## OCR Chemistry A



## Preparation work

To be completed over the summer holidays.
(Remember to hand it in first week of year 12!!!)

## Atomic Structure

## What Are Atoms Like?

1) All atoms have a nucleus at their centre containing neutrons and protons.
2) Almost all of the mass of the atom is contained in the nucleus which also has an overall positive charge.
3) The positive charge arises because each of the protons in the nucleus, has a +1 charge.
4) The nucleus is relatively tiny compared with the total volume occupied by the whole atom.
5) The neutrons in the nucleus have a similar mass to the protons but they are uncharged.
6) The electrons orbit the nucleus in shells (or energy levels). The electrons are much smaller and lighter than either the neutrons or protons.
7) The volume occupied by the orbits of the electrons determines the size of the atom.

## What is the Charge on an

## Atom?

The overall charge on an atom is zero.
This is because each +1 charge from each of the protons in the nucleus is cancelled out by a -1 charge on each of the electrons.
If an atom loses or gains electrons it becomes charged. These charged particles are called ions. The fact that the protons and electrons are oppositely charged also helps to explain why the electrons remain in orbit: opposites attract.

Have a go at these questions:

1) Copy and complete the table:

| Particle | Relative <br> Mass | Charge |
| :---: | :---: | :---: |
| Proton | 1 |  |
| Neutron |  |  |
| Electron | $1 / 1840$ |  |

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

The atomic number of an element is given the symbol $Z$.
It is sometimes called the proton number as it represents the number of protons in the nucleus of a particular element.
For 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.

The mass number of an element is given the symbol A. It represents the total number of neutrons and protons in the nucleus. Subtracting $Z$ from $A$ allows you to calculate the number of neutrons in the nucleus.

Try this question (you may need to refer to a Periodic):

1) Copy and complete the table:

| Element name | Symbol | Z | A | Number of <br> protons | Number of <br> neutrons | Number of <br> electrons |
| :---: | :---: | :---: | :---: | :---: | :---: | :---: |
| Sodium |  |  | 23 |  |  |  |
|  |  | 6 | 12 |  |  |  |
|  |  | 12 |  |  |  |  |
| Chlorine |  | 84 | 210 |  |  |  |
| chlorine |  | 17 | 35 |  |  |  |

## Isotopes

The last two examples in the table above show two chlorine atoms with different numbers of neutrons. These are called isotopes of chlorine. Both are chlorine atoms because they have the same number of protons - but they have different numbers of neutrons. In other words they have the same atomic number but a different mass number. Isotopes are very common: some occur naturally and some are man-made. Some elements may have a large number of isotopes.

Have a go at these questions:
2) In terms of the numbers of subatomic particles, state one difference and two similarities between two isotopes of the same element.
3) Give the chemical symbol, mass number and atomic number of an atom which has 3 electrons and 4 neutrons.
4) Three isotopes of carbon are: carbon-12, carbon-13 and carbon-14. State the numbers of protons, neutrons and electrons in each.

## Relative Atomic Mass

## Calculating the Average Mass Number

When looking at the mass number of an element on a detailed copy of the Periodic Table, it is not always a whole number. This is because the value given is the average mass number for two or more isotopes. This idea is further complicated by the fact that some isotopes are more abundant than others.

For example, in a sample of chlorine there are on average 3 atoms with mass number 35 to every atom with mass number 37. The average mass number is 35.5.

How can we calculate the mass number as an average value for a number of isotopes?
Example: Calculate the average mass number for chlorine.
chlorine-35 chlorine-37
Ratio of atoms 3 : 1
Relative $3 / 4 \quad 1 / 4$
abundance
Average mass- abundance for Cl-35 x $35+$ abundance for $\mathrm{Cl}-37 \times 37$

$$
\begin{aligned}
& =3 / 4 \times 35+1 / 4 \times 37 \\
& =26.25+9.25 \\
& =35.5
\end{aligned}
$$

Note: the relative abundance is arrived at by considering that 3 out of 4 , ie. $3 / 4$ atoms, will be $\mathrm{Cl}-35$ and 1 out of 4 ie. 1/4 will be Cl-37.

Now have a go at these:

1) What is the average mass number of a sample of magnesium comprising 1 atom $\mathrm{Mg}-24$ to each Mg -25 atom?
2) What is the average mass number of a sample of carbon with $99 \mathrm{C}-12$ atoms to every C-13?
3) What is the average mass number of a sample of sulphur comprising 9 atoms $\mathrm{S}-32$ to every atom of S-33?
4) What is the average mass number of a sample of boron comprising 4 atoms $B-10$ to every atom ofB-11?
5) What is the average mass number of a sample of argon comprising 16 atoms Ar-4O to every 3 atoms ofAr-39?

## Arrangement of Electrons

## Electrons are Arranged in Energy Levels

Electrons orbit the nucleus in energy levels (also called shells).
The first energy level can contain up to 2 electrons. It is called an 's'level.
The second energy level can contain up to 8 electrons. However it is actually split into 2 sub-levels. Two of the electrons are in an 's' level and the remaining 6 are in a 'p' level.
At GCSE the 's' and ' $p$ ' sub-levels are not distinguished. We simply combine the 2 's' electrons with the 6 ' $p$ ' electrons to make a total of 8.

## How can Electron Arrangements be

## Represented?

Concentric circles can be drawn to represent the energy levels, and electrons drawn on each level as shown below.
For an atom with 21 electrons:
The diagram on the right shows the energy levels filling up with electrons.

Remember you should always start filling the innermost levels first.
This atoms fills as follows: $\mathbf{2}$ in the $\mathbf{1}^{\text {st }}$ shell, $\mathbf{8}$ in the $\mathbf{2}^{\text {nd }}$ shell, $\mathbf{8}$ in the $\mathbf{3}^{\text {rd }}$ shell and $\mathbf{3}$ in the $\mathbf{4}^{\text {th }}$ shell.

A simpler way to show electron filling is:


6 electrons arranged as: 2,4
11 electrons arranged as: 2,8,1

Use a Periodic Table to help you answer the following questions:

1) Draw diagrams to show the electron arrangements of the following elements: carbon, fluorine, magnesium, sulphur
2) Write the electron arrangements of the following elements using the format shown above: lithium, sodium, potassium, beryllium, magnesium, calcium

## Formulae of Compounds

## Deducing the Formulae of Ionic Compounds

The formula of a compound indicates the ratio of the elements in the compound. This ratio is fixed and for ionic compounds it is easy to work out the formula from the charges on the ions.

Metal ions (and hydrogen ions) always carry a positive charge, whilst non-metal ions carry a negative charge. If you imagine that a positive charge is a 'hook' and a negative charge is an 'eye' then the formula can be deduced by exactly matching up the hooks and eyes.
$\mathrm{Na}^{+}$(sodium ion) has +1 charge so 1 hook $\mathrm{OH}^{-}$(hydroxide ion) has -1 charge so 1 eye $\mathrm{Mg}^{*}$ (magnesium ion) has +2 charge so 2 hooks $O^{2-}$ (oxide ion) has -2 charge so 2 eyes

Example 1: What is the formula of sodium oxide?


We need an extra $\mathrm{Na}^{+}$to give us a second hook to match the second of the eyes on the $\mathrm{O}^{2-}$ ion.

We have $2 \mathrm{Na}^{+}$ions to every $\mathrm{O}^{2-}$ ion, so the formula is $\mathrm{Na}_{2} \mathrm{O}$.

Example 2: What is the formula of magnesium hydroxide?


There are $2 \mathrm{OH}^{-}$ions to every $\mathrm{Mg}^{2+}$ ion so the formula is $\mathrm{Mg}(\mathrm{OH})$.

Now try these:
Use the charges on the ions at the bottom of the box to deduce the formulae of the following ionic compounds.

1) sodium chloride
2) potassium oxide
3) calcium bromide
4) aluminium chloride
5) sodium carbonate
6) potassium nitrate
7) aluminium oxide
8) aluminium sulphate
9) iron (II) chloride
10) iron (III) nitrate
aluminium: $\mathrm{Al}^{3+}$ bromide: $\mathrm{Br}{ }^{-}$calcium: $\mathrm{Ca}^{2+}$ carbonate: $\mathrm{CO}_{3}{ }^{2-}$ chloride Cl - iron(II): $\mathrm{Fe}^{2+}$ iron (III): $\mathrm{Fe}^{3+}$ nitrate: $\mathrm{NO}_{3}$ oxide: $\mathrm{O}^{2-}$ potassium: $\mathrm{K}^{+}$sodium: $\mathrm{Na}^{+}$sulphate: $\mathrm{SO}_{4}{ }^{2-}$

## Writing and balancing equations

## Rules for Working out the Products Formed in

## Reactions

In order to write a balanced symbol equation from scratch you need to be aware of the different ways that compounds react. Many examples involve the reactions of acids to form salts, and it helps if you are aware of some of the rules for working out the names of the products formed from particular reactants.

## Making Salts:

1. If sulphuric acid is used the salt will be ' $x x x$ ' sulphate, where ' $x x x$ ' is a metal.
2. If hydrochloric acid is used the salt will be ' $x x x$ ' chloride, where ' $x x x$ ' is a metal.
3. If nitric acid is used the salt will be ' $x x x$ ' nitrate where, ' $x x x$ ' is a metal.
(Sulphuric acid $=\mathrm{H}_{2} \mathrm{SO}_{4} \quad$ Hydrochloric acid $=\mathrm{HCl} \quad$ Nitric acid $=\mathrm{HNO}_{3}$ )
Reactions Involving Acids:
4. Metal + acid $\longrightarrow$ salt + hydrogen
5. Metal oxide + acid $\longrightarrow$ salt + water
6. Metal hydroxide + acid $\longrightarrow>$ salt + water
7. Metal carbonate + acid $\longrightarrow$ salt + water + carbon dioxide

## Some Other General Rules:

8. Combustion reactions result in the formation of oxides.
9. When fuels burn, carbon dioxide $\left(\mathrm{CO}_{2}\right)$ and water $\left(\mathrm{H}_{2} \mathrm{O}\right)$ are normally produced.
10. Atoms in elements that are gases often go round in pairs: $\mathrm{H}_{2}, \mathrm{~N}_{2}, \mathrm{O}_{2}, \mathrm{Cl}_{2}$

Use the ions on the previous page and the rules above to answer the following:
Give the names and formulae of all the products formed in these reactions.

1) aluminium + nitric acid
2) potassium hydroxide + sulphuric acid
3) the complete combustion of propanol $\left(\mathrm{C}_{3} \mathrm{H}_{7} \mathrm{OH}\right)$ in oxygen
4) calcium carbonate + hydrochloric acid
5) aluminium oxide + sulphuric acid
6) sodium hydroxide + nitric acid
7) zinc + hydrochloric acid

## Writing and Balancing Equations <br> Writing Balanced Equations

To write a balanced symbol equation where reactants are given there are 5 simple steps:

1. Write out the word equation first.
2. Write the correct formula for each compound below its name (see page 26).
3. Go through each element in turn, making sure the number of atoms on each side of the equation balances.
4. If you changed any numbers, do step 3 again.
5. Keep doing this 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:

$$
\mathrm{MgO}+2 \mathrm{HCl} \rightarrow \mathrm{MgCl}_{2}+\mathrm{H}_{2} \mathrm{O}
$$

There are two Cl on the right of the equation, so we need to have two HCl on the left-hand side. This also doubles the number of hydrogen atoms on the left-hand side, so that the hydrogen's balance as well. This always works. If you can't get an equation to balance then it's wrong.

The examples below use the rules from the previous page to write out the word and symbol equations. Read through them and then try the questions on the next page.

Example 1: Write a balanced equation for the combustion of methane $\left(\mathrm{CH}_{4}\right)$ in oxygen.
Step 1: $\quad$ Methane + oxygen $\longrightarrow$ carbon dioxide + water (using rule 9)
Step 2:

$$
\mathrm{CH}_{4}+\mathrm{O}_{2} \rightarrow \mathrm{CO}_{2}+\mathrm{H}_{2} \mathrm{O}
$$

(using rule 10)
Step 3:

$$
\mathrm{CH}_{4}+2 \mathrm{O}_{2} \rightarrow \mathrm{CO}_{2}+2 \mathrm{H}_{2} \mathrm{O}
$$

(the Cs already balance, put a 2 in front of $\mathrm{H}_{2} \mathrm{O}$ to balance the Hs , now put a 2 in front of $\mathrm{O}_{2}$ to balance the O's. Check that all still balances.)

Example 2: Write a balanced equation for the reaction of magnesium with hydrochloric acid.
Step 1:
Magnesium + hydrochloric acid $\rightarrow$ magnesium chloride + hydrogen
(using rules 2 and 4)
Step 2:

$$
\mathrm{Mg}+\mathrm{HCl}->\mathrm{MgCl}_{2}+\mathrm{H}_{2}
$$

(using rule 10)
Step3:

$$
\mathrm{Mg}+2 \mathrm{HCl} \rightarrow \mathrm{MgCl}_{2}+\mathrm{H}_{2}
$$

(the Mg's already balance, put a 2 in front of HCl to balance the H's and Cl's. Check that all still balances.)

## Writing and Balancing Equations

Combine what you've learnt on the previous three pages to answer the following:
Write balanced symbol equations for these reactions.
(The charges on the ions are given at the bottom of the questions.)

1) the complete combustion of propane $\left(\mathrm{C}_{3} \mathrm{H}_{8}\right)$ in oxygen
2) the complete combustion of ethanol $\left(\mathrm{C}_{2} \mathrm{H}_{5} \mathrm{OH}\right)$ in oxygen
3) sodium hydroxide + nitric acid
4) potassium oxide + hydrochloric acid
5) sodium hydroxide + sulphuric acid
6) magnesium carbonate + nitric acid
7) sodium carbonate + sulphuric acid
8) potassium carbonate + nitric acid
9) the complete combustion of octane $\left(\mathrm{C}_{8} \mathrm{H}_{18}\right)$ in oxygen
10) calcium hydroxide + hydrochloric acid
sodium: $\mathrm{Na}^{+}$carbonate: $\mathrm{CO}_{3}{ }^{2-}$ hydrogen: $\mathrm{H}^{+}$potassium: $\mathrm{K}^{+}$
chloride: $\mathrm{Cl}^{-}$nitrate: $\mathrm{NO}_{3}{ }^{-}$hydroxide: $\mathrm{OH}^{-}$calcium: $\mathrm{Ca}^{2+}$
oxide: $\mathrm{O}^{2-}$ magnesium: $\mathrm{Mg}^{+2}$ sulphate: $\mathrm{SO}^{2-}$

Answers (Use these to check your work once you have completed it) Atomic Structure - page1
1)

| Particle | Relative <br> Mass | Charge |
| :---: | :---: | :---: |
| Proton | 1 | $\mathbf{+}$ |
| Neutron | $\mathbf{1}$ | None |
| Electron | $1 / 1840$ | $\mathbf{-}$ |

2) $2+$
3) $\mathbf{2 -}$

## Atomic Number, Mass Number and Isotopes - page 2

1) 

| Element name | Symbol | Z | A | Number of <br> protons | Number of <br> neutrons | Number of <br> electrons |
| :---: | :---: | :---: | :---: | :---: | :---: | :---: |
| Sodium | $\mathbf{N a}$ | 11 | 23 | $\mathbf{1 1}$ | $\mathbf{1 2}$ | $\mathbf{1 1}$ |
| Carbon | $\mathbf{C}$ | 6 | 12 | $\mathbf{6}$ | $\mathbf{6}$ | $\mathbf{6}$ |
| Magnesium | $\mathbf{M g}$ | 12 | $\mathbf{2 4}$ | $\mathbf{1 2}$ | $\mathbf{1 2}$ | $\mathbf{1 2}$ |
| Polonium | $\mathbf{P o}$ | 84 | 210 | $\mathbf{8 4}$ | $\mathbf{1 2 6}$ | $\mathbf{8 4}$ |
| Chlorine | $\mathbf{C l}$ | 17 | 35 | $\mathbf{1 7}$ | $\mathbf{1 8}$ | $\mathbf{1 7}$ |
| Chlorine | $\mathbf{C l}$ | 17 | 37 | $\mathbf{1 7}$ | $\mathbf{2 0}$ | $\mathbf{1 7}$ |

2) Different number of neutrons, same number of protons, same number of electrons
3) Li, 7, 3. 4)

|  | Protons | Neutrons | electrons |
| :--- | :---: | :---: | :---: |
| Carbon-12 | 6 | 6 | 6 |
| Carbon-13 | 6 | 7 | 6 |
| Carbon-14 | 6 | 8 | 6 |

## Calculating the Average Mass Number - page 3

1) 24.5
2) 12.013$) 32.36$
3) 10.2
4) 39.84

## Arrangement of electrons - page 4

1. Carbon

Fluorine


2. Lithium - 2, 1 Sodium - 2, 8, 1

Potassium - 2, 8, 8, 1
Beryllium - 2, 2
Magnesium - 2, 8, 2
Calcium - 2, 8, 8, 2

## Formulae of compounds - page 5

| 1) | NaCl | 6) | $\mathrm{K}_{2} \mathrm{O}$ |
| :--- | :--- | :--- | :--- |
| 2) | $\mathrm{CaBr}_{2}$ | 7) | $\mathrm{AlCl}_{3}$ |
| 3) | $\mathrm{Na}_{2} \mathrm{CO}_{3}$ | 8) | $\mathrm{KNO}_{3}$ |
| 4) | $\mathrm{Al}_{2} \mathrm{O}_{3}$ | 9) | $\mathrm{Al}_{2}\left(\mathrm{SO}_{4}\right)_{3}$ |
| 5) | $\mathrm{FeCl}_{2}$ | 10) | $\mathrm{Fe}\left(\mathrm{NO}_{3}\right)_{3}$ |

## Writing and balancing equations - page 6

8) Aluminium nitrate $\left(\mathrm{AlNO}_{3}\right)+$ hydrogen $\left(\mathrm{H}_{2}\right)$
9) potassium sulphate $\left(\mathrm{K}_{2} \mathrm{SO}_{4}\right)+$ water $\left(\mathrm{H}_{2} \mathrm{O}\right)$
10) Carbon dioxide $\left(\mathrm{CO}_{2}\right)+$ Water $\left(\mathrm{H}_{2} \mathrm{O}\right)$
11) calcium chloride $\left(\mathrm{CaCl}_{2}\right)+$ water $\left(\mathrm{H}_{2} \mathrm{O}\right)+$ carbon dioxide $\left(\mathrm{CO}_{2}\right)$
12) aluminium sulphate $\left(\mathrm{Al}_{2}\left(\mathrm{SO}_{4}\right)_{3}\right)+$ water $\left(\mathrm{H}_{2} \mathrm{O}\right)$
13) sodium nitrate $\left(\mathrm{NaNO}_{3}\right)+$ water $\left(\mathrm{H}_{2} \mathrm{O}\right)$
14) zinc chloride $\left(\mathrm{ZnCl}_{2}\right)+$ hydrogen $\left(\mathrm{H}_{2}\right)$

## Writing and balancing equations - page 8

11) $\mathrm{C}_{3} \mathrm{H}_{8}+5 \mathrm{O}_{2} \rightarrow 3 \mathrm{CO}_{2}+4 \mathrm{H}_{2} \mathrm{O}$
12) $\mathrm{C}_{2} \mathrm{H}_{5} \mathrm{OH}+3 \mathrm{O}_{2} \rightarrow 2 \mathrm{CO}_{2}+3 \mathrm{H}_{2} \mathrm{O}$
13) $\mathrm{NaOH}+\mathrm{HNO}_{3} \rightarrow \mathrm{NaNO}_{3}+\mathrm{H}_{2} \mathrm{O}$
14) $\mathrm{K}_{2} \mathrm{O}+2 \mathrm{HCl} \rightarrow 2 \mathrm{KCl}+\mathrm{H}_{2} \mathrm{O}$
15) $2 \mathrm{NaOH}+\mathrm{H}_{2} \mathrm{SO}_{4} \rightarrow \mathrm{Na}_{2} \mathrm{SO}_{4}+2 \mathrm{H}_{2} \mathrm{O}$
16) $\mathrm{MgCO}_{3}+2 \mathrm{HNO}_{3} \rightarrow \mathrm{Mg}\left(\mathrm{NO}_{3}\right)_{2}+\mathrm{H}_{2} \mathrm{O}+\mathrm{CO}_{2}$
17) $\mathrm{Na}_{2} \mathrm{CO}_{3}+\mathrm{H}_{2} \mathrm{SO}_{4} \rightarrow \mathrm{Na}_{2} \mathrm{SO}_{4}+\mathrm{H}_{2} \mathrm{O}+\mathrm{CO}_{2}$
18) $\mathrm{K}_{2} \mathrm{CO}_{3}+2 \mathrm{HNO}_{3} \rightarrow 2 \mathrm{KNO}_{3}+\mathrm{H}_{2} \mathrm{O}+\mathrm{CO}_{2}$
19) $\mathrm{C}_{8} \mathrm{H}_{18}+121 / 2 \mathrm{O}_{2} \rightarrow 8 \mathrm{CO}_{2}+9 \mathrm{H}_{2} \mathrm{O}$ or $2 \mathrm{C}_{8} \mathrm{H}_{18}+25 \mathrm{O}_{2} \rightarrow 16 \mathrm{CO}_{2}+18 \mathrm{H}_{2} \mathrm{O}$ (both are correct!) 20) $\mathrm{Ca}(\mathrm{OH})_{2}+2 \mathrm{HCl} \rightarrow \mathrm{CaCl}_{2}+2 \mathrm{H}_{2} \mathrm{O}$

