Chemistry A Level – Preparatory Work 2022

Preparing to begin A-level Chemistry at Hereford Sixth Form College

INTRODUCTION

These pages aim to make sure that your knowledge and understanding from GCSE have put you in a good place to begin the A-level work next year. They focus on topics from GCSE and key skills (especially maths skills) which will continue to be really important as you move on to a broader and deeper understanding at A-level.

The work is in 2 sections:

Section A focusses on maths skills and includes questions that you need to **answer on paper and have ready to hand in when you come along to your first Chemistry lesson**.

Section B covers a number of topics from Chemistry which should (mainly, at least) be familiar from your GCSE work. This section also includes lots of questions but these are not for handing in: the idea in section B is for you to have a go at the questions and check the answers yourself: you won't be handing this in to us.

Don't worry if you can't do some of the questions or if we seem to be assuming you know about things you've actually never heard of! We will go through all this stuff during the first year of the course. Just try and pick up as much as you can from this material now, as that will make your learning easier when you do the course next year.

CONTENTS

Section A Maths Skills for Chemistry

There is a lot of maths involved in chemistry at A-level. The examples below are not chemical ones but they illustrate the kind of manipulations you will have to be able to carry out. **If you don't yet have these skills, please be aware that it's going to require a lot of extra work to develop them**.

- At the end of this section, there are a number of questions for you to have a go at. For this section only, you should **answer on paper and bring your answers in to hand to your teacher when you have your first lesson**.
- It is very important to show your working fully in answering these questions.

Rearranging Equations

This is a very important thing to be able to do, as it comes up a lot in chemical problems.

If your equation just involves terms multiplied or divided together (no extra terms added or subtracted) you can rearrange it by dividing both sides or multiplying both sides by the same thing:

e.g. you know that velocity = distance / time, which we can write as $v = s/t$

So what does t equal? To do this,

- first multiply both sides by t, so that the t cancels out on the right. This gives you: vt = s
- then divide both sides by v, so that v cancels on the left. This gives you: t = s/v

You can do this in a "visual" way, without thinking explicitly about multiplying and dividing, if you prefer. The key is to remember that terms slide diagonally from one side to the other:

- starting with v = s/t again. Slide the t diagonally from the bottom on the right to the top on the left, to get: vt = s
- then slide the v diagonally from the top on the left to bottom on the right, to get: t = s/v

Significant Figures

When we measure something in chemistry – like mass or temperature – we can't ever measure it exactly because we'll be limited by the precision of the device we're using (the balance or the thermometer). So when we report these measurements and use them in calculations, we use the idea of significant figures to indicate how precisely we know something.

For example, here are theresults of an experiment to measure what % of the mass of a rock sample is chromium:

- Mass of rock (measured on a balance measuring grams to 2 decimal places) = 5.30g
- Mass of chromium in the rock (measured by a chemical experiment) = 1.07g

Notice, first, that we wrote 5.30g and not just 5.3g. This is to show that we know the mass to the nearest 0.01g. So it is probably not exactly 5.30 but it's closer to 5.30 than to 5.29 or 5.31. We have given the mass to 3 significant figures.

So we can calculate the % by mass of chromium

= (mass of chromium / total mass) x 100 = (1.07 / 5.30) x 100

= 20.1886792 % (says your calculator)

We can't really know the number this precisely because we are limited by precision of the numbers we fed in. So, since the mass values we used were known to 3 significant figures, we should round our answer to the same precision, to get:

% by mass of chromium = 20.2% (to 3 SF)

Notice that the unrounded number had a 1 just before the point where we rounded off, but it got rounded up to 2 because the number following was greater than or equal to 5. You must make sure you know about rounding like this.

Remember too that for numbers less than one, the first significant figure is the first non-zero digit.

e.g. if you need to round 0.007438993 to 2 significant figures, you get 0.0074 (because that 7 was the first significant figure).

Standard Form

It is very common in science to give numbers – especially very small or very big ones – in standard form:

e.g. 52171 can be written as 5.2171×10^4 and 0.0000337 is 3.37×10^{-5}

Simple rules for converting to standard form:

For a big number, count how many places to the left the decimal point has to move to the left, to leave just one digit in front of it. If it has to move x places, the number will be … $\times 10^x$

e.g. for 52171, the point moves 4 places left to leave just 5 in front, so we get 5.2171×10^4

For a small number, count how many places to the right the decimal point must move to the right, to leave just one non-zero digit in front of it. If it has to move x places, the number will be … $\times 10^{-x}$.

e.g. for 0.0000337, the point moves 5 places right to leave just 3 in front,
so we get 3.37×10^{-5}

Now here are the questions: they aren't specifically related to Chemistry but use the key skills that you'll need to deal with A-level chemical calculations. Please answer on paper and have ready to hand to your teacher in your first lesson:

1. Speed = distance / time. Use this relationship to find how far a cormorant can travel in 20 minutes , if it is flying at 15.8 km per hour.

2. If a Labrador grows at a steady rate, we can write an expression for its mass:

mass (m) = birth mass (m_0) + growth rate (R) x time (t) i.e. $\boxed{m = m_0 + Rt}$

If a Labrador weighs 17.6 kg after 180 days and its growth rate is 0.0950 kg per day, what was its birth mass?

3. If $AB^2/C = K$ write an expression for B.

4. An average quince weighs 47.1g. A standard crate full of quinces weighs 3920g. An empty crate weighs 450g. What is the average number of quinces per crate? Give your answer:

(i) To 3 significant figures (ii) To the nearest whole quince (iii) To 2 decimal places

5. Put the following numbers in standard form: 356,000 0.00714

6. Express the following as conventional numbers: 1.04×10^5 2.22×10^{-4}

7. What is 247g in kg?

8. What is 4.45g per cm^3, in mg per cm^3?

9. Calculate: $28 \div (-4)$ $(-50) \times (-6)$ $(-10) - ((-20) / 10)$

10. $A = B - CD$. Find the value of C when A = -10, B = -80, D= +2

11. If I have eaten 30% of my Stinking Bishop cheese, and I now have 140g left, what mass of Stinking Bishop did I have to start with?

Section B Important Ideas for Chemistry

- **This section covers a number of important topics that you may already have learned something about at GCSE. Don't worry if you haven't covered any of these things: we will be going through them all in the early stages of your A-level course.**
- **Each topic has a few questions at the end for you to have a think about. <u>These are NOT for handing in to your teacher</u>. Instead, you can get worked answers to them from us, if you send an email to pae@hereford.ac.uk. These answers will let you check what you've done but also help you see what to do if you get stuck.**

1. The Periodic Table

The Periodic Table is probably the single thing you will use more often than anything else as you progress with your chemistry. The good news is that you don't have to learn it by heart – you will have a data sheet in lessons, and in your exams, which includes the table and key numbers for each element. The more you use the table, the quicker you'll get at finding particular elements, and seeing the relationships between them. The table, as it appears on the A-level data sheet is included on the next page.

Key features of the periodic table include:

- It is organised with atoms placed **in order of increasing atomic number**.
- Atoms of elements in the same **period** (a horizontal row) have at least one electron in the shell corresponding to their period number. For example, you'll find calcium (atomic number 20) in the fourth row down (period 3), so you know that a calcium atom has at least one electron in the fourth shell.
- Elements in the same **group** (vertical column) all have the same number of electrons in their outermost shell. This number is the same as the group number shown in brackets across the top of the table (eg nitrogen is in group 5 so it has 5 outer shell electrons). The consequence is that elements in the same group have similar physical and chemical properties, although reactivity changes as you descend the group.
- At A level you will learn about the way electrons behave in atoms and this will give you an understanding of why the periodic table has the shape it does.

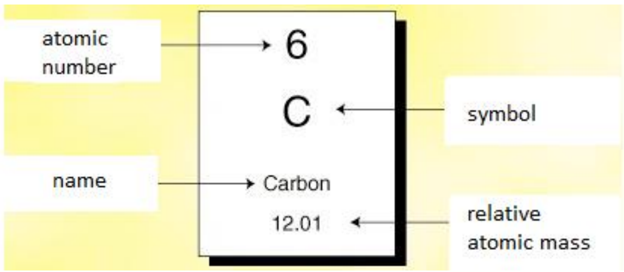
The Periodic Table of the Elements

Key
atomic number
Symbol
name
relative atomic mass

(1)	(2)											(3)	(4)	(5)	(6)	(7)	(0)
1	**2**	**3**	**4**	**5**	**6**	**7**	**8**	**9**	**10**	**11**	**12**	**13**	**14**	**15**	**16**	**17**	**18**
1 **H** hydrogen 1.0																	2 **He** helium 4.0
3 **Li** lithium 6.9	4 **Be** beryllium 9.0											5 **B** boron 10.8	6 **C** carbon 12.0	7 **N** nitrogen 14.0	8 **O** oxygen 16.0	9 **F** fluorine 19.0	10 **Ne** neon 20.2
11 **Na** sodium 23.0	12 **Mg** magnesium 24.3											13 **Al** aluminium 27.0	14 **Si** silicon 28.1	15 **P** phosphorus 31.0	16 **S** sulfur 32.1	17 **Cl** chlorine 35.5	18 **Ar** argon 39.9
19 **K** potassium 39.1	20 **Ca** calcium 40.1	21 **Sc** scandium 45.0	22 **Ti** titanium 47.9	23 **V** vanadium 50.9	24 **Cr** chromium 52.0	25 **Mn** manganese 54.9	26 **Fe** iron 55.8	27 **Co** cobalt 58.9	28 **Ni** nickel 58.7	29 **Cu** copper 63.5	30 **Zn** zinc 65.4	31 **Ga** gallium 69.7	32 **Ge** germanium 72.6	33 **As** arsenic 74.9	34 **Se** selenium 79.0	35 **Br** bromine 79.9	36 **Kr** krypton 83.8
37 **Rb** rubidium 85.5	38 **Sr** strontium 87.6	39 **Y** yttrium 88.9	40 **Zr** zirconium 91.2	41 **Nb** niobium 92.9	42 **Mo** molybdenum 95.9	43 **Tc** technetium	44 **Ru** ruthenium 101.1	45 **Rh** rhodium 102.9	46 **Pd** palladium 106.4	47 **Ag** silver 107.9	48 **Cd** cadmium 112.4	49 **In** indium 114.8	50 **Sn** tin 118.7	51 **Sb** antimony 121.8	52 **Te** tellurium 127.6	53 **I** iodine 126.9	54 **Xe** xenon 131.3
55 **Cs** caesium 132.9	56 **Ba** barium 137.3	57–71 lanthanoids	72 **Hf** hafnium 178.5	73 **Ta** tantalum 180.9	74 **W** tungsten 183.8	75 **Re** rhenium 186.2	76 **Os** osmium 190.2	77 **Ir** iridium 192.2	78 **Pt** platinum 195.1	79 **Au** gold 197.0	80 **Hg** mercury 200.6	81 **Tl** thallium 204.4	82 **Pb** lead 207.2	83 **Bi** bismuth 209.0	84 **Po** polonium	85 **At** astatine	86 **Rn** radon
87 **Fr** francium	88 **Ra** radium	89–103 actinoids	104 **Rf** rutherfordium	105 **Db** dubnium	106 **Sg** seaborgium	107 **Bh** bohrium	108 **Hs** hassium	109 **Mt** meitnerium	110 **Ds** darmstadtium	111 **Rg** roentgenium	112 **Cn** copernicium		114 **Fl** flerovium		116 **Lv** livermorium		

57 **La** lanthanum 138.9	58 **Ce** cerium 140.1	59 **Pr** praseodymium 140.9	60 **Nd** neodymium 144.2	61 **Pm** promethium 144.9	62 **Sm** samarium 150.4	63 **Eu** europium 152.0	64 **Gd** gadolinium 157.2	65 **Tb** terbium 158.9	66 **Dy** dysprosium 162.5	67 **Ho** holmium 164.9	68 **Er** erbium 167.3	69 **Tm** thulium 168.9	70 **Yb** ytterbium 173.0	71 **Lu** lutetium 175.0
89 **Ac** actinium	90 **Th** thorium 232.0	91 **Pa** protactinium	92 **U** uranium 238.1	93 **Np** neptunium	94 **Pu** plutonium	95 **Am** americium	96 **Cm** curium	97 **Bk** berkelium	98 **Cf** californium	99 **Es** einsteinium	100 **Fm** fermium	101 **Md** mendelevium	102 **No** nobelium	103 **Lr** lawrencium

Each element has its own square on the table which includes three key pieces of information:



The **atomic number** is the number of protons in the nucleus of an atom of that element. This number does not change in chemical reactions.

The **relative atomic mass** is the average mass of an atom of an element. More about this later.

Questions 1

1. **You've probably never met any rubidium (atomic number 37) but see if you can predict what sort of properties it might have, from its position in the table.**

2. **What do the following have in common, in terms of their electronic structures?**
   i. **Magnesium and sulfur**
   ii. **Boron and gallium**

2. Atoms and Ions

Atoms are made out of protons and neutrons (which are packed together in the nucleus) and electrons which occupy most of the space in the atom, outside the nucleus. You need to know the basic properties of these particles:

PARTICLE	RELATIVE MASS	CHARGE
Proton	1	+1
Neutron	1	0
Electron	approx. 1/2000	-1

A few key points:

- The number of electrons in a neutral atom is equal to the number of protons. Atoms have no charge because there are equal numbers of positive and negative charges, and protons and electrons have equal and opposite charges.
- The sum of the number of protons and the number of neutrons in an atom is called the **mass number**. This is written top left of the symbol

  e.g. an atom with 16 protons is a sulfur atom (atomic number 16). If this atom has 17 neutrons, its mass number is 33, so it can be written as ^{33}S.

- A simple subtraction sum of mass number – atomic number will determine the number of neutrons.
- The chemical properties of atoms are dominated by the electrostatic forces between the positively charged protons and negatively charged electrons. So neutrons, which have no charge, have only a small effect on chemical properties. For this reason you can get atoms of the same element, with virtually the same chemical properties but different numbers of neutrons. These are called **isotopes**.

Ions are particles formed when atoms lose or gain electrons (note that the number of protons and neutrons never changes in a *chemical* process). Since electrons are negatively charged, we can see that:

- Positive ions (also called **cations**) are formed when electrons are lost from an atom. e.g. Mg^{2+} is a magnesium ion, formed when a Mg atom loses two electrons.
- Negative ions (also called **anions**) are formed when extra electrons are gained by an atom. e.g. Cl^- is a chlori**d**e ion, formed when a chlori**n**e atom gains an electron. (Notice – and try to get in the habit of following this pattern – the change in the spelling when an atom becomes a negative ion).

Questions 2

1. **Write down the symbols for these atoms, including the mass number (follow the example of ^{33}S, above):**

	Protons	Neutrons	Electrons
(i)	1	1	1
(ii)	26	30	26
(iii)	92	143	92

2. **Write down how many protons, neutrons and electrons are in each of the following atoms:**
   (i) ^{37}Cl (ii) ^{94}Mo

3. **Write the name and the symbol for what you get when:**
   (i) An oxygen atom gains two electrons.
   (ii) An aluminium atoms loses three electrons

4. **Write the full symbol, including the mass number and the charge, for particles with following compositions:**

	Protons	Neutrons	Electrons
(i)	1	2	0
(ii)	35	46	36
(iii)	111	172	108

3. **Chemical Calculations**

The **mass number** refers to a specific isotope. However – with rare exceptions – we are usually working in chemistry not with individual atoms but with enormous numbers of atoms (e.g. a mugful of neon gas contains around 10 000 000 000 000 000 000 000 atoms). These will generally include a mixture of isotopes (our mugful of neon is a mixture of ^{20}Ne, ^{21}Ne and ^{22}Ne atoms, for example).

So in most practical situations in chemistry, what we care about is the average mass of an atom of an element in the mixture we are working with, rather than the mass of any one individual isotope. Fortunately the relative amounts of each isotope are fairly constant everywhere on Earth (although the small differences that exist are interesting and sometimes useful to measure). This means that once we've measured the average atomic mass, it will be valid for pretty much any calculation.

So we define **relative atomic mass**:

Relative atomic mass (A_r) is the weighted mean mass of an atom of an element relative to 1/12 of the mass of an atom of ^{12}C.

Expressing the mass of an atom on a relative scale like this (using 1/12 of a ^{12}C atom as our reference point) gives us numbers that are much easier to handle than tiny actual masses in grams would be).

The idea of a "weighted mean" is that an isotope contributes more towards the average, the more abundant it is on Earth. This leads to the following recipe for calculating A_r:

- For each isotope multiply the mass number x the relative amount of the isotope
- Add these terms together
- Divide by the sum of the relative amounts.

For example:

In a sample of chlorine, 75% of atoms are ^{35}Cl and 25% are ^{37}Cl.

The average mass is therefore:

$$A_{\mathrm{r}} = \frac{(35 \times 75) + (37 \times 75)}{100} \quad = \quad 35.5$$

Make sure you don't confuse mass numbers with relative atomic masses:

- Mass numbers refer to single isotopes and they are always whole numbers. Mass numbers should be used at the top left of an atom symbol to identify the isotope.

- Relative atomic mass refers to the isotopes that occur in nature, which usually means a mixture. They are often not whole numbers. These are the numbers you find for each element at the bottom of its square in the periodic table.

The Mole is a quantity that's used in chemistry to define an amount of a substance. The definition of a mole looks a bit strange:

One mole is the amount of substance containing the Avogadro number of atoms, molecules or groups of ions.

Avogadro's number is = 6.02 x 10^{23} mol^{-1} (ie 602 000 000 000 000 000 000 000)

There is a very good reason for using this weird looking definition: it allows us to take relative atomic masses (measured relative to 1/12 of a single atom) and apply them easily on the big scale (measured in g) that we normally work on:

Avogadro's number (written N_A for short) is special because if you have that number of atoms of ^{12}C, it has a mass of exactly 12g. Similarly, since for example, chlorine has a relative atomic mass of 35.5, and Avogadro's number of chlorine atoms has a mass of 35.5g.

One mole of any substance contains the same number (N_A) of particles. So we can make this incredibly important general statement:

One mole of atoms of any element has a mass (in grams) that is numerically equal to the relative atomic mass.

For example, Iron has a relative atomic mass of 55.8, so 1 mole of iron has a mass of 55.8g

Relative formula mass is a simple generalisation of this idea, to let us work with substances with more than one atom in their formula by simply adding up the individual atomic masses to get the relative formula mass (M_r).

e.g. Oxygen exists as molecules made up of pairs of oxygen atoms, so its formula is O_2. The relative formula mass is therefore: $M_r = 2 \times A_r = 2 \times 16.0 = 32.0$ for O_2.

Sodium oxide has the formula Na_2O. So its M_r is given by:

$2 \times A_r$ of sodium + A_r of oxygen $= (2 \times 23.0) + 16.0 = 62.0$

One mole of any substance has a mass (in grams) equal to the relative formula mass (M_r) of the substance.

We can translate this statement usefully into a formula, to use in most situations, where we don't have exactly one mole of a substance:

$$moles = \frac{mass\ (g)}{M_{\mathrm{r}}}$$

We can also rearrange this to find mass, or to find M_r:

$$Mass\ (g) = moles \times M_r$$

$$M_{\mathrm{r}} = \frac{mass\ (g)}{moles}$$

We can apply these formulae to calculations involving the masses of substances reacting together, as described by an equation.

e.g. the equation for the reaction of sodium and oxygen is: $4Na + O_2 \rightarrow 2Na_2O$

We can read this equation as meaning that 4 moles of sodium react with each mole of oxygen and 2 moles of sodium oxide are formed as a result.

So, suppose we have 1.00g of sodium. We can calculate what mass of oxygen will react with it or what mass of sodium oxide will be formed. Here are the steps:

(1) Find the moles of sodium that you have. Moles = mass(g) / A_r = 1.00 / 23.0 = 0.04348 mol.

(2) Now look at the equation. It tells you that the moles of oxygen that will react is only ¼ of the moles of sodium (that's what the 4 in the equation means). So:
Moles O_2 reacting = ¼ x moles Na = ¼ x 0.04348 mol = 0.01087 mol

(3) We can now turn this number of moles beck into mass, to find the mass of oxygen reacting:
Mass(g) = mol x M_r = 0.01087 x 32.0 = 0.348g

So now we know: 1.00g sodium will react with 0.348g of oxygen. <u>We could only do this by using moles</u>.

The method for finding reacting masses is really important to get used to:

- convert mass to moles,
- use the reaction equation to relate the moles of differences involved
- convert back to mass.

Questions 3

1. **Sulfur occurs as a mixture of 3 isotopes: ^{32}S: 94.9% ^{33}S: 0.80% ^{34}S : 4.3%**
   Calculate the relative atomic mass of sulfur.

2. **The most common isotope of sulfur is ^{32}S. The relative atomic mass of sulfur is 32.1. What does this tell you?**

3. **What is the mass of 1 mole of water (use relative atomic masses from the periodic table)?**

   How many water molecules does this contain?

4. **How many moles of sodium oxide (Na_2O) are in 13.8g of this substance?**
   How many sodium ions are in this amount of sodium oxide?
   How many oxide ions are in this amount of sodium oxide?

5. **Here is the equation for the reaction between sodium metal and chlorine gas:**

$$2Na + Cl_2 \rightarrow 2NaCl$$

This means that each mole of chlorine reacts with two moles of sodium to form 2 moles of sodium chloride. So:
Suppose you are given 142g of chlorine for your birthday.

(i) **How many moles of chlorine is this?**
(ii) **So how many moles of sodium will react with it?**
(iii) **What is the mass of this amount of sodium?**
(iv) **What mass of sodium chloride will be formed?**

6. **Here is the equation for the reaction of aluminium and carbon monoxide:**

$$2Al + 3CO \rightarrow Al_2O_3 + 3C$$

Using the same kind of ideas as in question 5, see if you can work out what mass of aluminium you would need to produce 240g of aluminium oxide.

4. Structure and Bonding

There are basically 3 ways in which atoms can form strong bonds to other atoms:

- **Ionic bonding is** electrostatic attraction between oppositely charged ions. It most commonly occurs when a metal forms a compound with a non-metal.
- **Covalent bonding** happens when two atoms share a pair of electrons. The bond is the attractive force between the shared pair and the nuclei of the two atoms. This type of bonding is common when non-metals react with one another.
- **Metallic bonding** occurs between the atoms of metallic elements – it is the attractive force between a lattice of positive ions and a sea of delocalised electrons surrounding them.

For the moment, we're going to focus just on ionic bonding.

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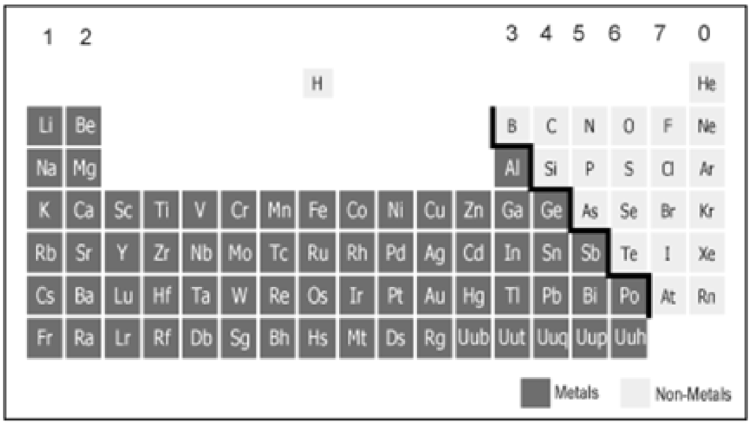
Ionic bonding

This most often occurs when a metal forms a compound with a non-metal (however this is not always the case, it is just a guideline for when it is reasonable that bonding will be ionic. So it is not the definition of ionic bonding. The correct definition is:

Ionic bonding is electrostatic attraction between oppositely charged ions.

So you can predict whether a compound will be ionic if you know which elements are metals and which are not. There is a rough boundary in the periodic table that makes this possible and you'll need to know where this boundary is:



When metals in the main groups of the periodic table (it's more complex for the transition elements) react with non-metals, the metal atoms usually lose all of their outer shell electrons to form positive ions. As a result the size of the positive charge is the same as the group number. e.g. calcium is in group 2, so a Ca atom has 2 outer shell electrons and when it reacts it loses both, to form a Ca^{2+} ion.

When non-metals react with metals, the non-metal atoms usually gain electrons to form negative ions. Specifically, they will gain electrons until they have 8 in the outer shell (we'll learn why this is the case during your A-level studies) e.g. oxygen is in group 6, so an O atom has 6 outer shell electrons. On reacting with a metal, it will therefore gain two electrons to form an O^{2-} ion (oxide) with 8 outer electrons.

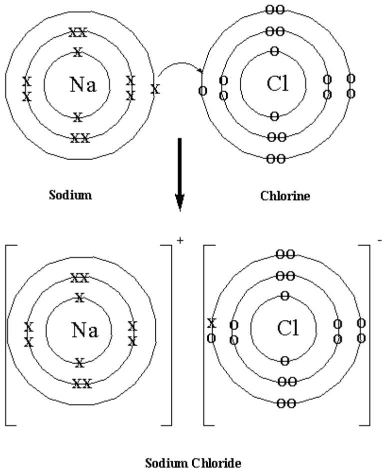
An ionic bond is the attraction between oppositely charged ions formed in this way. Ionic compounds are neutral overall, so the total charge of the positive and negative ions must cancel each other out. This enables you to work out the formula of an ionic compound

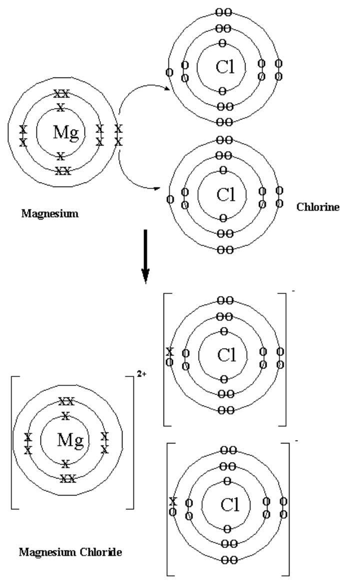
e.g. sodium forms Na^+ ions while oxygen forms O^{2-} ions. So to balance out the charges you need twice as many Na^+ as O^{2-} ions. The formula of sodium oxide is therefore Na_2O.

These websites may help you, if you're not yet comfortable predicting the formulae of ionic compounds:

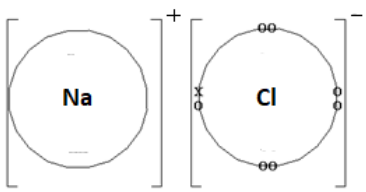
http://freesciencelessons.co.uk/charges-on-ions/ and
http://freesciencelessons.co.uk/formula-of-ionic-compounds/

You should remember how to draw dot and cross diagrams from GCSE for ionic compounds. Here are some examples:





At A-level we will still use these diagrams, but keep things simpler: We just show the electron arrangement once the bonds are formed (not what the atoms were like before this) and we just show just the outer shell electrons in the compound, eg for sodium chloride:



Key properties of ionic compounds can be understood in terms of this bonding model:

- Ionic compounds are usually solid at room temperature: melting them would require the ionic bonding to be partially broken down and, because this is very strong, it requires more energy than is available at room temperature.
- They do not conduct electricity in the solid state but they do when molten or dissolved in water. This is because conduction requires charged particles to be free to move. In the solid the strong ionic bonding locks the ions together so they can't move but these forces are weakened in the liquid and solution states, so the ions become mobile.

Some ionic compounds contain **polyatomic ions**: these are ions that are made up of a group of atoms joined together by covalent bonds, which between them share a positive or negative charge. There is a short list of common ones that you'll need to learn, so get ahead of the game by <u>learning them now</u>:

Polyatomic Ions to memorise:

Hydroxide OH^- **Sulfate SO_4^{2-}** **Nitrate NO_3^-** **Carbonate CO_3^{2-}**

Ammonium NH_4^+

Note: you have to be careful when writing the formula of a compound if it contains more than one polyatomic ion. E.g. calcium hydroxide is made up of Ca^{2+} and OH^-, and it obviously takes two OH^- for each Ca^{2+} to balance the charges. To show that there are two lots of the whole OH^- ion, you put a bracket around it in the formula:

$Ca(OH)_2$ is correct ….. $CaOH_2$ is wrong!.

Questions 4

1. **For each of the following reactions say whether the compound formed is ionic or covalent:**
   i. **Hydrogen and bromine**
   ii. **Potassium and sulfur**
   iii. **Carbon and aluminium**
   iv. **Uranium and fluorine**

2. **Write down the ions you would expect to be formed by the following elements:**
   i. Magnesium **ii. Iodine** **iii. Nitrogen**
   iv. Aluminium **v. Francium**

3. **For each of the following element combinations, suggest the formula of the compound formed if they react together and its name.**
   - **i. Potassium and chlorine**
   - **ii. Calcium and bromine**
   - **iii. Barium and sulfur**
   - **iv. Aluminium and oxygen**

4. **Suggest dot and cross diagrams (using the simplified, A-level format) for each of the compounds formed in the reactions in question 3.**

5. **Write down the formulae of the following compounds:**

i. Magnesium sulfate
ii. Ammonium fluoride
iii. Strontium nitrate
iv. Ammonium carbonate
v. Aluminium sulfate

6. **Carbon monoxide (CO) and magnesium oxide can both be turned into liquids by selecting a suitable temperature. Liquid CO is a non-conductor but liquid magnesium oxide is a good conductor. How do you explain this?**

5. Balancing Equations

At A-level you will need to be able to write balanced symbol equations for most of the reactions you come across, so practising balancing – and being able to do it quickly – is a skill that's worth developing at this point.

The first rule is that you must always make sure the individual substance formulae are correct before you start trying to do your balancing act. So:

- First decide if it's an ionic or covalent substance.
- If it's ionic, think about the ionic charge and then use the charge-balancing rule (see above) to get the formula.
- If it's covalent, there are various ways you can get the correct formula – but we won't get into that at the moment.

Now you can start putting your equation together:

1. Write down the correct chemical formula for each reactant and product
2. Check the equation to see if there are the same number of atoms of each element in the reactants and products. If there are not, you need to adjust the equation.
3. To set about balancing, start with an element that is only in one reactant and one product. Put a number IN FRONT of the particular reactant or product WITHOUT changing the formula, to make that element balance. Remember, these numbers in front are balancing numbers – they are not part of the formula of the individual substance. Any number in front multiplies every atom in that compound. These balancing numbers are now fixed.
4. Now do the same for a different element. Remember you can't change any balancing number you have already fixed. Continue to do this for each element, until both sides contain the same number of each type of atom.

   e.g. reaction of iron (III) oxide with magnesium:

   ……Fe_2O_3 + ……Mg → ……MgO + ……Fe

   Let's start with Fe: at the moment, we have 2 on the left but 1 on the right. We can fix this by putting in a balancing number:

   1 Fe_2O_3 + ……Mg → ……MgO + **2** Fe

   The underlined numbers are now fixed – we mustn't change them. So we can see that since we're stuck with **1** Fe_2O_3 , we'll have to have 3 oxygens on the right too. To make that happen, we need a balancing number for MgO:

   1 Fe_2O_3 + ……Mg → **3** MgO + 2 Fe

   Now we're almost there. We just need to make Mg balance, and that's easily done:

   1 Fe_2O_3 + **3** Mg → **3** MgO + **2** Fe

   We don't generally write "1" in an equation – it's just implied if there is no other balancing number. So we would write our final equation as:

   Fe_2O_3 + **3** Mg → **3** MgO + **2** Fe

If you would like more help in learning to balance equations, try these websites:

http://freesciencelessons.co.uk/balancing-chemical-equations/ and
http://www.wikihow.com/Balance-Chemical-Equations

Questions 5

1. **Balance the following equations. The formulae are all correct (so don't change them!)**

1.	____	KOH	+ ____	H_3PO_4	$\longrightarrow$	____	K_3PO_4	+ ____	H_2O
2.	____	NH_3	+ ____	O_2	$\longrightarrow$	____	NO	+ ____	H_2O
3.	____	Al	+ ____	O_2	$\longrightarrow$	____	Al_2O_3		
4.	____	CaS	+ ____	H_2O	$\longrightarrow$	____	$Ca(HS)_2$	+ ____	$Ca(OH)_2$
5.	____	Au_2S_3	+ ____	H_2	$\longrightarrow$	____	Au	+ ____	H_2S
6.	____	PCl_5	+ ____	H_2O	$\longrightarrow$	____	H_3PO_4	+ ____	HCl
7.	____	Ba_3N_2	+ ____	HF	$\longrightarrow$	____	BaF_2	+ ____	NH_3
8.	____	$CaCl_2$	+ ____	Na_3PO_4	$\longrightarrow$	____	$Ca_3(PO_4)_2$	+ ____	$NaCl$
9.	____	FeS	+ ____	O_2	$\longrightarrow$	____	Fe_2O_3	+ ____	SO_2
10.	____	$HClO_4$	+ ____	P_4O_{10}	$\longrightarrow$	____	H_3PO_4	+ ____	Cl_2O_7

2. For the following reactions, first make sure you have the right formula for each substance involved, then have a go at balancing them.

i. sodium carbonate + hydrochloric acid → sodium chloride + carbon dioxide + water
ii. sodium hydroxide + sulphuric acid → sodium sulphate + water
iii. sodium + water → sodium hydroxide + hydrogen
iv. sodium + chlorine → sodium chloride
v. iron (III) oxide + nitric acid →iron(III) nitrate + water
vi. burning magnesium in oxygen to give magnesium oxide
vii. reacting lithium with water to give lithium hydroxide and hydrogen
viii. burning C_5H_{12} in oxygen to give carbon dioxide and water
ix. reacting calcium with dilute hydrochloric acid (HCl) to give calcium chloride and hydrogen
x. reacting magnesium carbonate with dilute sulphuric acid (H_2SO_4) to give magnesium sulphate, carbon dioxide and water
xi. reacting sodium carbonate solution with magnesium chloride solution to give a precipitate of magnesium carbonate and another compound left in solution.

3. Aluminium reacts with hydrochloric acid (HCl) to make aluminium chloride and hydrogen. Write an equation for this reaction and use this to help you calculate what mass of hydrogen gas would be formed if 25.0g of aluminium reacts with the acid.