Chemistry - (Further Chemistry)

AQA GCSE Chemistry Revision Notes – Module Three Higher Tier

Disclaimer; I made this to help me revise when the exam comes around, but while I’ve tried to make this document as accurate as possible, I can’t guarantee that the information in this document is completely correct – I may have made mistakes.

# C3

## – The early periodic table

* Early chemistry was in the main done by people with lots of money, who had time on their hands to experiment with things.
* These people often didn’t know much about the chemistry behind what they were trying to do otherwise, so many wouldn’t have tried to turn lead to gold.
* Some people tried to categorise elements into tables, these tables were expanded on by many people until the periodic table we know today was made. Unfortunately, many early tables included false elements like ‘potash’.
* These tables attempted to classify elements.
* A major breakthrough came when Russian scientist named ‘Mendeleev’ thought to leave gaps for elements he hadn’t yet discovered, his table was the first *periodic* table.

## – The modern periodic table

* In the periodic table, there are eight groups (0-7), arranged in columns (although there are more than eight columns).
* The group number indicates the amount of electrons on the outer shell of each element in the group (e.g. Sodium has one electron in its outer shell).

## 1.3 – Group one (Alkali metals)

* Each group of elements share similar properties; In *group one:*
  + All the elements are metals.
  + All are soft.
  + All are shiny when freshly cut.
  + All react readily with oxygen and water (e.g. 4Na + O2 > 2NaO2).
  + All are less dense than water (except caesium)
  + Produce hydrogen and the metal hydroxide when reacted with water

(e.g. 2Na + 2H2O > 2NaOH + H2).

* As *group one* (alkali metals) descends the elements become more reactive, this is because the outer shell electron is further away from the nucleus (which holds it to the atom), and so it is easier to remove. A group one atom needs only to remove one electron in order to become stable.

## 1.4 – Group seven (Halogens)

* In *group seven* (halogens), as the elements get bigger (and the group descends) the elements become less reactive, as they need to gain one electron to become stable and the outer shell is far away from the nucleus.
* None of the halogens are ever found on their own, they are always in pairs (i.e. ‘F’ is always ‘F2’. This is because they are unstable until they covalently bond together, sharing one electron from each atom.
* When a halogen reacts with a metal, it forms a ionic bond, as when it reacts with a non-metal, it always makes a covalent bond (hence the last point).
* **These changes in the reactivity (for 1.3 - 1.4) are due to a phenomenon called ‘shielding’ (but that’s not in GCSE module three).**
* **To sum it up (1.2 – 1.4), the chemical properties of an element are determined by the number of electrons in the outer shell, the group that an element is in is determined by its atomic number and we can explain trends in reactivity by looking at the number of energy levels in atoms.**

## 1.5 – The transition elements (or metals)

* These elements are found in the middle of the periodic table, roughly in-between the alkali metals and the non-metals.
* The transition elements have a typical metallic structure – each metal is held together in a giant metallically bonded structure.
* The outer electrons of each atom are free to move around, so they are good conductors of electricity and heat.
* The transition metals are all tough, strong and hard, but we can bend or hammer them into useful shapes.
* The only transition element with a low melting point is mercury, which is a liquid at room temperature.
* Many of the transition elements are brightly coloured when in a compound because shell three is normally only partially full. This also makes them good catalysts, (E.g. Iron in the Haber process).
* All of the transition metals are less reactive than the metals in group one. This means that they don’t corrode as easily, so are useful as structural materials, especially when combined with other elements to form a steel (E.g. Iron and Carbon)
* If a compound has a name such as *copper(II) sulphate*, the Roman Numeral on the end denotes the charge of the ion with the number after it. In this example, we can see that the copper ion has a positive charge of two. (The transition elements can have more than one ion.

## 2.1 – Strong and weak acids/alkalis

* When measuring acidity (or not) we give a value from 0-14. 7 is neutral, below that is acid and below is alkali.
* An acid in an aqueous solution forms H+ ions, while alkalis in an aqueous solution form OH- ions.
* A strong acid or base is 100% ionised in water. (E.g. Hydrochloric acid is fully ionised)
* A weak acid or alkali is only partially ionised in water (it’s a reversible reaction – the position of the equilibrium determines the acidity).
* Acids donate protons and bases accept protons (H+ ions have a charge that is positive and so need to lose a positive electron, OH- ions have a negative charge, so need to gain a proton).

## 2.2 – Titrations

* Titration is a technique used to precisely measure and react an acid and alkali *and find the endpoint of the reaction*.
* The PH of the solution left when you’ve done your titration isn’t always 7. Different indicators change colour over different PH ranges, so it’s important to choose a suitable indicator for the reaction you are doing – i.e. don’t choose an indicator that changes colour at PH 4 if the endpoint of your reaction is PH 9.
* You can use a titration to accurately measure how much of an acid reacts with an alkali (vice versa)
* The end point of the reaction is (well guessed) when the reaction is complete.
* A suitable indicator must change colour at the end point of the reaction.

## 2.3 – Titration calculations

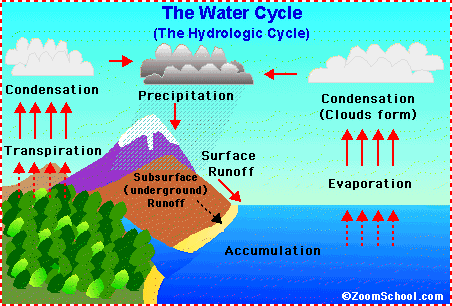
* The concentration **of a solution is measured in moles per cubic decimetre** (mol/dm3) (A cubic decimetre is really just a fancy name for one litre).
* If the amount of the substance dissolved in the solution is know, and the amount of the solution is known then we can work out the concentration. E.g. If we dissolve 40 grams of sodium hydroxide to make 1dm3 of solution we find the concentration simply by finding the mass of one mole of sodium hydroxide (NaOH). Do this by finding the mass of each element in the compound: 23+16+1 = 40 grams (how convenient, we put exactly one mole into the solution) which means we have NaOH with a concentration of 1mol/dm3.
* Here’s some examples of titration calculations:
  + I want the concentration of 500cm3 of solution, from 40 grams of NaOH
    - We need to convert the 500 to 1000 (we want the answer in terms of dm3) so we do (40/500)\*1000 which equals 80grams.
    - However, 80 grams is the amount of NaOH that would be dissolved in 1000cm3, so then you must do 80/40 which gives 2 moles
    - The answer is 2mol/dm3
  + What mass of H2SO4 is there in 250cm3­ of 1mol/dm3 sulphuric acid (H­2SO­4)?
    - 1dm3 of acid = 1 mole of H2SO4, so the mass of H2SO4 in 1dm3 is simply the weight of its atoms; (2\*1) + 32 + (4\*16) = 98g.
    - Now we do (98/1000)\*250 which equals 24.5 grams
  + 25cm3 of sodium hydroxide (unknown concentration) into pipette. It reacted with 20cm3 of 0.50mol/dm3 hydrochloric acid added from a burette. What was the concentration of the sodium hydroxide solution?
    - The reaction is: NaOH (aq) + HCL (aq) = NaCL (aq) + H2O (l)
    - The concentration of HCL = 0.50mol/dm3
    - (0.50/1000)\*20 = 0.010 moles dissolved in 20cm3 of HCL
    - We did this because 0.50 is the concentration of the known solution, 1000 = 1dm3 and 20 was how many moles of HCL dissolved in the solution.
    - This means that 0.010 moles of HCL reacts with the same quantity of NaOH, so we do (0.010/25)\*1000 (25 is the amount of sodium hydroxide there was and 1000 will convert it to mol/dm3) which gives 0.4mol/dm3.
* **To find the concentration:**
  + **Calculate the mass in grams of solute in 1cm3solution**
  + **Calculate the mass in grams of the solute in 1000cm3 solution**
  + **Convert to moles**
* **To find the mass of solute in a certain volume:**
  + **Calculate the mass in grams of the solute in 1dm3 of solution**
  + **Calculate the mass in grams of the solute in 1cm3 solution (divide by 1000)**
  + **Calculate the mass in grams of solute there is in the given volume of the solution**

## 2.4 – How ideas about acids and bases developed

* Alkalis were originally made from wood ash, which is how they got their name, which is derived from Arabic (al qalīy) which means *calcined ashes*).
* At first, chemists thought that air was responsible for the behaviour of acids, and called the gas oxygen (from Greek *acid forming*).
* Sir Humphrey Davy then linked hydrogen to acids, not oxygen.

## 3.1 – Water and solubility

* You need to know about the **water cycle**; as the water on the surface of the earth is warmed by the sun, it evaporates. As it rises through the atmosphere, it cools and condenses to form clouds, which fall as rain.

[[1]](#footnote-2)

* Many gasses are soluble in water to some extent, and so are many ionic compounds, while many covalently bonded substances just won’t dissolve in water. We call the amount of **solute** that we can dissolve in a certain amount of **solvent** the solubility of that substance. We measure the solubility in grams of solute per 100 grams of solvent at a specific temperature.
* **Saturated solutions** cannot have any more solute put into them, unless heated, in which case more solute can be added, but if the solution is cooled, the solute will precipitate out and crystallise.[[2]](#footnote-3)

## 3.2 – Solubility Curves

### Solubility of solids

* The amount of substance that will dissolve into a solvent is dependent on the temperature of the solvent.
* We can plot how much solute dissolves on a graph with the **solubility on the y axis and temperature on the x axis.**

### Solubility of gasses

* Like solids, the solubility of gasses is affected by temperature, but unlike solids, it is also affected by pressure.
* As the temperature rises, the solubility of a gas in water decreases, but as the pressure increases, the solubility increases. (Think of the fish in streams that are harmed when coal fired power plants use the stream water as a coolant – the water carries less oxygen)
* Fizzy drinks make you burp because the carbon dioxide becomes less soluble as it is warmed in the stomach.

## 3.3 – Hard water

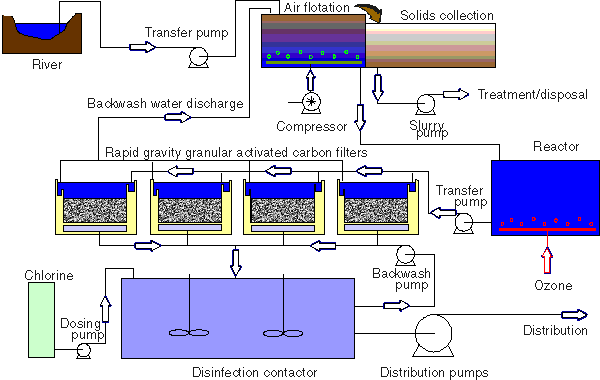
* Hard water makes it difficult to wash as it contains dissolved compounds that react with soap to form **scum**.
* Hard water usually contains Calcium or Magnesium compounds which have been picked up when the water passed over rocks in streams.
* Using hard water is more expensive than using soft water because more soap is needed to give a similar cleaning effect. The compounds in the water must be bonded to first, and then the soap can start to work. Here is the word equation:
  + Sodium Stearate (soap) + Ca2+ and Mg2+ ions (in calcium and magnesium salts) = calcium stearate magnesium stearate precipitate (scum) + Na+ ions (soluble in water)
* Hard water can also lead to scale building up in hot water pipes, here’s the equation:
  + Ca2+ + 2HCO3- (aq) + heat = CaCO3 (s) + H2O(l) + CO2 (g)
* Hard water can reduce the chance of heart disease and the calcium can reinforce teeth.

## 3.4 – Removing hardness

* Soft water doesn’t need to be softened, as it doesn’t contain any Ca or Mg ions. We soften water by removing these ions.
* By adding **sodium carbonate (aka washing soda)** to hard water, we can turn it soft. The carbonate in the washing soda bonds to the calcium ions, and the sodium that was in the washing soda is released instead (along with CaCO3).
* An **Ion Exchange Column** can be used to make water soft. They contain sodium ions, which are swapped for calcium and magnesium ions as the water filters through. Dishwashers contain this system. Eventually, the column runs out of sodium resin and must be replaced.

## 3.5 – Water treatment

* We can make water pure by distillation, here is how water is made safe to drink:



## 4.1 – Comparing the energy produced by fuels

* Fuels produce different amounts of energy when they burn, we can use a **Bomb Calorimeter** to find how much energy is given off. The amount of energy in food is measured in the same way.
* In the classroom, the fuel is burnt, and used to heat up a copper beaker filled with water. Then, the temperature difference is measured.

## 4.2 – Energy change in reactions

* Breaking bonds is an **endothermic** process.
* This means that making bonds is an **exothermic** process. (energy is released when bonds are made).
* In some reactions, the energy released when new bonds are formed is greater than that taken in when the old bonds are broken, so the reaction is **exothermic**
* In other reactions, the energy taken in when breaking the bonds was greater than that given off when the new bonds were formed, making the reaction **endothermic**.
* Energy changes are measured in **kJ/mol**. We can use **energy level diagrams** to view how the energy level of a substance has changed as it has reacted.
* Energy change is represented by **ΔH [[3]](#footnote-4)**
* In an exothermic reaction ΔH is always negative, as the reaction has given out energy, and in an endothermic reaction, the opposite is true.
* Activation energy is the minimum amount of energy needed for the reaction to start.

## 4.3 – Calculations using bond energies

* The energy needed to break a bond between two atoms is known as the **bond energy**.
* As with energy changes, bond energy is measured in **kJ/mol**. Using a list of common bond energies[[4]](#footnote-5), we can work out the ΔH of a reaction.
* Firstly, we need to work out how much energy to break all the bonds, and then we need to work out how much energy we get from forming new bonds.
* To work out ΔH, just do this: total **energy needed to break the bonds – total energy given out when making the bonds.**
* We can deduce certain facts about element from bond energies, for example Nitrogen is unreactive as N2 has a bond energy of 945!

## 5.1 –Tests for positive ions (must be learnt in AQA)

* Lots of compounds look similar to the naked eye, so telling substances apart can be a real problem for chemists. Fortunately, scientists have come up with ways to distinguish one substance from another.

### Flame testing:

* Many metals in group one and two produce a distinctive colour when on fire. To do this:
  + Put a small amount of the suspect compound on a platinum wire loop which has been dipped in concentrated hydrochloric acid.
  + Hold the loop under the blue flame.
  + Use the colour of the blue[[5]](#footnote-6) flame to determine what the metal is.

### Testing using sodium hydroxide:

* Aluminium, calcium and magnesium ions all produce a white precipitate[[6]](#footnote-7) when mixed with sodium hydroxide solution.
  + When lots more sodium hydroxide solution is added, the aluminium compound will dissolve, while calcium or magnesium will not. The latter two can be distinguished by a flame test.
* Some metal ions form coloured precipitates when they react with sodium hydroxide:
  + Copper (II) ions - light blue
  + Iron (II) ions – ‘dirty’ green
  + Iron (III) ions – red/brown
* It’s also used to test if ammonia ions are present, as they react with the sodium hydroxide to form ammonia and water. If this reaction is happening then we cannot tell until the solution has been warmed, in which case ammonia gas will be given off. When red litmus paper is held in the gas, it will turn blue as ammonia is alkaline.

## 5.2 – Tests for negative ions

### Carbonates:

* When a dilute acid is added to a carbonate, it fizzes and carbon dioxide is given off.
* Two metal carbonates give off a very distinctive colour when heated:
  + (Green) CuCO3 -> heat -> (black) CuO + CO2 [[7]](#footnote-8)
  + (White) ZnCO3 -> heat -> (lemon yellow when hot/white when cold) ZnO + + CO2 [[8]](#footnote-9)

### Halides (chloride, bromide and iodide):

* If silver nitrate solution and nitric acid is added to the suspect compound, then the following will happen:
  + Chloride ions – white precipitate
  + Bromide ions – cream precipitate
  + Iodide ions – pale yellow precipitate
* The ionic equation (where X is the halide) is:
  + Ag+ (aq) + X- (aq) -> AgX (s)

### Sulphates[[9]](#footnote-10):

* Sulphate ions produce a white precipitate when we add hydrochloric acid followed by barium chloride solution. The precipitate is barium sulphate.
* The ionic equation is as follows:
  + Ba2 (aq) + SO­42- (aq) -> BaSO4 (s)

### Nitrates:

* We can make use of the test for ammonia to test for nitrates (see *testing for sodium hydroxide* on the previous page).

## 5.3 – Testing for organic substances

* Chemicals that char on heating are **organic**, and ones that don’t are **inorganic**.
* A better definition of organic and inorganic substances is that organic chemicals are substances that are based on carbon.

### Detecting carbon-carbon double bonds (C=C)

* Unsaturated hydrocarbons will react with bromine water producing colourless compounds.[[10]](#footnote-11) This is an easy way to test for (C=C) bonds:
  + Unsaturated hydrocarbon (colourless) + bromine water (orange/yellow) -> no colour
  + Saturated hydrocarbon (colourless) + bromine water (orange/yellow) -> **no reaction, but the colour is orange brown**.

### Combustion analysis:

* We can work out the empirical formula of an organic compound and measuring the amounts of the products formed. E.g.
  + A substance (Z) contains carbon and hydrogen. A sample of Z is burnt in oxygen, and 1.8 grams of water and 3.52 grams of carbon dioxide are produced.
    - Calculate the moles of CO2:
      * The Mr of CO2 is 12 + (2\*16) = 44g
      * 3.54/44 = 0.08 moles (mass/mr)
    - Calculate the moles of H2O:
      * The Ar of H2O is (2\*1) + 16 = 18g
      * 1.8/18 = 0.10 moles (mass/mr)
    - As CO2 has only one carbon atom in, we times 0.08 by 1 = 0.08
    - AsH2O has two hydrogen atoms in, we times 0.10 by 2 = 0.20
    - 0.08:0.20 = 2:5
  + The empirical formula of Z is C2H5

## 5.4 – Instrumental analysis 1[[11]](#footnote-12)

* Modern analysis techniques are:
  + Highly accurate
  + Quicker than normal methods
  + Allow small quantities of chemicals to be analysed
  + Expensive
  + Needs special training to use
  + Liable to provide results that must be compared to known other results to interpret.
* **Atomic absorption spectroscopy (AAS)** is a technique used to measure the concentration of a particular metal in a liquid sample. This is how it works:
  + The sample is fed into a flame and vaporised
  + Light passing through the flame passes through a monochromator which selects a wavelength to be studied.
  + This light falls on a detector that produces an electric current depending on the light intensity. This current is amplified and processed.
  + A measurement and concentration comes out.
* A **mass spectrometer** is able to compare the mass of different atoms and so provides an important way of determining relative atomic masses as well as identifying the atoms.

## 5.5 – Instrumental analysis 2

* **Chromatography** is used to separate out different compounds within a mixture. There are four different types of chromatography:
  + **Gas-Liquid chromatography** – used to separate compounds that are easily vaporised
  + **Gel permeation chromatography** – used to separate compounds according to the size of their molecules
  + **Ion exchange chromatography** – used to separate compounds containing differently charged particles
  + **High performance liquid chromatography** – used to separate compounds in a solution
* Once the compounds have been separated, they can be identified by comparing them to how far known substances travelled, or they can be identified using other instruments.

1. Sorry for the appalling picture [↑](#footnote-ref-2)
2. Sometimes, it’s the other way around. [↑](#footnote-ref-3)
3. Of course you already knew that Δ stands for change and H stands for enthalpy (or total thermodynamic energy). [↑](#footnote-ref-4)
4. Usually, the table lists the energy required to break the bonds, so when you need to look up a value to make a bond, remember to add a ‘-‘ sign and make it negative. [↑](#footnote-ref-5)
5. Or not [↑](#footnote-ref-6)
6. Make it go cloudy [↑](#footnote-ref-7)
7. Check using lime water [↑](#footnote-ref-8)
8. Again, check using lime water [↑](#footnote-ref-9)
9. I think it can be spelt ‘Sulfate’ too [↑](#footnote-ref-10)
10. As you should remember from module one. [↑](#footnote-ref-11)
11. Thankfully, only the pros and cons of the modern techniques are to be committed to memory. [↑](#footnote-ref-12)