 10^{-27}$  kg

The **atomic number** is the number of protons in an atom. It tells us what element we have, we use this because the number of protons is the same as the number of electrons in an element. To work out the number of neutrons, we use the **mass number**. The mass number is the total number of protons and neutrons in a given atom.

The number of neutrons in an atom can be found by taking the atomic number from the mass number.

For example, in sodium, Na:

23 ← relative atomic mass,  $A_r$  (= protons + neutrons)

Na

11 ← atomic number,  $Z$  (= protons = electrons)

So in an atom of sodium, Na: number of protons = 11, number of electrons = 11, number of neutrons =  $23 - 11 = 12$

### **Exercise**

- |  |
|--|
| 1. Work out the number of protons, neutrons and electrons in the following atoms: carbon, C; lithium, Li; actinium, Ac; manganese, Mn. |
|--|

## Isotopes

Some elements have atoms with different numbers of neutrons, which means that some atoms of that element will be slightly more massive than others. These slightly different atoms of the same element are called **isotopes**. The relative atomic mass,  $A_r$ , is the average mass of all the isotopes of the element.

For example, chlorine has two isotopes in the ratio of about three to one:



three of this to every one of this  
so that their average mass is 35.5, as shown below.

$$\begin{aligned}\text{mass of four atoms} &= 35 + 35 + 35 + 37 = 142 \\ \text{average mass} &= 142 / 4 = 35.5\end{aligned}$$

## Exercise

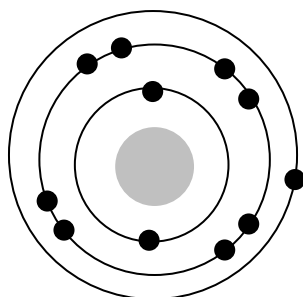
2. What would be the numbers of protons, neutrons and electrons, in the following isotopes?
- a)  $^{12}\text{C}$ ,  $^{13}\text{C}$ ,  $^{14}\text{C}$       b)  $^1\text{H}$ ,  $^2\text{H}$ ,  $^3\text{H}$ .
3. Neon, Ne, has isotopes  $^{20}\text{Ne}$  and  $^{22}\text{Ne}$ . They exist in the approximate ratio 9:1. What will be the average relative atomic mass of neon?

## Arrangement of electrons

The electron shells hold different numbers of electrons. In order of increasing distance from the nucleus, these are:

- First shell    up to 2
- Second shell    up to 8
- Third shell    up to 8 with reserve space for 10 more.

We can draw electron arrangements as in Fig. 2 or use shorthand, for example sodium: Na, 2,8,1.



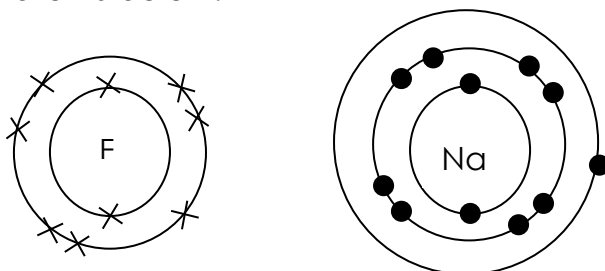
## Exercise

4. Draw the structures and write the electron arrangements for:  
a) boron, B;      b) fluorine, F;      c) lithium, Li.

### Why do atoms react?

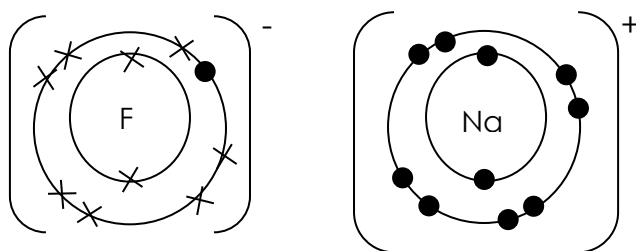
Atoms are stable when they have a full outer shell of electrons. Metals and non-metals can achieve this stability by either the transfer or sharing of electrons. The ease with which an atom can gain or lose electrons is linked to their reactivity. This is why elements in group 0 (the Noble Gases) are so unreactive.

Look at the two atoms below:



Fluorine is in group 7 of the Periodic Table, it has 7 electrons in its outer shell. To achieve stability, it needs to gain 1 electron. Sodium is in group 1 of the Periodic Table, it has 1 electron in its outer shell. To achieve stability, it needs to lose one electron.

When sodium and fluorine react together, the outer electron of sodium is transferred to the outer shell of fluorine (we show this by drawing one set of electrons as dots and the other as crosses). This results in both elements becoming stable. When they react together, both fluorine and sodium become **ions**.



Atoms that lose electrons become **positive ions**, atoms that gain electrons become **negative ions**. The number of charges tells you the number of electrons either lost or gained. The formula of sodium fluoride is NaF.

## Exercise

5. Neon atoms have 10 electrons (2,8) as do F<sup>-</sup> ions and Na<sup>+</sup> ions. Why don't fluorine and sodium become Ne when they react?
6. Give the charge of the ions you would expect to be formed by  
a. calcium    b. oxygen    c. aluminium    d. nitrogen.



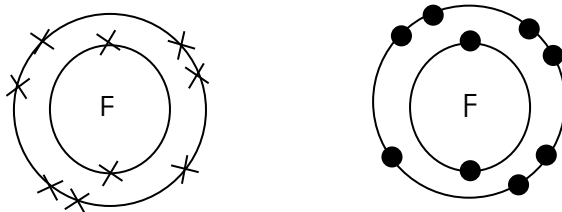
In **magnesium fluoride**, magnesium (group 2) needs to give away two electrons to get a full outer shell but fluorine needs to gain only one. So two fluorines are needed for each magnesium and the formula is  $\text{MgF}_2$  (one to two).

### Exercise

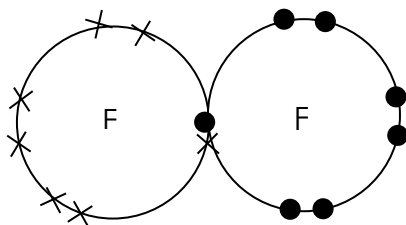
7. Draw similar diagrams to show the bonding between atoms of:
  - sodium and chlorine
  - sodium and oxygen
  - calcium and fluorine.
  
8. What is the formula and name of each of the products in question 7?

When non-metal atoms react together, they are still needing to gain electrons. These electrons are not transferred, they are instead shared between the atoms in order to achieve a full outer shell of electrons.

Look at the two atoms below:



Both of the fluorine atoms need to gain 1 electron to achieve a full outer shell of electrons. By sharing one electron each, both atoms will have a full outer shell of electrons. There is no transfer of electrons so the resulting fluorine ( $\text{F}_2$ ) molecule has no charge on it. To make it easier to show what is happening, only the outer electron shells are drawn.



### Exercise

9. Draw dot-cross diagrams to show the bonding in:
  - a. Chlorine,  $\text{Cl}_2$
  - b. Methane,  $\text{CH}_4$
  - c. Water,  $\text{H}_2\text{O}$
  - d. Oxygen,  $\text{O}_2$

## Equations

### Word and symbol equations

Word equations tell us which chemicals will react together and what products they form. Balanced symbol equations have a different function. They tell us in what proportions substances react together. State symbols are letters in brackets which can be added to say what state the reactants and products are in: (s) means solid; (l) means liquid; (g) means gas; (aq) means aqueous solution (dissolved in water).

### Exercise

10. Can you recognise a **balanced equation**?

Which of the following equations is/are balanced? If unbalanced, balance it.

- a.  $2\text{H}_2 + \text{O}_2 \longrightarrow 2\text{H}_2\text{O}$   
b.  $\text{CH}_4 + \text{O}_2 \longrightarrow \text{CO}_2 + \text{H}_2\text{O}$   
c.  $\text{Na} + \text{H}_2\text{O} \longrightarrow 2\text{NaOH} + \text{H}_2$   
d.  $\text{H}_2 + \text{Cl}_2 \longrightarrow 2\text{HCl}$

A balanced equation has the same number of each type of atom on each side of the arrow, because chemical reactions simply rearrange atoms, they do not create or destroy them.

Although equations can only be found for certain by experiment, we can often work out what the balanced equation for a given reaction must be if we know the reactants (starting materials), the products and their formulae.

### Exercise

11. Balance the two equations below. (Don't change the formulae!).

- a.  $\text{Li} + \text{F}_2 \longrightarrow \text{LiF}$   
b.  $\text{Fe} + \text{O}_2 \longrightarrow \text{Fe}_2\text{O}_3$

12. Balance the following equations:

- a.  $\text{Mg} + \text{HCl} \longrightarrow \text{MgCl}_2 + \text{H}_2$   
b.  $\text{Na} + \text{O}_2 \longrightarrow \text{Na}_2\text{O}$   
c.  $\text{Ca}(\text{OH})_2 + \text{HNO}_3 \longrightarrow \text{Ca}(\text{NO}_3)_2 + \text{H}_2\text{O}$   
d.  $\text{Ca} + \text{H}_2\text{O} \longrightarrow \text{Ca}(\text{OH})_2 + \text{H}_2$

## Moles

### Relative atomic mass, $A_r$

The relative atomic mass,  $A_r$ , shows how heavy an atom of an element is compared with an atom of hydrogen. We can find  $A_r$  from the Periodic Table.

### Exercise

13. Use the Periodic Table to look up the relative atomic masses of the following atoms.
- sodium, Na
  - oxygen, O
  - chlorine, Cl
  - lead, Pb

### Relative molecular mass, $M_r$

The relative molecular mass,  $M_r$ , is the mass of a molecule (small group of atoms covalently bonded together) compared with the mass of an *atom* of hydrogen.

The mass of a molecule is found by adding together the masses of the atoms it contains. We can either work in terms of **atomic masses** (in  $\mu$ ) or **relative atomic masses** (no units). It is called the **molecular mass** (in  $\mu$ ), or **relative molecular mass** (no units) which has the symbol  $M_r$ . Another term for the mass of one mole of a substance is the **molar mass**, this has units of  $\text{g mol}^{-1}$ .

### Example

- A. hydrogen fluoride, HF

Atomic masses of H = 1 and F = 19.  
Molecular mass =  $1 + 19 = 20$

- B. water,  $\text{H}_2\text{O}$

Atomic masses H = 1, O = 16  
Molecular mass =  $(2 \times 1) + (1 \times 16) = 18$

### Exercise

14. Work out (using relative atomic masses in the Periodic Table) the relative molecular mass of:
- ammonia,  $\text{NH}_3$
  - calcium hydroxide,  $\text{Ca}(\text{OH})_2$
  - oxygen molecule,  $\text{O}_2$
  - ethanol,  $\text{C}_2\text{H}_6\text{O}$

## The Mole

The number of atoms or other particles we get when we weigh out the relative atomic mass or the relative molecular mass in grams is enormous. It is called the Avogadro constant and is approximately equal to  $6 \times 10^{23}$ . The amount of substance that contains this number of atoms or other particles is called a **mole** (mol).

Whenever we weigh out the relative atomic mass of an element in grams we have a mole of atoms. Similarly, the relative molecular mass in grams gives a mole of a particular type of particles.

For example:

- a mole of carbon atoms, C, has a mass of 12 g
- a mole of copper atoms, Cu, has a mass of 63.5 g
- a mole of hydrogen atoms, H, has a mass of 1 g
- BUT: a mole of hydrogen *molecules*, H<sub>2</sub>, has a mass of 2 g
- a mole of calcium nitrate, Ca(NO<sub>3</sub>)<sub>2</sub>, has a mass of 164 g.

## Number of moles

We can find out how many moles of substance are present in a given mass of substance if we know its formula.

$$\text{Number of moles} = \frac{\text{mass (in g)}}{\text{molar mass}}$$

## Exercise

15. How many moles are there in:

- a. 40 g of calcium, Ca?
- b. 980 g of sulphuric acid, H<sub>2</sub>SO<sub>4</sub>?
- c. 22 g of carbon dioxide, CO<sub>2</sub>?
- d. 3.2 g of oxygen, O<sub>2</sub>?

You can also use this equation to determine the mass of a substance if given a number of moles, or the molar mass if given the mass and number of moles. Rewrite the formula above in terms of:

Mass =

Molar Mass =

## Using the idea of moles

### Empirical / Simplest Formulae

An **empirical formula** sometimes called the **simplest formula** tells us the simplest ratio of atoms in a chemical, a **molecular formula** tells us the actual composition of make-up of one molecule. Both simplest and molecular formulae must contain whole numbers.

e.g. CH<sub>4</sub> (methane) contains one atom of carbon and 4 atoms of hydrogen.

e.g. glucose has the molecular formula C<sub>6</sub>H<sub>12</sub>O<sub>6</sub>

but its simplest formula is CH<sub>2</sub>O

### Empirical Formula – Using Percentages

When a question gives the percentage, these refer to the percentage by mass.

#### Example

A white powder used in paints has the following composition:

Ba 69.58%    C 6.090%    O 16.0%

#### Answer

STEP 1 -	set out the symbols of the elements, in columns, putting the metal first.
STEP 2 -	put the % given in the Q under each symbol.
STEP 3 -	find the moles of each element (use moles = mass/A <sub>r</sub> )
STEP 4 -	find the simplest ratio of whole numbers
STEP 5 -	write the formula inserting whole numbers (omit number 1!)

	<b>Ba</b>	<b>C</b>	<b>O</b>
%	69.58	6.090	16.0
Moles	69.58 / 137.33	6.090 / 12.00	24.32 / 16
Ratio	0.5067	0.5075	1.520
whole numbers	1	1	3

Formula is **BaCO<sub>3</sub>** (notice that the number 1 is omitted).

### Empirical Formula – Using Masses

#### Example

An 18.3 g sample of hydrated compound contains 4.0 g of calcium, 7.1 g of chlorine and 7.2 g of water. Find its empirical formula.

Step 1: Set out the symbols in columns, putting the metal first. Although water is a molecule we treat it the same way as we do atoms.

Step 2: Calculate the amount of each substance (the number of moles) using  
 $\text{moles} = \text{mass} / A_r$ .

Step 3: Find the relative amount of each substance by dividing each amount by the smallest.

Step 4: Write down the formula by inserting whole numbers.

	<b>Ca</b>	<b>Cl</b>	<b>H<sub>2</sub>O</b>
Mass	4.0	7.1	7.2
Moles	<u>4.0</u>	<u>7.1</u>	<u>7.2</u>
	40	35.5	18
Ratio	0.1	0.2	0.4
÷ by smallest	<u>0.1</u> = 1	<u>0.2</u> = 2	<u>0.4</u> = 4
	0.1	0.1	0.1

Ratio is  $\text{CaCl}_2 \cdot 4\text{H}_2\text{O}$

### Exercise

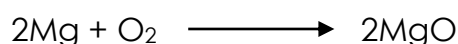
16. What is the formula of each of the following compounds?
- 196 g of sulphuric acid containing 4 g of hydrogen, 64 g of sulphur and 128 g of oxygen.
  - 12.4 g sodium oxide containing 9.2 g of sodium and 3.2 g of oxygen.
  - Find the simplest formula of the carbon oxide formed when 0.24 g carbon react with 0.32 g oxygen.
  - 0.2 g calcium reacted with fluorine to make 0.39 g calcium fluoride. Find the simplest formula of this compound.
  - A compound of carbon, hydrogen and oxygen contains 40.0% carbon, 6.6% hydrogen and 53.4% oxygen. Calculate its empirical formula.

### Reacting Masses

You will be asked to either calculate an amount of product made from an amount of reactant or work out how much reactant is needed to make a particular amount of product. This is a simple, 3-step process which requires a **BALANCED EQUATION**.

### Example

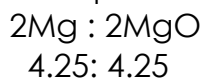
What mass of magnesium oxide can be made from 102g of magnesium?



STEP 1 – Work out the moles of magnesium used:

$$\text{Moles} = \text{mass} / \text{molar mass}$$
$$102/24 = 4.25 \text{ mol}$$

STEP 2 – Use the equation ratio to work out the moles of MgO made:

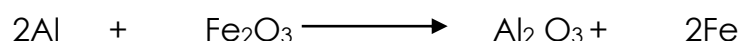


STEP 3 – Work out the mass of MgO produced:

$$\text{Mass} = \text{moles} \times \text{molar mass}$$
$$4.25 \times 40 = 170 \text{ g.}$$

### Exercise

17. What mass of iron (III) oxide must be used to produce exactly 86.0 g of iron in the following reaction?



### Concentration

Not all substances are used as solids. If a solution is involved, it is important to know its concentration. This is how much substance (the solute) is dissolved in water (or another solvent) when we make a solution.

Concentrations are measured in either mass per  $\text{dm}^3$  ( $\text{g dm}^{-3}$ ) or moles per  $\text{dm}^3$  ( $\text{mol dm}^{-3}$ )

In older books, concentration may be given the symbol M which represents molarity. **DO NOT** use this – use  $\text{mol dm}^{-3}$  instead.

**NOTE:** Often volumes in questions are given in  $\text{cm}^3$ . These MUST be converted into  $\text{dm}^3$  by dividing by 1000 ( $1 \text{ dm}^3 = 1000 \text{ cm}^3$ )

If we know the volume and concentration of a solution, we can work out the number of moles it contains:

$$\text{Number of moles} = \text{Volume (in dm}^3) \times \text{Concentration (in mol dm}^{-3})$$

### Exercise

18. How many moles of solute are there in
- 10  $\text{cm}^3$  of 1  $\text{mol dm}^{-3}$  solution?
  - 50  $\text{cm}^3$  of 2  $\text{mol dm}^{-3}$  solution?
  - 20  $\text{cm}^3$  of 0.01  $\text{mol dm}^{-3}$  solution?

## Making Solutions

As we cannot weigh quantities directly in moles, only in grams, we must convert the number of moles we require into grams when making a solution.

We can do this using the formulae:

$$\text{Number of moles} = \text{Mass} / \text{Molar Mass}$$

$$\text{Number of moles} = \text{Volume} \times \text{Concentration}$$

Remember to  
convert volume  
into  $\text{dm}^3$  if  
needed!

### Example

What mass of sodium hydroxide would be needed to prepare  $1 \text{ dm}^3$  solution of concentration  $2 \text{ mol dm}^{-3}$  sodium hydroxide?

$$\begin{aligned} \text{Number of moles} &= \frac{\text{Volume} \times \text{Concentration}}{1} \\ &= 2 \text{ mol NaOH} \end{aligned}$$

To find the mass of 2 moles of NaOH:

$$\begin{aligned} \text{Mass} &= \text{Number of moles} \times \text{Molar mass} \\ &= 2 \times 40 = 80 \text{ g NaOH needed} \end{aligned}$$

### Exercise

19. How much would you need to make the following:
- $1 \text{ dm}^3$  of  $2 \text{ mol dm}^{-3}$  sodium chloride solution ( $\text{NaCl}$ ).
  - $2 \text{ dm}^3$  of  $0.01 \text{ mol dm}^{-3}$  silver nitrate solution ( $\text{AgNO}_3$ ).
  - $250 \text{ cm}^3$  of  $0.5 \text{ mol dm}^{-3}$  sodium thiosulphate solution ( $\text{Na}_2\text{S}_2\text{O}_3$ ).
  - $100 \text{ cm}^3$  of  $1 \text{ mol dm}^{-3}$  iodine solution ( $\text{I}_2$ ).

### Titration

- During a titration a known volume (usually  $25 \text{ cm}^3$ ) of alkali (of known concentration) is put into a conical flask and is neutralised with an acid placed in a burette.
- A few drops of indicator are added.
- The acid is added to the alkali in the conical flask which is swirled until the indicator turns colour.
- The colour change indicates the "end point". This is when equivalent moles of acid and alkali have been added.

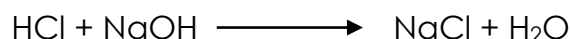


- e. This is repeated 4 times (3 accurate and 1 approximate).
- f. The average accurate volume of acid used is calculated.

Using the calculated average volume of acid, the concentration of the alkali and the volume of alkali it is then possible to calculate the exact concentration of the acid being used – as long as you have a **BALANCED EQUATION**.

### Example

100 cm<sup>3</sup> of hydrochloric acid is neutralised by 50 cm<sup>3</sup> of 2 mol dm<sup>-3</sup> sodium hydroxide. What is the concentration of the acid?

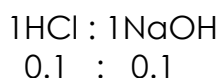


This is also a 3 step calculation, similar to the reacting masses:

STEP 1 – Work out the moles of sodium hydroxide used:

$$\text{Moles} = \frac{\text{Volume}}{1000} \times \text{Concentration} \\ (50/1000) \times 2 = 0.1 \text{ mol}$$

STEP 2 – Use the equation ratio to work out the moles of used:



STEP 3 – Work out the concentration of hydrochloric acid:

$$\text{Concentration} = \frac{\text{Moles}}{\text{Volume}} \\ 0.1 / (100/1000) = 1 \text{ mol dm}^{-3} \text{ acid}$$

### Exercise

20. Work out the concentrations of the following:

- a. 100 cm<sup>3</sup> of 1 mol dm<sup>-3</sup> sulphuric acid is just neutralised by 50 cm<sup>3</sup> potassium hydroxide. What is the concentration of the potassium hydroxide? (**Be careful**, 1 mole of sulphuric acid reacts with 2 moles of potassium hydroxide).
- b. 20 cm<sup>3</sup> of nitric acid is just neutralised by 10 cm<sup>3</sup> of 0.2 mol dm<sup>-3</sup> sodium hydroxide. What is the concentration of the acid? (1 mole of nitric acid reacts with one mole of sodium hydroxide).

21.

- a. Using your answer to 20a, work out what mass of potassium hydroxide is dissolved in 1 dm<sup>3</sup> of solution.
- b. Using your answer to 20b, work out what mass of nitric acid is dissolved in 1 dm<sup>3</sup> of solution.

Finally, some substances are in the form of gases. Gases are not measured by mass but by volume.

As the particles in a gas are spread out with lots of space between them, the volume of **ANY** gas is 24 dm<sup>3</sup> (or 24000 cm<sup>3</sup>) at room temperature (25°C) and pressure no matter what its relative formula mass.

### Example

A reaction produces 120 cm<sup>3</sup> of hydrogen. How many moles of gas is this?

$$\text{Moles of gas} = \frac{\text{volume of gas}}{\text{volume of 1 mole of gas (24dm}^3 \text{ or 24000 cm}^3\text{)}}$$

$$= \frac{120}{24\,000}$$

$$= 0.005 \text{ mol.}$$

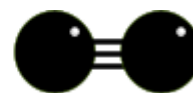
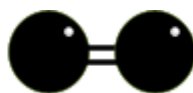
### Exercise

22. How many moles are there in the following volumes of oxygen (O<sub>2</sub>) at room temperature and pressure?
- 240 000 cm<sup>3</sup>
  - 120 cm<sup>3</sup>
  - 48 cm<sup>3</sup>
  - What difference (if any) would it make if the gas were neon?

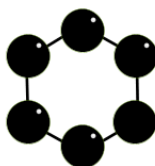
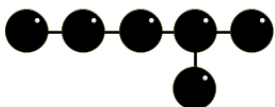
## Organic Chemistry

Organic chemistry is the study of carbon compounds. It is such a complex branch of chemistry because...

- carbon atoms can form strong covalent bonds with other carbon atoms;
- bonds between carbon atoms can be single, double or triple;



- carbon atoms can be arranged in straight lines, branched chains or in rings;



- Most organic molecules have lots of bonds between carbon and hydrogen atoms, although other atoms or groups of atoms can be put onto chains or swapped to make new types of organic molecules;

- there is often more than one way to arrange the atoms in the molecule.

### **The language of organic chemistry**

Success with organic chemistry relies on students knowing the definitions of key terms – it's a bit like learning a new language. For instance, you can't write a sentence in French if you don't know any vocabulary or grammar rules. It's the same with organic chemistry.

### **Exercise**

23. Find out and write definitions for the following key organic chemistry terms:

Homologous series, functional group, aliphatic, aromatic, alicyclic, isomer

## Types of formula

Chemists use different types of formula to help visualise molecules. These include:

1. Molecular formula
2. General formula
3. Empirical formula
4. Structural formula
5. Displayed formula
6. Skeletal formula
7. 3D formula

### Exercise

24. Find out and write definitions for the 7 types of formulae above, draw an example of a molecule to explain clearly what the definition means

## Organic families

During the A-level course, you will become familiar with lots of different types of organic molecules. You will need to be able to recognise these different molecules, name molecules that contain different functional groups and describe some of their reactions.

### Exercise

25. Copy and complete the table to identify the different types of organic molecule. You should show a displayed formula and a structural formula for each functional group.

Type of Molecule	Functional Group	Example
Alkane		
Alkene		
Alcohol		
Aldehyde		
Ketone		
Carboxylic acid		
Ester		
Halogenoalkane		
Amine		
Acyl chloride		

## Naming organic molecules

Organic molecules are always named using IUPAC convention to prevent any confusion. You will need to make sure you know the names of the first 10

### Exercise

26. Copy and complete the table to identify the different types of organic molecule. You should show a displayed formula and a structural formula for each functional group (the first one has been done for you as an example).

Number of carbon atoms	Name	Prefix
1	Meth-	methyl
2		
3		
4		
5		
6		
7		
8		
9		
10		

## Finally

Completing this work will put you in a really strong position to do well at A-level. It is important to get into good study habits as soon as possible, a lot of the content of our course builds on prior learning and knowledge, so get the basics sorted and then everything else should fall into place.

# The Periodic Table of the Elements

	1	2	3	4	5	6	7	0	
	7 <b>Li</b> lithium 3	9 <b>Be</b> beryllium 4		11 <b>B</b> boron 5	12 <b>C</b> carbon 6	14 <b>N</b> nitrogen 7	16 <b>O</b> oxygen 8	19 <b>F</b> fluorine 9	20 <b>Ne</b> neon 10
	23 <b>Na</b> sodium 11	24 <b>Mg</b> magnesium 12		27 <b>Al</b> aluminium 13	28 <b>Si</b> silicon 14	31 <b>P</b> phosphorus 15	32 <b>S</b> sulfur 16	35.5 <b>Cl</b> chlorine 17	40 <b>Ar</b> argon 18
	39 <b>K</b> potassium 19	40 <b>Ca</b> calcium 20		70 <b>Ga</b> gallium 31	73 <b>Ge</b> germanium 32	75 <b>As</b> arsenic 33	79 <b>Se</b> selenium 34	80 <b>Br</b> bromine 35	84 <b>Kr</b> krypton 36
	85 <b>Rb</b> rubidium 37	88 <b>Sr</b> strontium 38		115 <b>In</b> indium 49	119 <b>Sn</b> tin 50	122 <b>Sb</b> antimony 51	128 <b>Te</b> tellurium 52	127 <b>I</b> iodine 53	131 <b>Xe</b> xenon 54
	133 <b>Cs</b> caesium 55	137 <b>Ba</b> barium 56		204 <b>Tl</b> thallium 81	207 <b>Pb</b> lead 82	209 <b>Bi</b> bismuth 83	[209] <b>Po</b> polonium 84	[210] <b>At</b> astatine 85	[222] <b>Rn</b> radon 86
	[223] <b>Fr</b> francium 87	[226] <b>Ra</b> radium 88		201 <b>Hg</b> mercury 80	201 <b>Hg</b> mercury 80	201 <b>Hg</b> mercury 80	[272] <b>Rg</b> roentgenium 111		
				59 <b>Co</b> cobalt 27	59 <b>Co</b> cobalt 27	59 <b>Co</b> cobalt 27	[268] <b>Mt</b> meitnerium 109		
				56 <b>Fe</b> iron 26	56 <b>Fe</b> iron 26	56 <b>Fe</b> iron 26	[277] <b>Hs</b> hassium 108		
				55 <b>Mn</b> manganese 25	55 <b>Mn</b> manganese 25	55 <b>Mn</b> manganese 25	[264] <b>Bh</b> bohrium 107		
				52 <b>Cr</b> chromium 24	52 <b>Cr</b> chromium 24	52 <b>Cr</b> chromium 24	[266] <b>Sg</b> seaborgium 106		
				51 <b>V</b> vanadium 23	51 <b>V</b> vanadium 23	51 <b>V</b> vanadium 23	[262] <b>Db</b> dubnium 105		
				48 <b>Ti</b> titanium 22	48 <b>Ti</b> titanium 22	48 <b>Ti</b> titanium 22	[261] <b>Rf</b> rutherfordium 104		
				45 <b>Sc</b> scandium 21	45 <b>Sc</b> scandium 21	45 <b>Sc</b> scandium 21	[227] <b>Ac*</b> actinium 89		
				91 <b>Zr</b> zirconium 40	91 <b>Zr</b> zirconium 40	91 <b>Zr</b> zirconium 40			
				93 <b>Nb</b> niobium 41	93 <b>Nb</b> niobium 41	93 <b>Nb</b> niobium 41			
				[98] <b>Tc</b> technetium 43	[98] <b>Tc</b> technetium 43	[98] <b>Tc</b> technetium 43			
				101 <b>Ru</b> ruthenium 44	101 <b>Ru</b> ruthenium 44	101 <b>Ru</b> ruthenium 44			
				103 <b>Rh</b> rhodium 45	103 <b>Rh</b> rhodium 45	103 <b>Rh</b> rhodium 45			
				106 <b>Pd</b> palladium 46	106 <b>Pd</b> palladium 46	106 <b>Pd</b> palladium 46			
				108 <b>Ag</b> silver 47	108 <b>Ag</b> silver 47	108 <b>Ag</b> silver 47			
				112 <b>Cd</b> cadmium 48	112 <b>Cd</b> cadmium 48	112 <b>Cd</b> cadmium 48			
				65 <b>Zn</b> zinc 30	65 <b>Zn</b> zinc 30	65 <b>Zn</b> zinc 30			
				63.5 <b>Cu</b> copper 29	63.5 <b>Cu</b> copper 29	63.5 <b>Cu</b> copper 29			

1	<b>H</b>	hydrogen	1
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relative atomic mass
<b>atomic symbol</b>
name
atomic (proton) number

Elements with atomic numbers 112-116 have been reported but not fully authenticated

\* The lanthanoids (atomic numbers 58-71) and the actinoids (atomic numbers 90-103) have been omitted.  
The relative atomic masses of copper and chlorine have not been rounded to the nearest whole number.