

Year 11 revision, 2007

For

Daisy, Ellie, Eleanor, Esha, Naomi, Abinaya, Marina, Jade, Dasha, Nikki,
Aimee, Avisha, Hannah, Chelsie, Anna and Alysha

This is not designed to replace your textbook, notes or your CGP guide – it's just an extra resource. As it is difficult (and time consuming) to draw pictures on computer I haven't bothered...

Questions to drmckeating@hotmail.com please.

Other useful sites...

www.chemguide.co.uk

www.creative-chemistry.org.uk

www.bbc.co.uk/schools/gcsebitesize/chemistry/

Atomic structure

Atomic structure is a simple topic at GCSE but you need to be careful if you're going to get the best marks.

Particle	Charge	Mass	Location
Proton	+1	1	Nucleus
Neutron	0	1	Nucleus
Electron	-1	0	In shells orbiting the nucleus

You should know the above information as it forms the basis of most of the subject! (Remember that the **neutron** is **neutral** and that the electron has negligible or zero mass and the rest should come flooding back.)

You should be able to draw a simple atom and recall, or work out, the **electronic structure** of the first 20 elements in the periodic table. For instance sulphur has the electronic structure 2, 8, 6.

Atoms are neutral if they are found on their own, but most are involved in chemistry of some description, meaning that they are combined with other elements.

Bear in mind that electrons are negative...

If an atom **loses** an electron it becomes **positive** (the positive charge of the protons outnumbers the negative charge of the electrons)

If an atom **gains** an electron it becomes **negative** (the negative charge of the electrons outnumbers the positive charge of the protons)

The number of protons in an atom defines what element it is. For instance, every atom with 11 protons is sodium, Na, regardless of how many neutrons it has.

Some elements have different atoms – the difference is the number of neutrons in the nucleus. All this means is that the atoms have different weights (mass numbers). Atoms with the same number of protons but different numbers of neutrons are called **isotopes**.

${}_{35}^{79}\text{Br}$ and ${}_{35}^{81}\text{Br}$ are both isotopes of bromine. They exist in equal proportions and if you look at the periodic table you will see that the symbol for bromine is ${}_{35}^{80}\text{Br}$, even though there are no individual bromine atoms with a **mass number** of 80!

The smaller number (they can be printed either way round and should be labelled on the periodic table in your exam paper) is called the **atomic number**.

Bonding

For GCSE you need to know about 3 kinds of bonding...

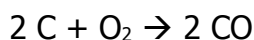
1. Covalent bonding
Covalent bonding involves the **sharing of pairs of electrons**
2. Ionic bonding
Ionic bonding involves the **exchange of electrons** to form **ions**
3. Metallic bonding
Metallic bonding involves an array of positive ions with a sea of delocalised electrons.

Covalent bonding happens between **non-metal** atoms. Examples of non-metal atoms you should be able to draw the bonding in are water, ammonia, carbon dioxide, hydrogen chloride.

There are also some molecules called the **homonuclear diatomics**. 7 elements occur naturally in pairs, so if you are writing the element in an equation you should write it as a molecule containing two atoms. These 7 elements are hydrogen, H₂, oxygen, O₂, nitrogen, N₂, iodine, I₂, fluorine, F₂, bromine, Br₂, and chlorine, Cl₂. They can be remembered by the handy mnemonic honifbrcl, which is better said out loud.

An example of when you might need this information is the question
Write an equation showing how carbon monoxide is formed from carbon and oxygen in the air.

Your answer would be



Remember that atoms share electrons in order to fill up their outer shells of electrons – if you are in any doubt count the electrons around the atoms.

Ionic bonding happens between **metals** and **non-metals**. This means you need to know where the metals and non-metals are on the periodic table! The metal will always lose electrons to the non-metal. The metal will form a positive ion and the non-metal will form a negative ion.

You can work out the charges on an atom by using the periodic table. The atoms gain or lose electrons in an effort to end up with a full outer shell. Group 1 (metals) lost 1 electron and will have a charge of +. Group 2 (metals) will have a charge of 2+. Group 3 (metals) will have a charge of 3+.

Atoms from groups 4 and 5 seldom form ions.

Group 6 (non-metals) gain 2 electrons to gain a charge of 2-. Group 7 (non-metals) will have a charge of -.

Group 8 is also known as group 0 – the atoms neither gain nor lose electrons and therefore do not commonly form ions.

Structures and intermolecular forces

This can be a really tricky topic. You often get questions asking why something has a high or a low boiling point and should remember that the answer to this will involve either talking about a giant structure or a simple molecular one.

Giant covalent structures normally have a three-dimensional lattice (graphite is layered however). Examples you should know are graphite and diamond (both allotropes of carbon) and silicon dioxide, SiO_2 . A substance with a giant covalent structure will be a solid at room temperature. This is because to melt it all of the bonds need to be broken to allow the atoms or molecules to move around. Silicon dioxide is sand and you may be asked to compare it to carbon dioxide, because carbon and silicon are in the same group so you might expect their oxides to have similar physical properties. CO_2 has a simple molecular structure.

Simple molecular structures are any non-metal element or compound. They are normally gases (but sometimes liquids) at room temperature. They have weak forces BETWEEN the molecules. Remember that covalent bonds are always 'strong'. If you try to explain the low boiling point of a molecular substance by saying there are weak covalent bonds then you are wrong! When CO_2 , for instance, changes from solid to gas (it sublimates) the individual molecules of CO_2 remain as CO_2 units – only the gaps between the molecules get bigger.

Giant ionic structures form in all ionic compounds. They have a giant, repeating array or lattice of ions. The ions are held together by strong forces. To melt an ionic compound all of these forces between all of these ions have to be overcome enough to allow the ions to move around. For this reason ionic substances normally have high melting points and are solids at room temperature.

When a metal is on its own as an element, like a block of aluminium, it has a giant **metallic** structure. This is 'an array of positive metal ions in a sea of electrons'. The electrons can move, which is why all metals can conduct electricity. The electrons come from the outer shell of each metal atom, leaving a positive ion (remember that metals lose electrons to leave an ion with a full outer shell).

The Periodic Table

The periodic table is your weapon in the fight to get a good grade, even if you hate it. A bit of familiarity is a good thing.

- The columns are called **groups**
- The rows are called **periods**
- Elements in the same group have the same number of outer electrons and this (mostly) means that they have very similar chemistry
- The elements are arranged in order of increasing atomic number – H is 1, He is 2, Li is 3 and so on

The group 1 metals are called alkali metals because they react with water to produce an alkali (a soluble metal hydroxide). You should be able to write an equation for any of the group 1 metals reacting with water to produce the correct hydroxide and hydrogen gas. The metals are soft (easy to cut) and have low melting points, which makes them unlike most metals. They get more reactive as you go down the group – lithium fizzes a bit in water, caesium explodes! This is because the one outer electron they all lose when they react is further from the positively charged nucleus in the bigger atoms – it is more easily lost.

Group 7 contains the halogens. These are fluorine, chlorine, bromine, iodine and astatine. The first and last of these are not often asked about at GCSE. As you go down the group the melting points increase and they get darker in appearance – chlorine is a pale green gas, bromine is a reddish liquid, iodine is a dark grey solid. The melting points get higher because the forces between the molecules get stronger. They all react by gaining an electron. The argument for their reactivity is therefore the opposite of that for the alkali metals: The smaller atoms are better at gaining electrons because the attraction of the positive nucleus is greater than in the big atoms. You should learn a use of each halogen.

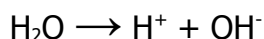
Group 8 (or 0) contains the noble gases. They have a full outer shell of electrons and so don't react – they don't gain or lose any electrons to form bonds. They are **inert**. As the atoms get bigger the forces between them get larger and their boiling points increase, but they're still gases at room temperature. You should find some uses of each of them.

The transition metals are the 'big block' 10 elements wide in the middle of the table. They often form coloured compounds. As they are metals they too form positive ions when they react. You are not expected to know how many electrons they lose apart from that silver forms Ag^+ ions and zinc forms Zn^{2+} ions. You should otherwise be told what the charge on the metal is in a compound. Iron (II) oxide contains Fe^{2+} and copper (I) oxide contains Cu^+ for instance.

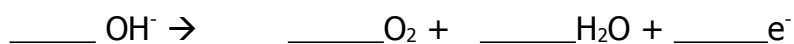
Electrolysis

Ah, electrolysis. Splitting up with electricity. Rules and guidelines:

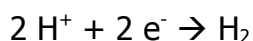
- You can only electrolyse liquids (molten ionic compounds normally) or solutions – this is so that the 'electric charge' can be transferred by the free-to-move ions
- PANC – positive anode, negative cathode
- Opposites attract – negative ions will go to the anode and positive ions to the cathode
- Electrolysing a molten substance will produce the elements that make it up – sodium chloride will give sodium and chlorine; aluminium oxide will give aluminium and oxygen (the hot oxygen will react with a carbon electrode to form carbon dioxide!)
- Electrolysis of solutions is more complicated because water splits into hydrogen and hydroxide ions:



- This means that there are other positive and negative ions which can be attracted to the electrodes
- If the solution contains iodide, bromide or chloride then iodine, bromine and chlorine are usually formed, otherwise oxygen is formed from the hydroxide ions (balance the following)



- If the solution contains ions of a metal MORE reactive than hydrogen, then hydrogen is formed at the cathode

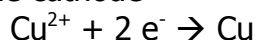


- If the solution contains ions of a metal LESS reactive than hydrogen then the metal is formed at the cathode

Examples would be:

1. What is formed when copper sulphate solution is electrolysed?

Copper metal is formed at the cathode



Oxygen forms at the anode (see the equation you balanced earlier)

2. What is formed when molten aluminium chloride is electrolysed?

Aluminium and chlorine are formed at the cathode and anode respectively

3. What is formed when salt solution is electrolysed?

Salt solution is sodium chloride, NaCl. However, Na is more reactive than hydrogen and hydrogen gas forms at the cathode. Chlorine gas forms at the anode.

Organic Chemistry

It's all about carbon. Carbon can form bonds with other carbon atoms, meaning that long chains and rings can form. The properties of organic molecules met at GCSE are dominated by their sizes and shapes. Organic compounds normally only include non-metal elements. What sort of bonding do you expect to find in organic molecules? Try drawing a 'dot and cross' diagram of CH_4 and then name it.

Alkanes

Alkanes are examples of hydrocarbons – compounds containing **only** hydrogen and carbon (the 'only' is really important). What is the general formula for alkanes?

You only need to know about alkanes up to chains of four carbon atoms – that means methane, ethane, propane and butane.

Alkenes

The alkenes are another homologous series – a group of chemicals with the same basic formula. In this case C_nH_{2n} . The difference between them and alkanes is that the alkanes are **saturated** – they contain only single bonds. The alkenes all contain a carbon-carbon double bond, $\text{C}=\text{C}$. This makes them more reactive than alkanes. They react with bromine water to turn it from orange to colourless – this is a useful **test** for alkenes.

Polymers

The polymers you need to know about for GCSE chemistry are called **addition polymers**. This means that they are formed by lots of smaller molecules (**monomers**, usually alkenes) adding together to form a long chain molecule. Double bonds in alkenes can 'spring open' to join up the molecules. What monomer leads to the formation of the polymer shown below?

Extraction and uses

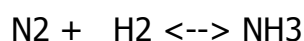
The alkanes can be extracted from crude oil which is a mixture of lots of different alkanes. Fractional distillation separates the components of the mixture because bigger molecules have higher boiling points. Big molecules can be converted into more useful short-chain molecules using a process called **cracking**.

Most of the alkanes are used as fuels of some sort – you should know the uses of the fractions of crude oil, as this is a common question. It's something you just have to learn...

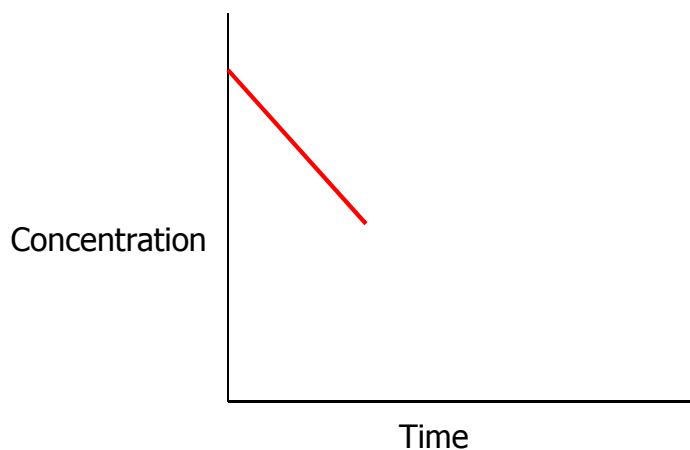
Equilibrium

Equilibrium describes processes which go 'forwards' and 'back'. By 'forwards' we mean the normal 'reactants go to products' and by 'back' we mean that the products can recombine to form reactants.

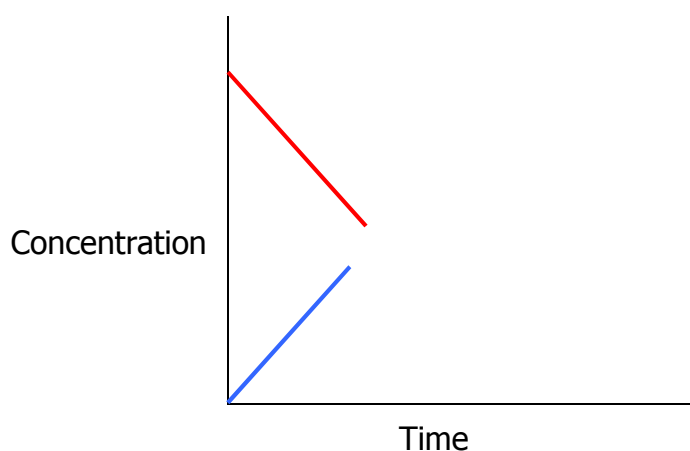
It is important to remember that in equilibrium systems no stuff can get in or out of the reacting system – we can only change the pressure, or change the temperature by heating or cooling. We will use the example of the Haber process, as it is very often asked about and it contains all the ideas we need. The equation for the Haber process is



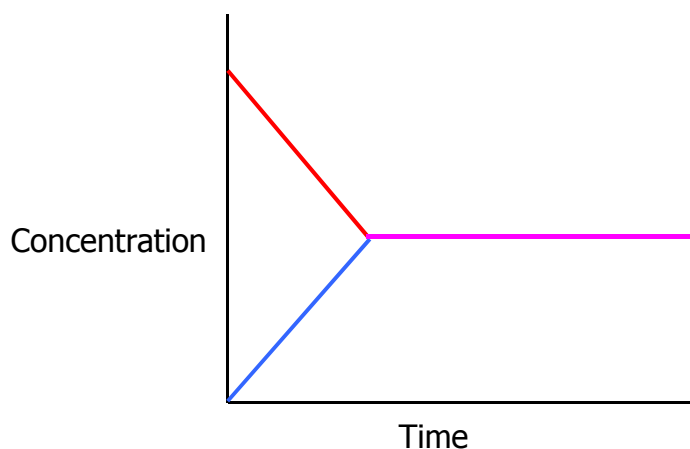
Imagine we start off with a mixture of nitrogen and hydrogen. As they react their concentrations will fall (red line).



As they react they produce ammonia, so the concentration of NH₃ rises.



When the two lines meet the rates of the forward and back reactions are the same (remember that rate is a change in some quantity per unit time, in this case change in concentration per second) and the concentrations of the two things stop changing.



The system has reached a state of **dynamic equilibrium**: 'Dynamic' because the forward and back reactions are still happening; 'equilibrium' because the overall concentrations of the hydrogen, nitrogen and ammonia are no longer changing.

What happens when you change the conditions?

This is controlled by Le Chatelier's principle, which can be stated a bit like this:

A system at dynamic equilibrium shifts the position of equilibrium to oppose any change made.

When the temp is increased, the position of equilibrium (the relative proportions of the reactants and products) moves to favour the substances on the endothermic side of the reaction.

When the temp is decreased, the position of equilibrium moves to favour the substances on the exothermic side of the reaction.

When the pressure is increased, the position of equilibrium moves to favour the side with fewer **gaseous** particles.

When the pressure is decreased, the position of equilibrium moves to favour the side with more **gaseous** particles.