Transition Pack for A Level Chemistry

SARACENS
HIGH SCHOOL

Get ready for A-level!
A guide to help you get ready for A-level Chemistry, including everything from topic guides to days out and online learning courses.

Contents and instructions

Section 1: Fundamental topics information and questions

This section is compulsory

Task: read through the assigned chapters. Answer all questions on these pages, optional: make flash cards or revision notes from these pages.

- Chapter 1 – structure of the atom
- Chapter 4 – bonding and properties
- Chapter 5 – chemical equations
- Chapter 11 – calculations
- Chapter 12 – enthalpy
- Chapter 13 – investigating and interpreting

Section 2: Pre-Knowledge quizzes

This section is compulsory

Task: read the information and watch the relevant videos/ read the relevant notes (scan the QR code). Then answer the questions that follow for topic 1-10. Your notes from your previous task may help you!

Section 3: Further information

This section is not compulsory

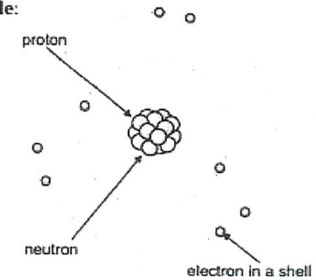
Task: pick one further activity to complete. Do this and write a reflection of the chemistry you saw or experienced here. You can hand this to your teacher in September for feedback if you opt to do this!

- Recommended books
- Chemistry related films to watch
- Research activities
- Places to visit

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.**

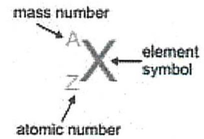
Neutrons are the perfect criminals — they never get charged...

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...

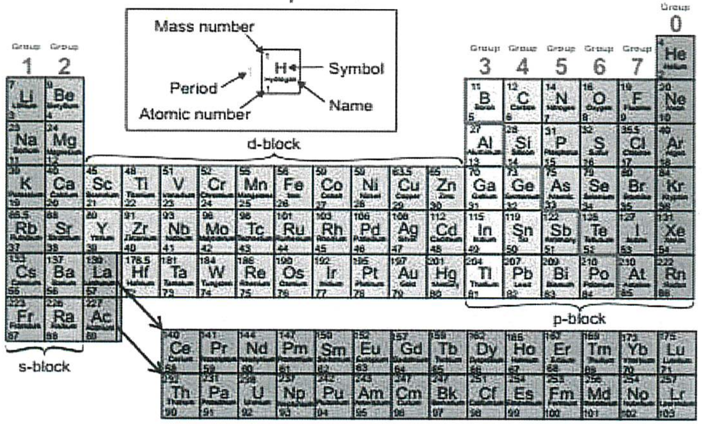
1) Use the periodic table to work out how many neutrons are in a neutral phosphorus atom.
2) In terms of the numbers of subatomic particles, state two similarities and one difference between two isotopes of the same element.
3) Three neutral isotopes of carbon have mass numbers 12, 13 and 14. State the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

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.



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 $1s^2$.

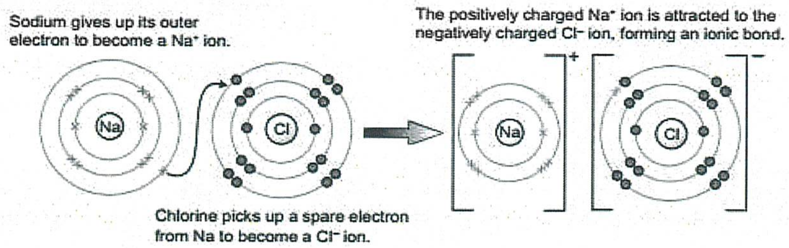
The mystery of the periodic table? It's elementary, my dear Watson...

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.

Ionic Bonding

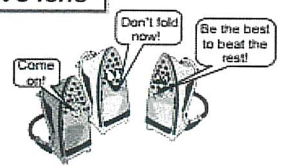
Ionic Bonds Involve the Transfer of Electrons

1) Ions form when **electrons** are transferred from **one atom** to **another**. Atoms that **lose electrons** form **positive ions** and atoms that **gain electrons** become **negative ions**.
2) These oppositely charged ions are **attracted** to each other by **electrostatic attraction**. When this happens, an **ionic bond** is formed.
3) The simplest ions form when atoms lose or gain 1, 2 or 3 electrons to get a **full outer shell**.
4) You can show the transfer of electrons to form an ionic compound using a **dot-and-cross** diagram. For example, sodium and chlorine will react to form sodium chloride (NaCl):



5) In the example above, the **dots** represent the electrons that come from the chlorine atom, and the **crosses** represent the electrons that come from the sodium atom.

You Can Find The Ratio of Positive to Negative Ions



1) The **ratio** of positive ions to negative ions in an ionic compound depends on the **charges** of the ions.
2) The **overall charge** of an ionic compound is **zero**, so the **sum** of all the **positive charges** in the compound must be **equal** to the **sum** of the **negative charges**.
3) If you know the **individual charges** of each of the ions in a compound, you can work out their **ratio**. You can use this to find the **ionic formula** of the compound.

EXAMPLE: In the compound calcium chloride, what is the ratio of Ca^{2+} to Cl^- ions?

For the compound to be neutral it must contain **two Cl^-** ions (2×-1) to **balance** the charge of **each Ca^{2+}** ion ($1 \times +2$).
So the ratio of Ca^{2+} ions to Cl^- ions in the compound must be **1 : 2**.
The ionic formula will be **$CaCl_2$**.

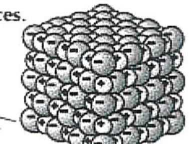
I can't afford Mg^{2+} — the charge is just too high...

1) Draw a diagram showing how a magnesium atom reacts with an oxygen atom to form magnesium oxide, MgO. Your diagram should show the electron transfer process.
2) In potassium oxide, what is the ratio of K^+ ions to O^{2-} ions? What is the ionic formula?

Ionic Compounds

Ionic Bonds Produce Giant Ionic Structures

1) Ionic bonds do not work in any particular direction. The electrostatic attraction is just as strong in **all directions** around the ion.
2) This means that when ionic compounds form, they produce **giant lattices**.
3) The lattice is a closely packed **regular** array of ions, with each negative ion **surrounded** by positive ions and vice versa. The **forces** between the **oppositely charged** ions are very **strong**.
4) **Sodium chloride** forms a lattice like this one. This is called the sodium chloride structure.



Ionic Bond Strength Depends on the Charge on the Ions

The **strength** of the bonds between ions is **not the same** in all ionic structures:

The **bigger** the charges on the ions, the **stronger** the attraction.

For example, the bonds between the ions in **MgO** ($Mg^{2+}O^{2-}$) will be **stronger** than those between the ions in **NaCl** (Na^+Cl^-).

Physical Properties of Ionic Compounds

Melting points

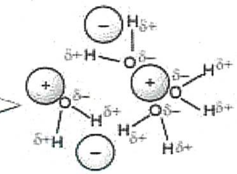
In order to **melt** a solid, the forces holding the particles together have to be **overcome**. In an ionic solid, these bonds are very **strong**, so a **large** amount of energy is required to break them. So, ionic compounds have very **high** melting points.

Electrical conductivity

In their solid form, ionic compounds are electrical **insulators** (they don't conduct electricity). They have **no free ions** or electrons to carry electric current.
When **molten** or **dissolved**, the ions **separate** and are **free** to move and conduct electricity.
So **all** ionic compounds **conduct** electricity when **molten** or **dissolved**.

Solubility

In many cases ionic compounds are **soluble** in water.
This happens because water is a **polar** molecule (see page 10) — the positive end of the molecule points towards the negative ions and the negative end towards the positive ions.
Although **lots of energy** is required to break the strong bonds within the lattice, it is provided by the formation of **many weak bonds** between the water molecules and the ions in solution.



Rabbits love studying ionic compounds — all those giant lettuces...

1) Put these ionic compounds in order of melting point, highest to lowest: Lithium oxide (Li_2O), Beryllium oxide (BeO), Lithium fluoride (LiF). Explain why you have put them in that order.
2) Explain why the ionic compound, potassium chloride (KCl), can conduct electricity when molten or dissolved, but not when it is solid.

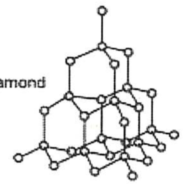
Giant Covalent Structures

Giant Covalent Structures

Carbon is ideally placed to share electrons and form covalent bonds, because it has a **half-full** outer shell. Carbon atoms can share their electrons with four other carbons to gain a full outer shell. This can result in the formation of a single massive carbon molecule — a **giant structure**.

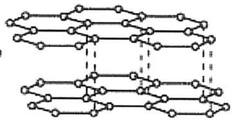
Carbon can form various different **giant covalent structures** such as **diamond** and **graphite**.

diamond



Each carbon atom forms **four** covalent bonds in a very **rigid** structure.
This structure makes diamond very **hard**.

graphite



Each carbon atom forms **three** covalent bonds in the same **plane**. This results in a series of **layers** which can **slide** over each other.
The fourth electron from each carbon atom is **free**.

Properties of Giant Covalent Structures

Giant covalent structures have some different **physical properties** from small molecules.

Melting points

Unlike small molecules, melting points are **extremely high**, as all of the atoms are held together by **strong covalent bonds**. These millions of covalent bonds need to be **broken** to allow the atoms within the structure to move freely, which requires a lot of energy.

This contrasts with small molecules where no covalent bonds (only intermolecular bonds) need to be broken in order for the substance to melt.

Electrical conductivity

Giant covalent structures are **electrical insulators**. This is because they don't contain **charged particles**, and the atoms aren't free to move.
Even a **molten** covalent compound will not conduct electricity.
Graphite is the only exception to this, as the loosely held **electrons** between the layers of atoms can move through the solid structure. Graphite conducts in both its solid and liquid forms.

Solubility

Giant covalent structures are **not soluble** in water. To get a giant covalent structure to dissolve, all the covalent bonds joining the atoms together would need to be **broken**. There is no way to get the energy required to do this, since the individual **neutral atoms** in the structure will **not** form intermolecular bonds with the water molecules.

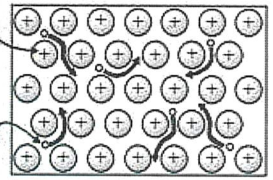
Diamonds — don't mess with 'em — they're well 'ard...

1) Devise a series of tests that would allow you to distinguish between two unknown crystalline solids, one of which is an ionic compound and the other a giant covalent structure.
2) Why won't diamond dissolve in water when sodium chloride will?

Metallic Bonding

Metals have Giant Structures Too

1) In a metal, the **outer electrons** from each atom are **delocalised** (they're not stuck on one atom) — this leaves **positive metal ions**.
2) The positive metal ions are arranged regularly in a **giant structure**, surrounded by a 'sea' of delocalised electrons.
3) Metals are held together because of the **electrostatic attractions** between the **positive metal ions** and the **delocalised** 'sea' of **electrons**.
This is called **metallic bonding**.



Properties of Metals

Metallic bonding explains the **physical properties** of metals:

Melting points

Metals generally have **high** melting points. This is because a lot of energy is required to overcome the **strong metallic bonding** between the particles.

The **more** electrons that are **delocalised** from **each atom**, the **stronger** the bonding will be and the **higher** the melting point.

> **EXAMPLE:** Predict, with reasoning, whether magnesium or sodium will have a higher melting point.
>
> Magnesium is made up of Mg^{2+} ions with **two** delocalised electrons per atom.
> Sodium is made up of Na^+ ions and only **one** delocalised electron per atom.
> So **magnesium** will have a **higher melting point** than sodium, because the metallic bonds will be **stronger** and require **more energy** to break.

Electrical conductivity

The **delocalised electrons** in metals are **free to move** around and can carry a **current**. This makes metals **good electrical conductors**.

Solubility

The **strong metallic bonds** mean that metals are generally **insoluble**.

Metallica bonds — friendships based on a love of '80s rock music...

1) Predict, with reasoning whether potassium or calcium will have a higher melting point.
2) Draw a diagram to show the bonding in a sample of sodium.
3) Sodium has a metallic structure, whilst sodium chloride (NaCl) is an ionic compound. Give one similarity and one difference between the physical properties of these substances.

Writing and Balancing Equations

Reaction Equations Show How Chemicals React Together

1) A reaction equation shows what happens during a chemical reaction.
   The **reactants** are shown on the **left hand side**, and the **products** on the **right hand side**.
2) **Word equations** just give the **names** of the components in the reaction.
   For example: propane + oxygen → carbon dioxide + water
3) **Symbol equations** give the chemical formulae of all the different components.
   They show all the **atoms** that take part in the reaction, and how they rearrange.
   For example: $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$
4) Symbol equations have to **balance** — there has to be the **same number** of each **type** of atom on each side of the equation. The big numbers in front of each substance tell you how much of that particular thing there has to be for all the atoms to balance.

Writing **Balanced** Equations

To write a balanced symbol equation for a reaction there are 4 simple steps:

1) Write out the **word equation** first.
2) Write the correct **formula** for each substance below its name.
3) Go through each element in turn, making sure the **number of atoms** on each **side** of the equation **balances**. If your equation isn't balanced, you can only add more atoms by adding **whole reactants** or **products**.
4) If you changed any numbers, do step 3 again, and repeat until **all** the elements **balance**.

Doing the third step:

If the atoms in the equation don't balance you **can't** change the **molecular formulae** — only the numbers in **front** of them.
For example: $CaO + HCl \rightarrow CaCl_2 + H_2O$
There are **two Cl** atoms on the **right-hand side** of the equation, so we need to have **two HCl** on the **left-hand** side: $CaO + 2HCl \rightarrow CaCl_2 + H_2O$
This also doubles the number of **hydrogen atoms** on the left-hand side, so that the hydrogens **balance** as well.

EXAMPLE: Write a balanced equation for the reaction of magnesium with hydrochloric acid.

Step 1 — Write the word equation:
magnesium + hydrochloric acid → magnesium chloride + hydrogen

Step 2 — Write the symbol equation: $Mg + HCl \rightarrow MgCl_2 + H_2$

Step 3 — Go through the equation and balance the elements one by one:
$Mg + \mathbf{2}HCl \rightarrow MgCl_2 + H_2$
(the Mgs balance, but there are different amounts of H and Cl on each side. Put a 2 in front of HCl to balance the Hs and Cls. Check everything still balances.)

Writing and Balancing Equations

In **Ionic Equations** Make Sure the **Charges** Balance

1) In some reactions, particularly those in solution, not all the particles take part in the reaction.
2) **Ionic equations** are chemical equations that just show the **reacting particles**.
3) As well as having the same number of **atoms** of each element on each side of the equation, in ionic equations you need to make sure the **charge** is the same on both sides.

EXAMPLE: Balance the following ionic equation: $Na + H^+ \rightarrow Na^+ + H_2$

First, balance the **number of atoms** of each element using the method on the last page:
$Na + \mathbf{2}H^+ \rightarrow Na^+ + H_2$

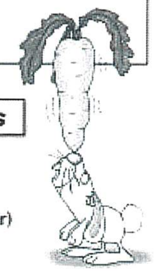
Then check the **charge** is the same on both sides of the equation:

- On the left hand side, each H^+ ion contributes +1, so the charge is 2 × +1 = **+2**.
- On the right hand side, the sodium ion contributes +1, so the charge is 1 × +1 = **+1**.

To get the charges to balance, you need another positive charge on the right-hand side. One way of doing this is by adding another sodium ion to the products:
$Na + 2H^+ \rightarrow \mathbf{2}Na^+ + H_2$

Now check that the number of atoms still balances:
The Hs balance, but there are 2Nas on the right-hand side, and only one on the left.
So put a 2 in front of the left-hand side Na:
$\mathbf{2}Na + 2H^+ \rightarrow 2Na^+ + H_2$

The atoms **and** charges on each side balance, so that's your final answer.



Chemical Equations Sometimes Include **State Symbols**

State symbols show the **physical state** that a substance is in.
The state symbols you need to know about are in the box below:

(l) — liquid (g) — gas (s) — solid (aq) — aqueous (dissolved in water)

So the balanced equation for the reaction between hydrochloric acid and magnesium, including state symbols is: $Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$.

Hold one ear and stare at something still — it'll help you balance...

1) Write a balanced symbol equation for the combustion of methane (CH_4) in oxygen.
   Step 1 has been done for you.
   Step 1: Methane + oxygen → carbon dioxide + water
2) Write balanced symbol equations for the following reactions.
   a) The complete combustion of ethanol (C_2H_5OH) in oxygen (O_2) to give carbon dioxide (CO_2) and water (H_2O).
   b) The reaction of calcium hydroxide ($Ca(OH)_2$) with hydrochloric acid (HCl) to give calcium chloride ($CaCl_2$) and water (H_2O).
3) Balance the following ionic equation: $Cl_2 + Fe^{2+} \rightarrow Cl^- + Fe^{3+}$.
   Include state symbols given that Cl_2 is a gas and everything else is aqueous.

The Mole

A Mole is a Number of Particles

If you had a sample of a substance, and you wanted to **count** the number of atoms that were in it, you'd have to use some very **big numbers**, and spend a very long time counting. So you need a **unit** to describe the **amount** of a substance that you have — that unit is the **mole**.

One mole of a substance contains **6.02×10^{23}** particles.
6.02×10^{23} mol^{-1} is known as **Avogadro's constant**.

The particles can be **anything** — e.g. atoms or molecules (or even giraffes).
So 6.02×10^{23} atoms of **carbon** is 1 mole of carbon,
and 6.02×10^{23} molecules of CO_2 is 1 mole of CO_2.

Molar Mass is the Mass of One Mole

One mole of atoms or molecules has a **mass in grams** equal to the **relative formula mass** (A_r or M_r) of that substance.

For **carbon**, A_r = **12.0** so 1 mole of carbon weighs **12 g** and the **molar mass** is **12 g mol^{-1}**.
For **CO_2**, M_r = **44.0** so 1 mole of CO_2 weighs **44 g** and the **molar mass** of CO_2 is **44 g mol^{-1}**.
So, **12.0 g** of **carbon** and **44.0 g** of **CO_2** must contain the **same number of particles**.

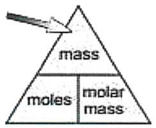
You can use molar mass in calculations to work out how many moles of a substance you have.

Just use this formula: $\text{Number of moles} = \dfrac{\text{Mass of substance (g)}}{\text{Molar mass (g mol}^{-1})}$ ← g mol^{-1} is the same as g/mol.

EXAMPLE: How many moles of sodium oxide are present in 24.8 g of Na_2O?

Molar mass of Na_2O = $(2 \times 23.0) + (1 \times 16.0) = 62.0$ g mol^{-1}
Number of moles of Na_2O = 24.8 g ÷ 62.0 g mol^{-1} = **0.400 moles**

You can **rearrange** the formula above and use it to work out the mass of a substance or its relative formula mass (see page 3). It can help to remember this triangle:



EXAMPLE: What is the mass of 1.30 moles of magnesium oxide (MgO)?

Molar mass of MgO = $(1 \times 24.3) + (1 \times 16.0) = 40.3$ g mol^{-1}
Rearranging the formula, mass = moles × molar mass
So mass of MgO = 1.30×40.3 = **52.4 g** (3 s.f.)

Avocado's constant: how much I need to satisfy my guacamole craving...

1) Find the molar mass of sulfuric acid, given that 0.700 moles weighs 68.6 g.
2) How many moles of sodium chloride are present in 117 g of NaCl?
3) I have 54.0 g of water (H_2O) and 84.0 g of iron (Fe). Do I have more moles of water or of iron?

Determination of Formulae from Experiments

Empirical and Molecular Formulae

The **empirical formula** of a compound is the **simplest ratio** of the atoms of each element in the compound.
The **molecular formula** of a compound gives the **actual number** of atoms of each element in the compound.
For example, a compound with the molecular formula **C_2H_6** has the empirical formula **CH_3**.
The **ratio** of the atoms is one C to every three Hs.

Calculating Empirical Formulae

Often, the only way to find out the formula of a compound is through **experimentation** and **calculation**. You can calculate the formula of a compound from the **masses** of the **reactants**.
Here is a simple set of rules to follow when calculating a formula:

1) Write the **mass** or **percentage mass** of each element.
2) Find the number of **moles** of each substance by dividing by the atomic or molecular mass.
3) Divide all answers by the **smallest** answer.
4) If required: multiply to make up to **whole numbers**.
5) Use the **ratio** of atoms to write the formula (this gives the empirical formula).

EXAMPLE: Find the formula of an oxide of aluminium formed from 9.00 g aluminium and 8.00 g oxygen.

1) First write down the mass of each substance:
   Al: 9.00 g O: 8.00 g
2) Divide the mass by the atomic masses to find the number of moles of each substance:
   Al: 9.00 ÷ 27.0 = 0.333 moles O: 8.00 ÷ 16.0 = 0.500 moles
3) Divide by the smallest number, which is 0.333:
   Al: 0.333 ÷ 0.333 = 1.00 O: 0.5 ÷ 0.333 = 1.50
4) Multiply by 2 to give whole numbers:
   Al: 1.00 × 2 = 2 O: 1.50 × 2 = 3
5) The ratio of Al : O is **2:3**.
   The empirical formula is **Al_2O_3**.

Roman empirical formula — 1 Caesar, 3 gladiators & 8 straight roads...

1) Find the empirical formulae of the following oxides:
   a) An oxide containing 12.9 g of lead to every 1.00 g of oxygen.
   b) An oxide containing 2.33 g of iron to every 1.00 g of oxygen.
   (Relative atomic mass values: Pb = 207.2, O = 16.0, Fe = 55.8)
2) Calculate the empirical formula of the carboxylic acid that is comprised of 4.30% hydrogen, 26.1% carbon and 69.6% oxygen.
   (Relative atomic mass values: H = 1.0, C = 12.0, O = 16.0)

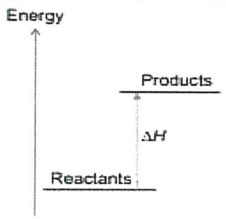
Endothermic and Exothermic Reactions

In an **exothermic** reaction, **heat** energy is **given out** (the room temperature rises).
In an **endothermic** reaction, **heat** energy is **taken** from the surroundings (the room temperature drops).

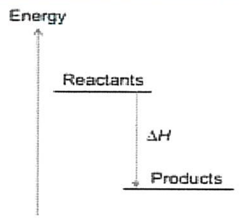
Making and Breaking Bonds

1) It takes energy to **break bonds**. When two atoms joined by a bond are **separated**, the energy required to do this must be provided from the surroundings.
2) However, energy is **released** when bonds are made. When two atoms become **joined together** by forming a bond, energy is **released** to the surroundings.
3) In a reaction, if more energy is taken in to break bonds than is given out when bonds are made, the process is **endothermic** — it will take in heat energy. The overall **enthalpy change** of the reaction (ΔH) is **positive**.
4) But, if more energy is given out when bonds are made than is taken in when bonds are broken, the process is **exothermic** — it will give out heat energy. The overall **enthalpy change** of an exothermic reaction (ΔH) is **negative**.

Reactions can be Represented by Energy Level Diagrams



In an **endothermic** reaction, the reactants **take in** energy from the surroundings. The products therefore have **more energy** than the reactants, and ΔH is **positive**.



In an **exothermic** reaction, the reactants **release** energy to the surroundings. The products therefore have **less energy** than the reactants and ΔH is **negative**.

After that I think I need a cup of tea. It'll help improve my energy level...

1) Are the following reactions exothermic or endothermic?
   a) burning coal
   b) sodium hydrogencarbonate + hydrochloric acid (temperature drops)
   c) acid + hydroxide (gets hotter)
   d) methane + steam (cools as they react)
2) a) Draw an energy level diagram for the following reaction:
   $C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$ $\quad \Delta H = -2809$ kJ mol^{-1}
   You should label the products, reactants and enthalpy change on your diagram.
   b) Is the reaction in part a) endothermic or exothermic?

Bond Energy

Average Bond Energy

Bonds between **different atoms** require different amounts of **energy** to break them. When the **same two atoms** bond in the same way, the amount of energy needed is always about the same. The average bond energy values for some common bonds are given below:

C–H 413	C–O 360	C=C 612
O=O 498	H–H 436	C=O 743
C–C 348	O–H 463	

All these values are in kJ mol^{-1}.

The values tell you that:

e.g. It takes 413 kJ of energy to break 1 mole of C–H bonds.
It takes 463 × 2 = 926 kJ to break 1 mole of water (which has 2 O–H bonds per molecule) into oxygen and hydrogen atoms.
743 × 2 = 1486 kJ are released when 1 mole of CO_2 (which has 2 C=O bonds) forms.

Calculating the Change in Energy

When a reaction takes place, the change in energy is simply:

sum of energy required to break old bonds – sum of energy released by new bonds formed

EXAMPLE: Calculate the energy change involved when 1 mole of methane burns in oxygen.

The equation for the reaction is: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

This tells you that 1 mole of methane reacts with 2 moles of oxygen to form 1 mole of carbon dioxide and 2 moles of water.

Step 1: Calculate the energy required to break all of the bonds between the reactant atoms:
4 C–H bonds = 4 × 413 = 1652 kJ
2 O=O bonds = 2 × 498 = 996 kJ
Total = 2648 kJ

Step 2: Calculate the energy released by all the new bonds formed in the products:
2 C=O bonds = 2 × 743 = 1486 kJ
4 O–H bonds = 4 × 463 = 1852 kJ
Total = 3338 kJ

Step 3: Combine the two values to give the overall value for the energy change:
The overall energy change is: 2648 – 3338 = **–690 kJ mol^{-1}**.

The negative sign shows that energy is being released to the surroundings, indicating that this is an **exothermic** reaction. This is expected, since this is a combustion reaction.

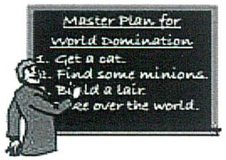
Ian Fleming was like an exothermic reaction — he made lots of Bonds...

1) Calculate the energy change of the following reactions:
   (Use the values for the average bond energies given at the top of the page).
   a) burning 1 mole of propane $\quad C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$
   b) burning 1 mole of ethanol $\quad C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$
   c) hydrogenation of 1 mole of ethene $\quad C_2H_4 + H_2 \rightarrow C_2H_6$

Planning Experiments

Make Sure You Plan Your Experiment Carefully

To get accurate and precise results from your experiments, you first need to plan them carefully...



1) Work out the **aim** of the experiment.
2) Identify the **variables** (see below).
3) Decide what **data** to collect.
4) Decide the right **equipment** to use.
5) Plan how to reduce any **risks** in your experiment.
6) Write out a **detailed method**.
7) Carry out **tests** to address the aim of your experiment.

You Need to Control All the Variables

A **variable** is a quantity that might **change** during an experiment, for example temperature. There are two types of variables to know about when carrying out an experiment:

- The **independent variable** is the quantity that you **change**.
- The **dependent variable** is the thing that you measure.

When you plan an experiment you need to work out how you will **control** the variables so that the only one that changes is the one you're investigating — all the others are kept **constant**.

EXAMPLE: Measuring the effect of surface area on reaction rate.

In this experiment, the **independent variable** is the **surface area**, and the **dependent variable** is the **rate** of reaction.

Everything else, such as temperature and concentration, has to stay exactly the same between different experiments. Surface area is the only variable that you change.

Choose the Right Equipment

You need to think carefully about selecting the right **equipment** for your experiment...

1) The equipment has to be **appropriate** for the specific experiment — for example, in an experiment where you're collecting a **gas** the equipment you use needs to be properly **sealed** so that the gas can't **escape**.
2) The equipment needs to be the right **size**.
3) The equipment needs to be the right level of **sensitivity** — for example, if you want to measure out 4.2 g of a compound, you'll need a balance that measures to at least the nearest 0.1 g, not the nearest gram.

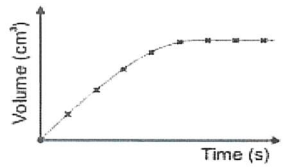
Reduce Risk — and play poker instead...

1) A student is measuring the effect of temperature on the time taken for a lump of magnesium to react completely in a sample of concentrated hydrochloric acid.
   a) What is the dependent variable in the student's experiment?
   b) Name two variables that the student should control to make the experiment a fair test.

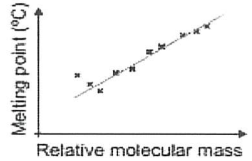
Presenting and Interpreting Data

You Can Represent Your Data in a Table or on a Graph

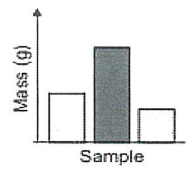
When you do an experiment, it's a good idea to set up a table to **record** your **results** in. Make sure you **include** enough **rows** and **columns** to **record all of the data** you need. Tables are good for **recording** data, but it can be easier to interpret your results if you **plot** them on a **graph**. Depending on the **type** of experiment, the **graph** you plot will vary:



Line graphs show how two sets of data are related.



Scatter plots show **trends** in data. Don't join all the points — just draw a **line of best fit**.



If one of your sets of data can be split into **groups**, draw a **bar graph**.

Repeating an Experiment Makes Your Results More Reliable

1) If you **repeat** an experiment, your results will usually **differ slightly** each time you do it. You can use the **mean** (or average) of the measurements to represent all these values. The more times you repeat the experiment the **more reliable** the average will be. To find the mean:

Add together all the data values then **divide** by the total number of values in the sample.

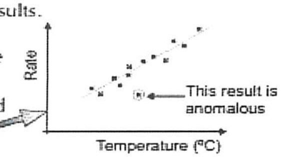
EXAMPLE: Calculate the mean result for the volume of hydrogen gas produced after 30 seconds in the reaction between hydrochloric acid and magnesium.

Run 1	Run 2	Run 3
23 cm^3	22 cm^3	25 cm^3

There are **three** values in this sample, so to find the mean result, just add together the results and divide by three:

$(23 + 22 + 25) \div 3 = \mathbf{23.3\ cm^3}$

2) Repeating experiments also lets you spot any **weird results** that stick out like a hedgehog in a tea cup. These are called **anomalous** results. For example — if one of the results above was only 5 cm^3, then something probably went wrong. You should **ignore** the anomalous result when you calculate the mean.
3) Anomalous results are really easy to spot on **scatter plots** and **line graphs** as they sit miles away from the line of best fit.



I was hoping for a nice result, but it ended up being mean...

1) Kay measured the volume of gas given off in a reaction. Her results were 22.0 cm^3, 23.0 cm^3, 22.0 cm^3, 19.0 cm^3 and 24.0 cm^3. Identify any anomalous results and calculate the mean.

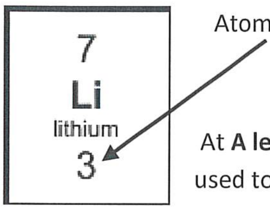
Section 2: Pre-Knowledge Topics

Chemistry topic 1 – Electronic structure, how electrons are arranged around the nucleus

A periodic table can give you the proton / atomic number of an element, this also tells you how many electrons are in the ***atom***.

You will have used the rule of electrons shell filling, where:

The first shell holds up to 2 electrons, the second up to 8, the third up to 8 and the fourth up to 18 (or you may have been told 8).



Atomic number =3, electrons = 3, arrangement 2 in the first shell and 1 in the second or

Li = 2,1

At **A level** you will learn that the electron structure is more complex than this, and can be used to explain a lot of the chemical properties of elements.

The 'shells' can be broken down into 'orbitals', which are given letters:'s' orbitals, 'p' orbitals and 'd' orbitals.

You can read about orbitals here:

http://bit.ly/pixlchem1

http://www.chemguide.co.uk/atoms/properties/atomorbs.html#top

Now that you are familiar with s, p and d orbitals try these problems, write your answer in the format:

$1s^2$, $2s^2$, $2p^6$ etc.

Q1.1 Write out the electron configuration of:

a) Ca b) Al c) S d) Cl e) Ar f) Fe g) V h) Ni i) Cu j) Zn k) As

Q1.2 Extension question, can you write out the electron arrangement of the following ***ions***:

a) K^+ b) O^{2-} c) Zn^{2+} d) V^{5+} e) Co^{2+}

Chemistry topic 3 – Isotopes and mass

You will remember that an isotopes are elements that have differing numbers of neutrons. Hydrogen has 3 isotopes; H_1^1 H_1^2 H_1^3

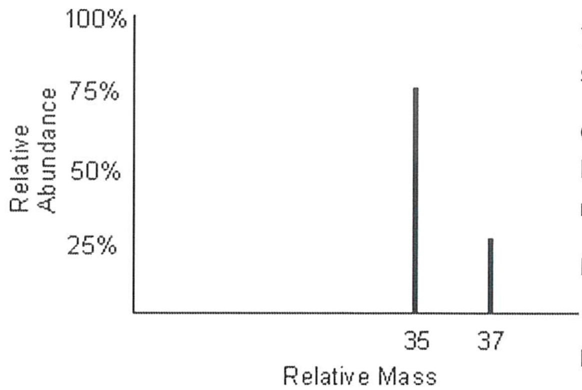
Isotopes occur naturally, so in a sample of an element you will have a mixture of these isotopes. We can accurately measure the amount of an isotope using a **mass spectrometer**. You will need to understand what a mass spectrometer is and how it works at A level. You can read about a mass spectrometer here:

http://bit.ly/pixlchem3
http://www.kore.co.uk/tutorial.htm
http://bit.ly/pixlchem4
http://filestore.aqa.org.uk/resources/chemistry/AQA-7404-7405-TN-MASS-SPECTROMETRY.PDF

Q3.1 What must happen to the atoms before they are accelerated in the mass spectrometer?

Q3.2 Explain why the different isotopes travel at different speeds in a mass spectrometer.

A mass spectrum for the element chlorine will give a spectrum like this:



75% of the sample consist of chlorine-35, and 25% of the sample is chlorine-37.

Given a sample of naturally occurring chlorine ¾ of it will be Cl-35 and ¼ of it is Cl-37. We can calculate what the **mean** mass of the sample will be:

$$\text{Mean mass} = \frac{75}{100} \times 35 + \frac{25}{100} \times 37 = 35.5$$

If you look at a periodic table this is why chlorine has an atomic mass of 35.5.

http://www.avogadro.co.uk/definitions/ar.htm

An A level periodic table has the masses of elements recorded much more accurately than at GCSE. Most elements have isotopes and these have been recorded using mass spectrometers.

GCSE

11 B boron 5	12 C carbon 6	14 N nitrogen 7	16 O oxygen 8	19 F fluorine 9
27 Al aluminium 13	28 Si silicon 14	31 P phosphorus 15	32 S sulfur 16	35.5 Cl chlorine 17

A level

10.8 B 5 boron	12.0 C 6 carbon	14.0 N 7 nitrogen	16.0 O 8 oxygen	19.0 F 9 fluorine
27.0 Al 13 aluminium	28.1 Si 14 silicon	31.0 P 15 phosphorus	32.1 S 16 sulphur	35.5 Cl 17 chlorine

Given the percentage of each isotope you can calculate the mean mass which is the accurate atomic mass for that element.

Q3.3 Use the percentages of each isotope to calculate the accurate atomic mass of the following elements.

a) Antimony has 2 isotopes: Sb-121 57.25% and Sb-123 42.75%

Q5.1 Balance the following equations

a. $H_2 + 0_2 \rightarrow H_20$

b. $S_8 + 02 \rightarrow S0_3$

c. $HgO \rightarrow Hg + 0_2$

d. $Zn + HCl \rightarrow ZnCl_2 + H_2$

e. $Na + H_20 \rightarrow NaOH + H_2$

f. $C_{10}H_{16} + Cl_2 \rightarrow C + HCl$

g. $Fe + 0_2 \rightarrow Fe_20_3$

h. $C_6H_{12}0_6 + 0_2 \rightarrow C0_2 + H_20$

i. $Fe_20_3 + H_2 \rightarrow Fe + H_20$

j. $Al + FeO \rightarrow Al_2O_3 + Fe$

Chemistry topic 6 – Measuring chemicals – the mole

From this point on you need to be using an A level periodic table, not a GCSE one you can view one here:

http://bit.ly/pixlpertab

https://secondaryscience4all.files.wordpress.com/2014/08/filestore_aqa_org_uk_subjects_aqa-2420-w-trb-ptds_pdf.png

Now that we have our chemical equations balanced, we need to be able to use them in order to work out masses of chemicals we need or we can produce.

The ***mole*** is the chemists equivalent of a dozen, atoms are so small that we cannot count them out individually, we weigh out chemicals.

For example: magnesium + sulfur → magnesium sulfide

$$Mg + S \rightarrow MgS$$

We can see that one atom of magnesium will react with one atom of sulfur, if we had to weigh out the atoms we need to know how heavy each atom is.

From the periodic table: Mg = 24.3 and S = 32.1

If I weigh out exactly 24.3g of magnesium this will be 1 mole of magnesium, if we counted how many atoms were present in this mass it would be a huge number (6.02×10^{23}!!!!), if I weigh out 32.1g of sulfur then I would have 1 mole of sulfur atoms.

So 24.3g of Mg will react precisely with 32.1g of sulfur, and will make 56.4g of magnesium sulfide.

Here is a comprehensive page on measuring moles, there are a number of descriptions, videos and practice problems.

Chemistry topic 8 – Titrations

One key skill in A level chemistry is the ability to carry out accurate titrations, you may well have carried out a titration at GCSE, at A level you will have to carry them out very precisely **and** be able to describe in detail how to carry out a titration - there will be questions on the exam paper about how to carry out practical procedures.

You can read about how to carry out a titration here, the next page in the series (page 5) describes how to work out the concentration of the unknown.

http://bit.ly/pixlchem11

http://www.bbc.co.uk/schools/gcsebitesize/science/triple_aqa/further_analysis/analysing_substances/revision/4/

Remember for any titration calculation you need to have a balanced symbol equation; this will tell you the ratio in which the chemicals react.

E.g. a titration of an unknown sample of sulfuric acid with sodium hydroxide.

A 25.00cm^3 sample of the unknown sulfuric acid was titrated with 0.100moldm^{-3} sodium hydroxide and required exactly 27.40cm^3 for neutralisation. What is the concentration of the sulfuric acid?

Step 1: the equation $2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$

Step 2; the ratios 2 : 1

Step 3: how many moles of sodium hydroxide 27.40cm^3 = 0.0274dm^3

number of moles = c x v = 0.100 x 0.0274 = 0.00274 moles

step 4: Using the ratio, how many moles of sulfuric acid

for every 2 NaOH there are 1 H_2SO_4 so, we must have 0.00274/2 =0.00137 moles of H_2SO_4

Step 5: Calculate concentration. concentration = moles/volume ←in dm^3 = 0.00137/0.025 = ***0.0548 moldm^{-3}***

Here are some additional problems, which are harder, ignore the questions about colour changes of indicators.

http://bit.ly/pixlchem12

http://www.docbrown.info/page06/Mtestsnotes/ExtraVolCalcs1.htm

Use the steps on the last page to help you

Q8.1 A solution of barium nitrate will react with a solution of sodium sulfate to produce a precipitate of barium sulfate.

$Ba(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaNO_3(aq)$

What volume of 0.25moldm^{-3}sodium sulfate solution would be needed to precipitate all of the barium from 12.5cm^3 of 0.15 moldm^{-3} barium nitrate?

Chemistry topic 10 – Acids, bases, pH

At GCSE you will know that an acid can dissolve in water to produce H^+ ions, at A level you will need a greater understanding of what an acid or a base is.

Read the following page and answer the questions

http://bit.ly/pixlchem15

http://www.chemguide.co.uk/physical/acidbaseeqia/theories.html#top

Q10.1 What is your new definition of what an acid is?

Q10.2 How does ammonia (NH_3) act as a base?

http://bit.ly/pixlchem16

http://www.chemguide.co.uk/physical/acidbaseeqia/acids.html#top

Q10.3 Ethanoic acid (vinegar) is a weak acid, what does this mean?

Q10.4 What is the pH of a solution of 0.01 moldm^{-3} of the strong acid, hydrochloric acid?

Videos to watch online

Rough science – the Open University – 34 episodes available

Real scientists are 'stranded' on an island and are given scientific problems to solve using only what they can find on the island.

Great fun if you like to see how science is used in solving problems.

There are six series in total

http://bit.ly/pixlchemvid1a

http://www.dailymotion.com/playlist/x2igjq_Rough-Science_rough-science-full-series/1#video=xxw6pr

or

http://bit.ly/pixlchemvid1b

https://www.youtube.com/watch?v=lUoDWAt259I

A thread of quicksilver – The Open University

A brilliant history of the most mysterious of elements – mercury. This program shows you how a single substance led to empires and war, as well as showing you come of the cooler properties of mercury.

http://bit.ly/pixlchemvid2

https://www.youtube.com/watch?v=t46lvTxHHTA

10 weird and wonderful chemical reactions

10 good demonstration reactions, can you work out the chemistry of .... any... of them?

http://bit.ly/pixlchemvid3

https://www.youtube.com/watch?v=0Bt6RPP2ANI

Chemistry in the Movies

Dantes Peak 1997: Volcano disaster movie.

Use the link to look at the Science of acids and how this links to the movie.
http://www.open.edu/openlearn/science-maths-technology/science/chemistry/dantes-peak

http://www.flickclip.com/flicks/dantespeak1.html

http://www.flickclip.com/flicks/dantespeak5.html

Fantastic 4 2005 &2015: Superhero movie

Michio Kaku explains the "real" science behind fantastic four http://nerdist.com/michio-kaku-explains-the-real-science-behind-fantastic-four/

http://www.flickclip.com/flicks/fantastic4.html

Task 5: ITO and the future of touch screen devices

ITO – indium tin oxide is the main component of touch screen in phones and tablets. The element indium is a rare element and we are rapidly running out of it. Chemists are desperately trying to find a more readily available replacement for it. What advances have chemists made in finding a replacement for it?

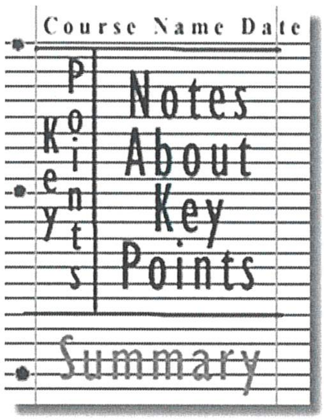


Figure 1: http://coe.jmu.edu/learningtoolbox/images/noteb4.gif

Places to visit

1. Go outdoors!
   Have you actually spent any time observing the geology of the area you live in? What rocks or minerals are found in your area? Does your area have a history of extracting minerals? If so what were they, what were they used for, how did they obtain them? Are there any working or remains of mineral extraction industries?

2. Are there any chemical or chemistry based businesses in your area? A big ask, but one that could be really beneficial to you, write them a letter explaining that you are taking A level chemistry and you want to see how chemistry is used in industry and you would like to visit / have some work experience. You never know this could lead to great things!!!!

3. You could also try writing to / searching for your nearest university to see if they are running any summer schools for chemistry – they are usually free and give you the opportunity to experience the laboratories in a university.

4. Science museums.
   You could visit your nearest science museum. They often have special exhibitions that may be of interest to you.
   https://en.wikipedia.org/wiki/List_of_science_museums#United_Kingdom