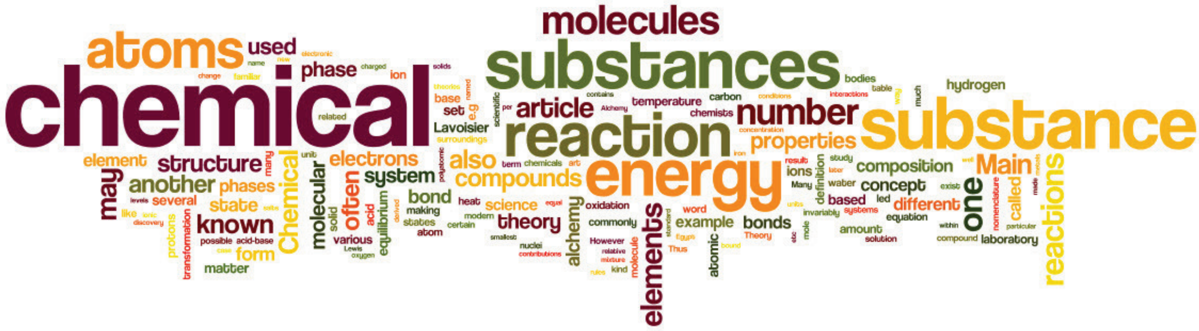
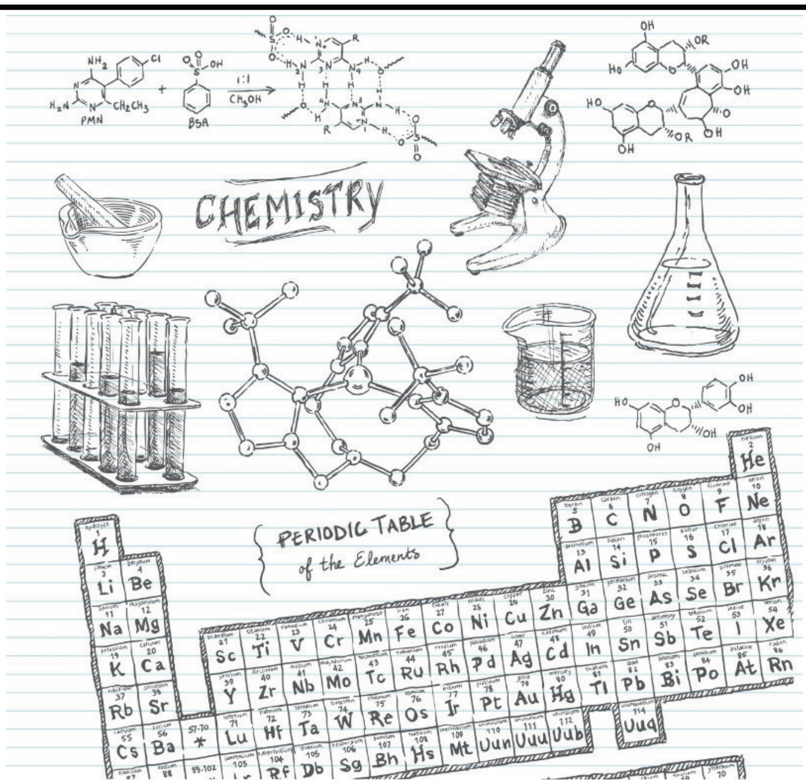


Transition Unit for Year 12 Chemistry

You should receive this booklet when you register to study A-Level chemistry at Dormers Wells High School. In preparation for the start of the course please complete the following:

1. The worksheets included in this booklet
2. Revision notes on the topic of The Structure of the Atom, following completion of the CGP Head Start to A-Level Chemistry book section 1 (included)
3. A piece of work based on your own research

Completion of the tasks will be assessed early in the course.



Preparatory tasks for A-Level chemistry

A-Level chemistry is a very demanding subject, for which reason it is very important that you make a confident start to your studies from the very beginning of year 12. The following tasks should be undertaken in order to prepare you to make such a start.

1. Recap and extension of GCSE content

Please complete the worksheets included in this booklet to help prepare you for the topics you will first encounter in year 12.

2. Making Use of the "Head Start to A-Level Chemistry" book

You are strongly encouraged to purchase a copy of the CGP 'Head Start to A-Level Chemistry' book designed to bridge the gap between GCSE and A-Level (ISBN 978 1 78294 280 1, www.cgpbooks.co.uk). If you make full use of this book then you will have the required prior knowledge from GCSE, ready for the relevant section of the A-Level chemistry course; this will make it much easier for you to make the transition from GCSE to A-Level.

Section 1 (The Structure of the Atom) is attached to this document: please complete it as part of this 'transition unit' and make some revision notes on this topic. You are strongly encouraged to complete sections 2 (Formation of Ions), 5 (Chemical Equations) and 11 (Calculations) by October half-term, and all remaining sections by the end of the Autumn term.

3. Research

Undertake some individual research based on one of the following topics and present your work (a minimum of 500 words plus relevant diagrams) as a written document; make sure you list and make reference to sources that you have consulted.

Suggested topics:

1. The development of the Periodic Table
2. The contributions of chemistry to drug development
3. The work of a particular notable chemist – eg. Harry Kroto and buckminsterfullerene: 'how does this relate to the development of nanochemistry/nanotechnology?' or 'how does it compare to other allotropes of carbon?'
4. A particular element, detailing its properties and value to mankind
5. The importance of the 'mole concept' in science
6. If you are interested in studying medicine, evaluate why chemistry is considered to be an essential subject to study at A-Level.

Summary sheet

Writing formulae

Compounds should have no overall charges, so the positive and negative charges should cancel each other out.

Apart from working out the charges on ions made up of one element, you need to know the following compound ions and their charges.

Name	Formula	Charge
hydroxide	OH^-	1–
nitrate	NO_3^-	1–
sulfate	SO_4^{2-}	2–
carbonate	CO_3^{2-}	2–
ammonium	NH_4^+	1+

Follow these steps.

Write the name of the compound	Magnesium bromide	Sodium sulfate
Work out the charge of your positive ion = group number, or 1+ for ammonium.	Mg^{2+}	Na^+
Work out the charge of your negative ion = group number – 8 *or* known charge for a compound ion.	Br^-	SO_4^{2-}
Rewrite the symbols; put a bracket around any compound ion.	Mg^{2+} Br^- Mg Br	Na^+ SO_4^{2-} Na (SO_4)
Swap the numbers of the charges and drop them to the opposite ion.	$MgBr_2$	$Na_2(SO_4)$

Writing ionic equations

- Make sure all state symbols are included.
- Identify the species that are aqueous, using the rules of solubility.

1 Look at the cation – is it Group 1 or ammonium? If so → soluble.

2 Look at the anion – is it a nitrate? If so → soluble.

- Proceed only if you have ruled out 1 and 2.

3 Is the anion a halide (chloride, bromide or iodide)?

4 If so, look at the metal – lead or silver? If so → insoluble.

5 Is the anion a sulfate?

6 If so, look at the metal – barium, calcium, lead? If so → insoluble.

7 Is the anion a hydroxide?

8 If so, look at the metal – transition metal or Group 2 (after Ca)? If so → insoluble.

- Split all the soluble salts into their aqueous ions on both sides – remember to write the numbers in front of the ions for multiples.
- Cancel out the ions that appear on both sides – again pay attention to numbers.
- Write your final equation (always keep the state symbols unless specifically told not to!).

Reacting masses

To work out masses of reactants and products from equations, follow these steps.

Steps to follow	**Example**	**Example**
	5 g of Ca reacted with excess chlorine. What mass of $CaCl_2$ is formed?	When $MgCO_3$ was heated strongly, 4 g of MgO was formed. What is the mass of $MgCO_3$ that was heated?
Write the balanced equation.	$Ca + Cl_2 \rightarrow CaCl_2$	$MgCO_3 \rightarrow MgO + CO_{2(g)}$
Write the masses given.	5 g (excess) **?**	**?** 4 g
Find the A_r or M_r.	40 111	84 40
Divide by the atomic or molecular mass (step 2 ÷ step 3).	$\frac{5}{40} : \frac{?}{111}$	$\frac{?}{84} : \frac{4}{40}$
Treat these like ratios, rearrange to find the unknown (**?**).	Mass of $CaCl_2$ = $(5 \times 111) \div 40 = 13.9$ g	Mass of $MgCO_3$ = $(4 \times 84) \div 40 = 8.4$ g

Note: if you are told something is in excess, do not use it in the calculation!

Percentage yield

The calculations above dealt with the masses you get or use if the reaction is 100% complete.

Most reactions are not 100% complete for the following reasons:

- not all the reactant reacts
- some is lost in the glassware as you transfer the reactants and the products
- some other products might be formed that you do not want.

This is a problem in industry. Less of the desired product has been made, so there is less to use or sell, and the waste has to be disposed of. Waste products can be harmful to the environment, e.g. the one above produces the greenhouse gas CO_2. Industries try to choose reactions that minimise waste and do not produce harmful products. They also try to make the rate of reaction high enough to make the reaction turnover fast so they can increase production and make money.

To work out % yield: use the balanced equation to work out how much of the given product you should get if the reaction is 100% efficient – this is the theoretical yield.

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

Worksheet A1: Atomic structure and the Periodic Table

Complete the following sentences and definitions to give a summary of this topic.

Structure of an atom

The nucleus contains …

____

The electrons are found in the …

____

To work out the number of each sub-atomic particle in an atom we use the Periodic Table (PT). The number of protons is given by …

____

In a neutral atom the number of electrons is …

____

To work out the number of neutrons we …

____

Vocabulary

State what is meant by the following terms.

9 Relative atomic mass

10 Relative molecular mass

11 Isotope

12 Relative isotopic mass

Structure of an ion

When an atom becomes an ion, only the number of ________________________ changes.

For positive ions this ___________________________________ by the number equivalent to the charge on the ion.

For negative ions this ____________________________________ by the number equivalent to the charge on the ion.

Worksheet B1: Chemical formulae

Write the formulae of the following compounds.

Copper(II) sulfate	______________________
Nitric acid	______________________
Copper(II) nitrate	______________________
Sulfuric acid	______________________
Sodium carbonate	______________________
Aluminium sulfate	______________________
Ammonium nitrate	______________________
Nitrogen dioxide	______________________
Sulfur dioxide	______________________
Ammonia	______________________
Ammonium sulfate	______________________
Potassium hydroxide	______________________
Calcium hydroxide	______________________

Worksheet B2: Cations and anions

Complete the table below to show the substance, its formula and its individual ions.

Substance	Formula	Cation (exact number)	Anion (exact number)
Sodium bromide			
	KI		
Silver nitrate			
Copper(II) sulfate			
	$NaHCO_3$		
Magnesium carbonate			
Lithium carbonate			
	$Ca(HSO_4)_2$		
Aluminium nitrate			
Calcium phosphate			
Potassium hydride			
Sodium ethanoate			
	$KMnO_4$		
Potassium dichromate(VI)			
Zinc chloride			
Strontium nitrate			
Sodium chromate(VI)			
Calcium fluoride			
Potassium sulfide			
Magnesium nitride			
Lithium hydrogensulfate			
	$(NH_4)_2SO_4$		

Worksheet B3: Writing equations

Write: **(a) the chemical equation and (b) the ionic equation for each of the following reactions.**

1 Magnesium with sulfuric acid

2 Calcium carbonate with nitric acid

3 Hydrochloric acid with sodium hydroxide

4 Aqueous barium chloride with aqueous sodium sulfate

5 Aqueous sodium hydroxide with sulfuric acid

6 Aqueous silver nitrate with aqueous magnesium chloride

7 Solid magnesium oxide with nitric acid

8 Aqueous copper(II) sulfate with aqueous sodium hydroxide

9 Aqueous lead(II) nitrate with aqueous potassium iodide

10 Aqueous iron(III) nitrate with aqueous sodium hydroxide

Section 4

The mole

When chemists measure how much of a particular chemical reacts they measure the amount in grams; or they measure the volume of a gas. However, chemists find it convenient to use a unit called a ***mole***. You need to know several definitions of a mole and be able to use them.

- The **mole** is the amount of substance, which contains the same number of particles (atoms, ions, molecules, formulae or electrons) as there are carbon atoms in 12 g of carbon -12
- This **number** is known as the *Avogadro constant, L,* and is equal to 6.02×10^{23} mol^{-1}
- The **molar mass** of a substance is the mass, in grams, of one mole
- The **molar volume** of a gas is the volume occupied by one mole at room temperature and atmospheric pressure (r.t.p). It is equal to 24 dm^3 at r.t.p
- *Avogadro's Law* states that equal volumes of all gases, under the same conditions of temperature and pressure contain the same number of moles or molecules. If the volume is 24 dm^3, at room temperature and pressure, this number, once again, is the Avogadro constant.

When you talk about moles, you must always state whether you are dealing with atoms, molecules, ions, formulae etc. To avoid any ambiguity it is best to show this as a formula.

Example calculations involving the use of moles

These calculations form the basis of many of the calculations you will meet at A level.

Example 1

Calculation of the number of moles of material in a given mass of that material

a Calculate the number of moles of oxygen atoms in 64 g of oxygen atoms. *You need the mass of one mole of oxygen atoms. This is the Relative Atomic Mass in grams; in this case it is 16 g mol*$^{-1}$.

$$\textbf{number of moles of atoms} = \frac{\textbf{mass in grams}}{\textbf{molar mass of atoms}}$$

$$\therefore \text{ number of moles of oxygen} = \frac{64\text{ g of oxygen atoms}}{\text{molar mass of oxygen of } 16\text{ g mol}^{-1}}$$

= **4 moles of oxygen atoms**

b Calculate the number of moles of chlorine molecules in 142 g of chlorine gas.

$$\textbf{number of moles of atoms} = \frac{\textbf{mass in grams}}{\textbf{molar mass of atoms}}$$

The first stage of this calculation is to calculate the molar mass of Chlorine molecules.
Molar mass of Cl_2 = 2 x 35.5 = 71 g mol^{-1}

$$\therefore \text{ number of moles of chlorine} = \frac{142\text{ g of chlorine gas}}{\text{molar mass of chlorine of } 71\text{ g mol}^{-1}}$$

= **2 moles of chlorine molecules**

c Calculate the number of moles of $CuSO_4.5H_2O$ in 100 g of the solid.

The Relative Molecular Mass of $CuSO_4.5H_2O$ =

[63.5 + 32 + (4 x 16) + 5{(2x1) + 16}] = 249.5 g mol^{-1}

$$\therefore \text{ number of moles of } CuSO_4.5H_2O = \frac{100\text{ g of } CuSO_4.5H_2O}{\text{molecular mass of } CuSO_4.5H_2O \text{ of } 249.5\text{ g mol}^{-1}}$$

= **0.4008 moles of $CuSO_4.5H_2O$ molecules**

Example 2

Calculation of the mass of material in a given number of moles of that material

The mass of a given number of moles	**=**	**the mass of 1 mole**	**x**	**the number of moles of material concerned**

a Calculate the mass of 3 moles of sulphur dioxide SO_2

1 mole of sulphur dioxide has a mass = 32 + (2x16) = 64 g mol^{-1}

$\therefore$ 3 moles of SO_2 = 3 x 64 **= 192 g**

b What is the mass of 0.05 moles of $Na_2S_2O_3.5H_2O$?

1 mole of $Na_2S_2O_3.5H_2O$ = [(23 x 2) + (32 x 2) + (16 x 3)] + 5[(2 x 1) +16]
= 248 g mol^{-1}

$\therefore$ 0.05 moles of $Na_2S_2O_3.5H_2O$ = 0.05 x 248 **= 12.4 g**

Example 3

Calculation of the volume of a given number of moles of a gas

All you need to remember is that **1 mole of any gas has a volume of 24 dm^3 (24000 cm^3) at room temperature and pressure**.

$\therefore$	**The volume of a given number of moles of gas**	=	**number of moles**	**x**	**24000 cm^3**

a What is the volume of 2 mol of carbon dioxide?

Remember you do not need to work out the molar mass to do this calculation as it does not matter what gas it is.

$\therefore$ 2 moles of carbon dioxide = 2 x 24000 cm^3 = 48000 cm^3 **= 48 dm^3**

b What is the volume of 0.0056 moles of chlorine molecules?

Volume of 0.0056 moles of chlorine = 0.0056 x 24000 cm^3 **= 134.4 cm^3**

Example 4

Calculation of the number of moles of gas in a given volume of that gas

$$\textbf{number of moles of gas} = \frac{\textbf{volume of gas in cm}^3}{\textbf{24 000 cm}^3}$$

a Calculate the number of moles of hydrogen molecules in 240 cm^3 of the gas.

$$\text{number of moles} = \frac{240\,\text{cm}^3}{24\ 000\,\text{cm}^3} = 0.010\,\text{moles}$$

b How many moles of a gas are there in 1000 cm^3 of the gas?

$$\text{number of moles of gas} = \frac{1000\,\text{cm}^3}{24\,000\,\text{cm}^3} = 0.0147\,\text{moles}$$

Example 5

Calculation of the volume of a given mass of gas

This calculation require you to apply the skills covered in the previous examples

Calculate the volume of 10 g of hydrogen gas.

This is a two-stage calculation a) you need to calculate how many moles of hydrogen gas are present, and b) you need to convert this to a volume.

$$\therefore \text{ number of moles of hydrogen } (H_2) = \frac{10\text{ g of hydrogen } (H_2)}{\text{molecular mass of hydrogen } (H_2) \text{ of } 2\text{ g mol}^{-1}}$$

$$= \quad 5 \text{ moles}$$

$\therefore$ 5 moles of hydrogen = 5 x 24000 cm^3 = 120000 cm^3 **= 120 dm^3**

Example 6

Calculation of the mass of a given volume of gas

This calculation require you to apply the skills covered in the previous examples

Calculate the mass of 1000 cm^3 of carbon dioxide

Again this is a two-stage calculation a) you need to calculate the number of moles of carbon dioxide and then b) convert this to a mass.

$$\therefore \text{ number of moles of } CO_2 = \frac{1000\text{ cm}^3 \text{ of } CO_2}{\text{volume of 1 mole of } CO_2 \text{ of } 24\,000\text{ cm}^3}$$

$$= \quad 0.0147 \text{ moles}$$

$\therefore$ 0.0147 moles of carbon dioxide = 0.0147 x 44 g **= 1.833 g**

Example 7

Calculation of the molar mass of a gas from mass and volume data for the gas

Calculations of this type require you to find the mass of 1 mole of the gas, ie 24 000 cm^3. This is the molar mass of the gas.

eg Calculate the Relative Molecular Mass of a gas for which 100 cm^3 of the gas at room temperature and pressure, have a mass of 0.0667 g

100 cm^3 of the gas has a mass of 0.0667 g

$$\therefore\ 24\,000\,\text{cm}^3 \text{ of the gas must have a mass of } = \frac{0.0667\,\text{g} \times 24\,000\,\text{cm}^3}{100\,\text{cm}^3}$$

$$= 16\text{ g}$$

$\therefore$ **The molar mass of the gas is 16 g mol^{-1}**

Exercise 4a

Calculation of the number of moles of material in a given mass of that material

In this set of calculations all the examples chosen are from the list of compounds whose molar mass you calculated in exercise 1.

In each case calculate the number of moles of the material in the mass stated.

1	9.00 g of H_2O
2	88.0 g of CO_2
3	1.70 g of NH_3
4	230 g of C_2H_5OH
5	560 g of C_2H_4
6	0.640 g of SO_2
7	80.0 g of SO_3
8	18.0 g of HBr
9	0.0960 g of H_2SO_4
10	3.15 g of HNO_3
11	19.3 g of NaCl
12	21.25 g of $NaNO_3$
13	2.25 g of Na_2CO_3
14	0.800 g of NaOH
15	17.75 g of Na_2SO_4

16	3.16 g of $KMnO_4$
17	32.33 g of K_2CrO_4
18	100 g of $KHCO_3$
19	7.63 g of KI
20	3.90 g of $CsNO_3$
21	0.111 g of $CaCl_2$
22	41.0 g of $Ca(NO_3)_2$
23	1.48 g of $Ca(OH)_2$
24	3.40 g of $CaSO_4$
25	41.6 g of $BaCl_2$
26	14.95 g of $CuSO_4$
27	13.64 g of $ZnCl_2$
28	1.435 g of $AgNO_3$
29	13.76 g of NH_4Cl
30	13.76 g of $(NH_4)_2SO_4$
31	23.4 g of NH_4VO_3
32	10.0 g of $KClO_3$
33	10.7 g of KIO_3
34	100 g of NaClO

35	1.70 g of $NaNO_2$
36	50.9 g of $CuSO_4.5H_2O$
37	19.6 g of $FeSO_4.7H_2O$
38	9.64 g of $(NH_4)_2SO_4.Fe_2(SO_4)_3.24H_2O$
39	12.4 g of $Na_2S_2O_3.5H_2O$
40	32.0 g of $(COOH)_2.2H_2O$
41	3.075 g of $MgSO_4.7H_2O$
42	40.0 g of $Cu(NH_3)_4SO_4.2H_2O$
43	6.00 g of CH_3CO_2H
44	3.10 g of CH_3COCH_3
45	0.530 g of $C_6H_5CO_2H$
46	4.79 g of $AlCl_3$
47	56.75 g of $Al(NO_3)_3$
48	8.35 g of $Al_2(SO_4)_3$
49	3.8 g of $FeSO_4$
50	200 g of $FeCl_2$

Exercise 4b

Calculation of the mass of material in a given number of moles of the material

In each case calculate the mass in grams of the material in the number of moles stated.

1	2 moles of H_2O
2	3 moles of CO_2
3	2.8 moles of NH_3
4	0.50 moles of C_2H_5OH
5	1.2 moles of C_2H_4
6	0.64 moles of SO_2
7	3 moles of SO_3
8	1 mole of HBr
9	0.012 moles of H_2SO_4
10	0.15 moles of HNO_3
11	0.45 moles of NaCl
12	0.70 moles of $NaNO_3$
13	0.11 moles of Na_2CO_3
14	2.0 moles of NaOH
15	0.90 moles of Na_2SO_4
16	0.050 moles of $KMnO_4$

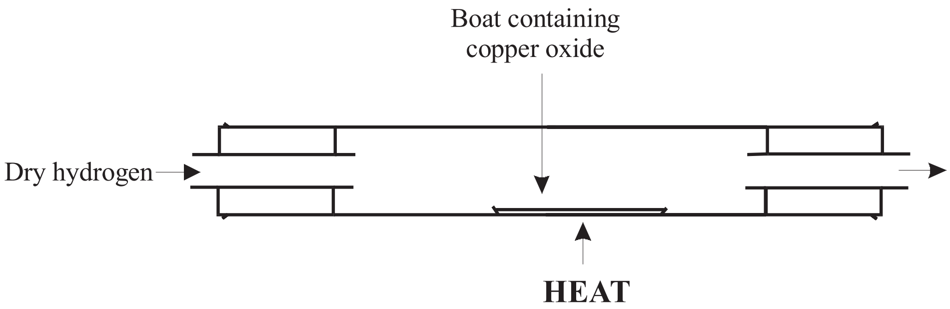
17	0.18 moles of K_2CrO_4
18	0.90 moles of $KHCO_3$
19	1.5 moles moles of KI
20	0.12 moles of $CsNO_3$
21	0.11 moles of $CaCl_2$
22	4.1 moles of $Ca(NO_3)_2$
23	0.0040 moles of $Ca(OH)_2$
24	0.10 moles of $CaSO_4$
25	0.21 moles of $BaCl_2$
26	0.10 moles of $CuSO_4$
27	0.56 moles of $ZnCl_2$
28	0.059 moles of $AgNO_3$
29	0.333 moles of NH_4Cl
30	1.1 moles of $(NH_4)_2SO_4$
31	0.025 moles of NH_4VO_3
32	0.10 moles of $KClO_3$
33	0.10 moles of KIO_3
34	10 moles of NaClO
35	0.0010 moles of $NaNO_2$

36	0.20 moles of $CuSO_4.5H_2O$
37	0.10 moles of $FeSO_4.7H_2O$
38	0.0050 moles of $(NH_4)_2SO_4.Fe_2(SO_4)_3.24H_2O$
39	0.040 moles of $Na_2S_2O_3.5H_2O$
40	2.4 moles of $(COOH)_2.2H_2O$
41	3.075 moles of $MgSO_4.7H_2O$
42	0.15 moles of $Cu(NH_3)_4SO_4.2H_2O$
43	0.17 moles of CH_3CO_2H
44	0.20 moles of CH_3COCH_3
45	0.080 moles of $C_6H_5CO_2H$
46	0.0333 moles of $AlCl_3$
47	0.045 moles of $Al(NO_3)_3$
48	0.12 moles of $Al_2(SO_4)_3$
49	2.0 moles of $FeSO_4$
50	11 moles of $FeCl_2$

Section 5

Using the idea of moles to find formulae

You can find the formula of copper(II) oxide by passing a stream of hydrogen over a known mass of copper oxide and weighing the copper formed.



- A known mass of copper(II) oxide is used.
- A stream of hydrogen from a cylinder is passed over the copper until all the air has been swept out of the apparatus.
- It is heated to constant mass (until two consecutive mass determinations at the end of the experiment are same) in a stream of *dry* hydrogen.
- The mass of the copper is finally determined.

Note:

- Excess hydrogen must **not** be ignited until it has been tested (by collection in a test tube) to make sure that all the air has been expelled from the apparatus. If the hydrogen in the test tube burns quietly, without a squeaky pop, then it is safe to ignite it at the end of the tube.
- The combustion tube is tilted to prevent the condensed steam from running back on to the hot part of the tube.
- When the reduction process is complete, ie after heating to constant mass, the tube is allowed to cool with hydrogen still being passed over the remaining copper. This is to prevent the copper from being oxidized to copper(II) oxide.

The working on the next page shows you how to calculate the results:

Typical results

Mass of copper (II) oxide	=	5 g
Mass of copper	=	4 g
Mass of oxygen	=	1 g

	÷ by relative atomic mass (r.a.m)	÷ by smallest	Ratio of atoms
Moles of copper atoms	$\frac{4}{64} = 0.0625$	$\frac{0.0625}{0.0625} = 1$	1
Moles of oxygen atoms	$\frac{1}{16} = 0.0625$	$\frac{0.0625}{0.0625} = 1$	1

Therefore, the simplest (or empirical) formula is CuO.

The apparatus may be modified to determine the formula of water. Anhydrous calcium chloride tubes are connected to the end of the combustion tube and the excess hydrogen then ignited at the end of these tubes. Anhydrous calcium chloride absorbs water; the mass of the tubes is determined at the beginning and end of the experiment. The increase in mass of the calcium chloride tubes is equal to the mass of water produced.

Typical results

Mass of water = 1.125 g
Mass of oxygen (from previous results) = 1.000 g
Mass of hydrogen = 0.125 g

	÷ by r.a.m	÷ by smallest	Ratio of atoms
Moles of hydrogen atoms	$\frac{0.125}{1} = 0.125$	$\frac{0.125}{0.0625} = 2$	2
Moles of oxygen atoms	$\frac{1}{16} = 0.0625$	$\frac{0.0625}{0.0625} = 1$	1

Since the ratio of hydrogen to oxygen is 2:1 the simplest (or empirical) formula is H_2O.

In examination calculations of this type the data is often presented not as mass, but as percentage composition of the elements concerned. In these cases the calculation is carried out in an identical fashion as percentage composition is really the mass of the element in 100 g of the compound.

Example 1

Sodium burns in excess oxygen to give a yellow solid oxide that contains 58.97% of sodium. What is the empirical formula of the oxide?

N.B. This is an oxide of sodium. It must contain only Na and O. Since the percentage of Na is 58.97 that of O must be 100 – 58.97 = 41.03%

	÷ by r.a.m	÷ by smallest	Ratio of atoms
Moles of sodium atoms	$\frac{58.97}{23} = 2.564$	$\frac{2.564}{2.564} = 1$	1
Moles of oxygen atoms	$\frac{41.03}{16} = 2.564$	$\frac{2.564}{2.564} = 1$	1

Therefore the empirical formula is **NaO**.

The result of the above calculation does not seem to lead to a recognisable compound of sodium. This is because the method used only gives the **simplest** ratio of the elements - but see below.

Consider the following series of organic compounds:

C_2H_4 ethene, C_3H_6 propene, C_4H_8 butene, C_5H_{10} pentene. These all have the same empirical formula CH_2.

To find the Molecular Formula for a compound it is necessary to know the Relative Molecular Mass (M_r).

Molecular Formula Mass = **Empirical Formula Mass** x **a whole number (n)**

In the example above the oxide has an M_r = 78 g mol^{-1}.

Thus

$$\begin{aligned} \text{Molecular Formula Mass} &= 78 \\ \text{Empirical Formula Mass} &= (\text{Na} + \text{O}) = 23 + 16 = 39 \\ \therefore\ 78 &= 39 \times n \\ \therefore\ n &= 2 \end{aligned}$$

The Molecular Formula becomes **$(NaO)_2$ or Na_2O_2**

Example 2

A compound **P** contains 73.47% carbon and 10.20% hydrogen by mass, the remainder being oxygen. It is found from other sources that **P** has a Relative Molecular Mass of 98 g mol^{-1}. Calculate the molecular formula of **P**.

NB It is not necessary to put in all the details when you carry out a calculation of this type. The following is adequate.

	C	**H**	**O**
	73.47	10.20	(100 – 73.47 – 10.20) = 16.33
By r.a.m	$\frac{73.47}{12}$ = 6.1225	$\frac{10.20}{1}$ = 10.20	$\frac{16.33}{16}$ = 1.020
By smallest	$\frac{6.1255}{1.020}$	$\frac{10.20}{1.020}$	$\frac{1.020}{1.020}$
Ratio of atoms	6	10	1

Therefore the empirical formula is $\mathbf{C_6H_{10}O}$.

To find molecular formula:

$$\begin{aligned} \text{Molecular Formula Mass} &= \text{Empirical Formula Mass x whole number (n)} \\ 98 &= [(6x12) + (10x1) + 16] \text{ x n} = 98 \text{ x n} \\ \therefore \text{n} &= 1 \end{aligned}$$

The molecular formula is the same as the empirical formula $\mathbf{C_6H_{10}O}$.

A warning

In calculations of this type at GCE Advanced level you may meet compounds that are different yet have very similar percentage composition of their elements. When you carry out a calculation of this type you should never round up the figures until you get right to the end. For example NH_4OH and NH_2OH have very similar composition and if you round up the data from one you may well get the other. If you are told the Relative Molecular Mass and your Empirical Formula Mass is not a simple multiple of this you are advised to check your calculation.

Example 3

Calculate the empirical formula of a compound with the following percentage composition:

C 39.13%; O 52.17%; H 8.700%

	C	**O**	**H**
By r.a.m	$\frac{39.13}{12}$	$\frac{52.17}{16}$	$\frac{8.700}{1}$
	= 3.26	= 3.25	= 8.70
Divide by smallest	1	1	2.66

It is clear at this stage that dividing by the smallest has not resulted in a simple ratio. **You must not round up or down at this stage**. You must look at the numbers and see if there is some factor that you could multiply each number by to get each one to a whole number. In this case if you multiply each by 3 you will get:

C	**O**	**H**
3	3	8

Thus $\mathbf{C_3H_8O_3}$ is the empirical formulae not $\mathbf{C_1H_{2.66}O_1}$

You need to watch carefully for this, the factors will generally be clear and will be 2 or 3. What you must not do is round 1.33 to 1 or 1.5 to 2.

Calculations involving the moles of water of crystallization

In calculations of this type you need to treat the water as a *molecule* and divide by the *Relative Molecular Mass*.

Example 4

24.6 grams of a hydrated salt of $MgSO_4 . xH_2O$, gives 12.0 g of anhydrous $MgSO_4$ on heating. What is the value of x ?

Your first job is to find the mass of water driven off .

Mass of water evolved = 24.6 – 12.0 = 12.6 g

	$\mathbf{MgSO_4}$	$\mathbf{H_2O}$
	12.0	12.6
Divide by M_r	$\frac{12.0}{120}$	$\frac{12.6}{18}$
	= 0.100	= 0.700
Ratio of Atoms	1	7

Giving a formula of $\mathbf{MgSO_4.7H_2O}$

Exercise 5

Calculation of a formula from experimental data

In *Section a.* calculate the empirical formula of the compound from the data given. This may be as percentage composition or as the masses of materials found in an experiment. For *Section b.* you are given the data for analysis plus the Relative Molecular Mass of the compound, in these cases you are to find the empirical formula and thence the molecular formula. *Section c.* is more difficult, the data is presented in a different fashion but the calculation of the empirical formula/ molecular formula is essentially the same.

Section a

1	Ca 40%; C 12%; O 48%
2	Na 32.4%; S 22.5%; O 45.1%
3	Na 29.1%; S 40.5%; O 30.4%
4	Pb 92.8%; O 7.20%
5	Pb 90.66%; O 9.34%
6	H 3.66%; P 37.8%; O 58.5%
7	H 2.44%; S 39.0%; O 58.5%
8	C 75%; H 25%

9	C 81.81%; H 18.18%
10	H 5.88% ; O 94.12%
11	H 5%; N 35%; O 60%
12	Fe 20.14%; S 11.51%; O 63.31%; H 5.04%

Section b

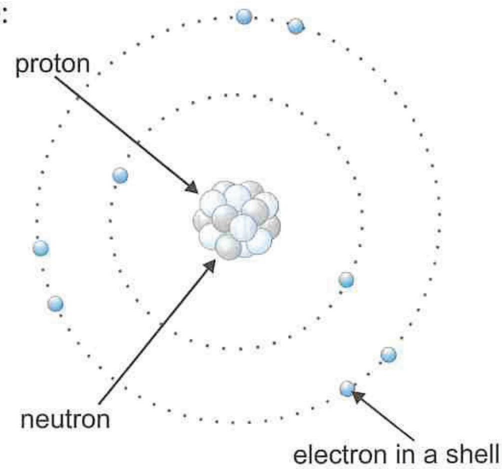
13	A hydrocarbon with a Relative Molecular Mass (M_r) of 28 g mol^{-1} has the following composition: Carbon 85.7%; Hydrogen 14.3%. Calculate its molecular formula.
14	A hydrocarbon with a Relative Molecular Mass (M_r) of 42 g mol^{-1} has the following composition: Carbon 85.7%; Hydrogen 14.3%. Calculate its molecular formula.
15	P 10.88%; I 89.12%. M_r = 570 g mol^{-1}
16	N 12.28%; H 3.51%; S 28.07%; O 56.14%. M_r = 228 g mol^{-1}
17	P 43.66%; O 56.34%. M_r = 284 g mol^{-1}
18	C 40%; H 6.67%; O 53.3%. M_r = 60 g mol^{-1}
19	Analysis of a compound with a M_r = 58 g mol^{-1} shows that 4.8 g of carbon are joined with 1.0 g of hydrogen. What is the molecular formula of the compound?
20	3.36 g of iron join with 1.44 g of oxygen in an oxide of iron that has a M_r = 160 g mol^{-1} What is the molecular formula of the oxide?
21	A sample of an acid with a M_r = 194 g mol^{-1} has 0.5 g of hydrogen joined to 16 g of sulphur and 32 g of oxygen. What is the molecular formula of the acid?
22	Analysis of a hydrocarbon showed that 7.8 g of the hydrocarbon contained 0.6 g of hydrogen and that the M_r = 78 g mol^{-1}. What is the formula of the hydrocarbon?

Section c

23	22.3 g of an oxide of lead produced 20.7 g of metallic lead on reduction with hydrogen. Calculate the empirical formula of the oxide concerned.
24	When 1.17 g of potassium is heated in oxygen 2.13 g of an oxide is produced. In the case of this oxide the empirical and molecular formulae are the same. What is the molecular formula of the oxide produced?
25	A hydrocarbon containing 92.3% of carbon has a Relative Molecular Mass of 26 g mol^{-1}. What is the molecular formula of the hydrocarbon?
26	When 1.335 g of a chloride of aluminium is added to excess silver nitrate solution 4.305 g of silver chloride is produced. Calculate the empirical formula of the chloride of aluminium. *Hint; you will need to work out how much chlorine there is in 4.305 g of AgCl. This will be the amount of chlorine in the initial 1.335 g of the aluminium chloride.*
27	16 g of a hydrocarbon burn in excess oxygen to produce 44 g of carbon dioxide. What is the empirical formula of the hydrocarbon. *Hint; you will need to work out what mass of carbon is contained in 44 g of CO_2. This is the mass of C in 16 g of the hydrocarbon.*
28	When an oxide of carbon containing 57.1% oxygen is burnt in air the percentage of oxygen joined to the carbon increases to 72.7%. Show that this data is consistent with the combustion of carbon monoxide to carbon dioxide.
29	When 14.97 g of hydrated copper(II) sulphate is heated it produces 9.60 g of anhydrous copper(II) sulphate. What is the formula of the hydrated salt?
30	When the chloride of phosphorus containing 85.1% chlorine is heated a second chloride containing 77.5% chlorine is produced. Find the formulae of the chlorides and suggest what the other product of the heating might be.

Atomic Structure

What Are *Atoms* Like?



1) Atoms are made up of **three** types of **subatomic particle**: **protons**, **neutrons** and **electrons**.
2) In the **centre** of all atoms is a **nucleus** containing **neutrons** and **protons**.
3) Almost all of the **mass** of the atom is contained in the **nucleus** which has an overall **positive** charge. The positive charge arises because each of the **protons** in the nucleus have a **+1** charge.
4) The **neutrons** in the nucleus have a very similar **mass** to the protons but they are **uncharged**.
5) **Electrons** are much **smaller** and **lighter** than either the neutrons or protons. They have a **negative charge** (–1) and **orbit** the nucleus in **shells** (or energy levels).
6) There's an **attraction** between the **protons** in the nucleus and the **electrons** in the shells.
7) The nucleus is **tiny** compared with the total volume occupied by the whole atom.
8) The **volume** occupied by the **shells** of the electrons determines the **size** of the atom.

Here's a round up of the **properties** of the subatomic particles:

Particle	Relative Mass	Relative Charge
Proton	1	+1
Neutron	1	0
Electron	$\frac{1}{2000}$	–1

What is the *Charge* on an Atom?

The overall charge on an atom is **zero**.

This is because each **+1** charge from a **proton** in the nucleus is **cancelled out** by a **–1** charge from an **electron**.

If an atom **loses** or **gains** electrons it becomes **charged**. These charged particles are called **ions**.

EXAMPLE: How many electrons has an Al^{3+} ion lost or gained?

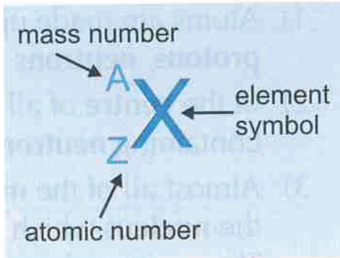
The Al^{3+} ion has a charge of **+3**, so there must be **3 more protons** than **electrons**.
Ions are formed when **electrons** are lost or gained, so Al^{3+} must have **lost 3 electrons**.

Neutrons are the perfect criminals — they never get charged...

1) Which subatomic particles are found in the nucleus?
2) What is the charge on an ion formed when an atom loses two electrons?
3) What is the charge on an ion formed when an atom gains two electrons?

Atomic Number, Mass Number and Isotopes

Atomic and Mass Numbers



1) If you look at an element in the periodic table, you'll see it's given **two numbers**. These are the **atomic number** and the **mass number**.
2) The **atomic number** of an element is given the symbol **Z**. It's sometimes called the **proton number** as it represents the number of **protons** in the nucleus of the element.
3) For **neutral** atoms the number of **protons equals** the number of **electrons**, but you need to take care when considering ions as the number of electrons changes when an ion forms from an atom.
4) The **mass number** of an atom is given the symbol **A**. It represents the **total** number of **neutrons** and **protons** in the nucleus.
5) **Subtracting Z** from **A** allows you to calculate the number of **neutrons** in the nucleus.

EXAMPLE: Use the periodic table to complete the following information about sodium.

Element	Symbol	Z	A	No. Protons	No. Neutrons	No. Electrons
Sodium			23			

The periodic table tells you that the **symbol** for sodium is **Na** and **Z** is **11**.
The number of **protons** in sodium is the same as the **atomic number**, which is **11**.
You work out the number of **neutrons** by **subtracting** Z from A: 23 – 11 = **12**.
The number of **electrons** is the **same** as the number of protons, which is **11**.

Isotopes

1) Atoms of the same **element** always have the same number of **protons**, so they'll always have the same **atomic number**, but their **mass numbers** can **vary** slightly.
2) Atoms of the same **element** with different **mass numbers** are called **isotopes**.
3) Isotopes have the same number of **protons** but different numbers of **neutrons** in their nuclei.

EXAMPLE: Copper has an atomic number of 29. Its two main isotopes have mass numbers of 63 and 65. How many neutrons does each of the isotopes have?

The ^{63}Cu isotope has 63 – 29 = **34 neutrons**.
The ^{65}Cu isotope has 65 – 29 = **36 neutrons**.

Finding the number of neutrons — it's as easy as knowing your A – Z...

1) Use the periodic table to work out how many neutrons are in a neutral phosphorus atom.
2) In terms of the numbers of subatomic particles, state two similarities and one difference between two isotopes of the same element.
3) Three neutral isotopes of carbon have mass numbers 12, 13 and 14. State the numbers of protons, neutrons and electrons in each.

Relative Atomic Mass

Calculating the **Relative Atomic Mass**

1) The average mass of an element is called its **relative atomic mass**, or A_r.
2) When you look up the **relative atomic mass** of an element on a **detailed** copy of the periodic table, you'll see that it isn't always a **whole number**. This is because the value given is the **average** mass number of two or more **isotopes**.
3) The **value** of the relative atomic mass is further complicated by the fact that some isotopes are **more abundant** than others. It's a **weighted average** of all the element's different isotopes.
4) You can use the **relative abundances** and **relative isotopic masses** (the mass number of a single, specific isotope) of each isotope to work out the **relative atomic mass** of an element.
5) Relative abundances of isotopes are often given as **percentages**. To work out the **relative atomic mass** of an element, all you need to do is multiply **each isotopic mass** by its **relative abundance**, add all the values together and divide by **100**.

EXAMPLE: What is the relative atomic mass of chlorine given that 75% of atoms have an atomic mass of 35 and 25% of atoms have an atomic mass of 37?

Average mass = (abundance of $^{35}Cl \times 35$ + abundance of $^{37}Cl \times 37$) $\div$ 100

$= [(75 \times 35) + (25 \times 37)] \div 100$

$= (2625 + 925) \div 100$

$= 3550 \div 100$

$=$ **35.5** (You can check your answer against a periodic table to see if it's right.)

Calculating the **Relative Formula Mass**

If you **add up** the relative atomic masses of all the atoms in a chemical formula, you get the **relative formula mass**, or M_r, of that compound.

(If the compound is molecular, you might hear the term relative molecular mass used instead, but it means pretty much the same.)

EXAMPLE: Calculate the relative formula mass of $CaCl_2$.

Ca has an atomic mass of 40.1 and Cl has an atomic mass of 35.5.

$M_r = (1 \times 40.1) + (2 \times 35.5)$

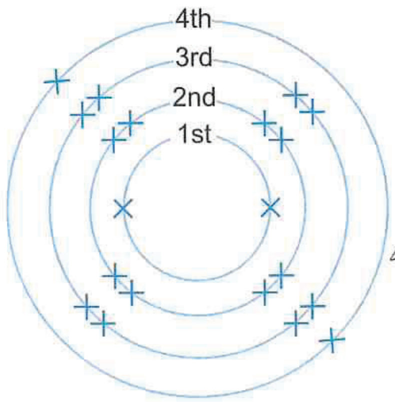
$=$ **111**

Together, my brother and I weigh 143 kg — it's our relative mass...

1) Find the relative atomic mass of lithium if its composition is 8% ^{6}Li and 92% ^{7}Li.
2) Find the relative atomic mass of carbon if its composition is 99% ^{12}C and 1% ^{13}C.
3) Find the relative atomic mass of silver if its composition is 52% ^{107}Ag and 48% ^{109}Ag.
4) Find the relative formula mass of sodium fluoride, NaF.
5) Find the relative formula mass of chloromethane, CH_3Cl.

Electronic Structure

Electrons are Arranged in *Energy Shells*



1) Electrons orbit the nucleus in **shells** (also called **energy levels**).
2) You can draw concentric **circles** to represent the different **shells**. Then add **crosses** to represent the **electrons** at each level.
3) For example, this diagram shows the energy levels, for an atom with 20 electrons, filling up with electrons. It has two electrons in the first shell, eight in the second shell, eight in the third shell and two in the fourth shell.
(Remember you should always **start** filling the **innermost levels** first.)

Here's another way to show electron arrangements using simple notation:

An atom with 6 electrons:	2, 4
An atom with 11 electrons:	2, 8, 1
An atom with 20 electrons:	2, 8, 8, 2

The first number tells you how many electrons are in the first shell, the second number tells you how many electrons are in the second shell, and so on.

Energy Levels are *Split* into *Subshells*

Subshell	Maximum electrons
s	2
p	6
d	10

1) Energy levels can be **split** into **sub-levels** called **subshells**. The first three subshells are called '**s**', '**p**' and '**d**'. They can each hold a different **number** of electrons.
2) The first energy level has **one subshell** — an 's' level. So the first energy level can contain up to **2 electrons**.
3) At GCSE you learnt that the second energy level can contain up to **8 electrons**. It's actually split into **2 sub-levels**. **Two** of the electrons are in an 's' level and the remaining **six** are in a '**p**' level. If you combine the 2 's' electrons with the 6 'p' electrons you get a total of 8.
4) Electrons generally start by filling the **energy level** with the **lowest energy**. So the **first** energy level will be completely **filled** before any electrons go into the **second** energy level. Within an energy level, electrons will fill the **subshells** in the order **s**, then **p**, then **d**.
5) As well as telling you how many electrons are in each **shell**, the **electron configuration** of an atom also tells you what **subshells** the electrons are in. For an atom with 10 electrons:

The big number tells you the energy level.

$$1s^2\,2s^2\,2p^6$$

The little number tells you how many electrons are in that subshell.

The letter tells you the subshell.

I'm trying to be calm, but my energy level is too high...

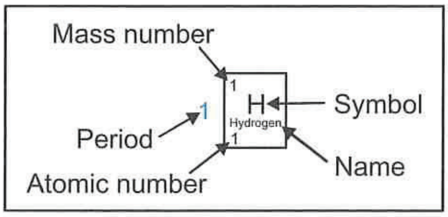
1) Draw diagrams to show the electron arrangements of the following elements: carbon, fluorine, magnesium, sulfur.
2) Use the simple notation shown above to write the electron arrangements of these elements: lithium, sodium, potassium, beryllium, magnesium, calcium.
3) Give the electron configurations of oxygen and chlorine.

The Periodic Table

*The **Periodic Table***

The periodic table contains:

- All of the elements in order of atomic number.
- Vertical groups of elements which have similar properties.
- Horizontal rows of elements called periods.



Period	Group 1	Group 2											Group 3	Group 4	Group 5	Group 6	Group 7	Group 0
																		4 He Helium 2
2	7 Li Lithium 3	9 Be Beryllium 4											11 B Boron 5	12 C Carbon 6	14 N Nitrogen 7	16 O Oxygen 8	19 F Fluorine 9	20 Ne Neon 10
3	23 Na Sodium 11	24 Mg Magnesium 12											27 Al Aluminium 13	28 Si Silicon 14	31 P Phosphorus 15	32 S Sulfur 16	35.5 Cl Chlorine 17	40 Ar Argon 18
4	39 K Potassium 19	40 Ca Calcium 20	45 Sc Scandium 21	48 Ti Titanium 22	51 V Vanadium 23	52 Cr Chromium 24	55 Mn Manganese 25	56 Fe Iron 26	59 Co Cobalt 27	59 Ni Nickel 28	63.5 Cu Copper 29	65 Zn Zinc 30	70 Ga Gallium 31	73 Ge Germanium 32	75 As Arsenic 33	79 Se Selenium 34	80 Br Bromine 35	84 Kr Krypton 36
5	85.5 Rb Rubidium 37	88 Sr Strontium 38	89 Y Yttrium 39	91 Zr Zirconium 40	93 Nb Niobium 41	96 Mo Molybdenum 42	98 Tc Technetium 43	101 Ru Ruthenium 44	103 Rh Rhodium 45	106 Pd Palladium 46	108 Ag Silver 47	112 Cd Cadmium 48	115 In Indium 49	119 Sn Tin 50	122 Sb Antimony 51	128 Te Tellurium 52	127 I Iodine 53	131 Xe Xenon 54
6	133 Cs Caesium 55	137 Ba Barium 56	139 La Lanthanum 57	178.5 Hf Hafnium 72	181 Ta Tantalum 73	184 W Tungsten 74	186 Re Rhenium 75	190 Os Osmium 76	192 Ir Iridium 77	195 Pt Platinum 78	197 Au Gold 79	201 Hg Mercury 80	204 Tl Thallium 81	207 Pb Lead 82	209 Bi Bismuth 83	210 Po Polonium 84	210 At Astatine 85	222 Rn Radon 86
7	223 Fr Francium 87	226 Ra Radium 88	227 Ac Actinium 89															

s-block (Groups 1–2); d-block (Sc–Zn columns); p-block (Groups 3–0)

140 Ce Cerium 58	141 Pr Praseodymium 59	144 Nd Neodymium 60	147 Pm Promethium 61	150 Sm Samarium 62	152 Eu Europium 63	157 Gd Gadolinium 64	159 Tb Terbium 65	162 Dy Dysprosium 66	165 Ho Holmium 67	167 Er Erbium 68	169 Tm Thulium 69	173 Yb Ytterbium 70	175 Lu Lutetium 71
232 Th Thorium 90	231 Pa Protactinium 91	238 U Uranium 92	237 Np Neptunium 93	242 Pu Plutonium 94	243 Am Americium 95	247 Cm Curium 96	247 Bk Berkelium 97	251 Cf Californium 98	254 Es Einsteinium 99	253 Fm Fermium 100	256 Md Mendelevium 101	254 No Nobelium 102	257 Lr Lawrencium 103

*What Are **Groups** and **Periods**?*

Chemical reactions involve atoms **reacting** to gain a **full outer shell** of electrons.
All of the elements in a group have the same **number** of **electrons** in their **outer shell**.
As a result, the elements in a group **react** in a **similar** way.

The properties of elements in the same **period change gradually** as you move from one side of the periodic table to the other.

*The **Periodic Table** is Split into **Blocks***

1) As well as being split into groups and periods, the periodic table has four **blocks**. You only need to worry about two of them at the moment though — the '**s**' block and the '**p**' block.
2) Groups **1** and **2** are called the **s-block** elements. Their outer electrons are in energy levels called **s subshells**. S subshells can accommodate up to 2 electrons (see page 4).
3) Groups **3** to **0** are called the **p-block** elements. Their outer electrons are in energy levels called **p subshells**. P subshells can accommodate up to 6 electrons (see page 4).
4) There's always one exception. **Helium** (He) is an **s-block** element, even though it's in **Group 0**. Its electron configuration is $1s^2$.

The mystery of the periodic table? It's elementary, my dear Watson...

1) Sort the following elements into a table to show which ones are from the s-block, and which are from the p-block: caesium, potassium, phosphorus, calcium, aluminium, barium and sulfur.
2) Give one similarity between elements that are in the same group.