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When an acid dissolves in water, it forms $H^+$ ions. This is a hydrogen atom which has lost an electron – in other words, it is a proton. These produced protons become surrounded by water molecules to keep them in solution – we call it hydrated. Hydrated hydrogen ions are shown with $H^+ (aq)$. An alkali is a base which dissolves in water, and produces $OH^-$ ions (hydroxide ions).

Because acids act as a source of protons, we call them proton donors. The hydroxide ions from an alkali combine with protons to form water:

$$OH^- (aq) + H^+ (aq) \rightarrow H_2O (l)$$

And because alkalis behave like this, we call them proton acceptors.

The strength of an acid depends on the extent to which it ionises in water. A strong acid or alkali is one which is 100% ionised in water. Hydrochloric acid, sulphuric acid and nitric acid are all strong acids. Sodium hydroxide and potassium hydroxide are both strong alkalis. A weak acid or alkali is only partly ionised in water. Ethanoic acid, citric acid and carbonic acid are all weak acids; and ammonia solution is a weak alkali.

We can detect strong and weak acids using their pH. This scale is a measure of the concentration of hydrogen ions in a solution.
A strong acid, e.g. hydrochloric, will be completely ionised, so the concentration of hydrogen ions is 1 mol/dm³. However, a weak acid, such as citric acid is only partly ionised, so the concentration of hydrogen ions will be much lower than 1 mol/dm³.

**Titrations**

Adding an acidic solution to an alkaline solution will produce a **neutralisation reaction**. They react together and neutralise each other, producing a salt in the process. When a neutralisation reaction takes place, the quantities of each solution used must be correct, because if a very strong acid and a very strong alkali were mixed, if there was more acid solution, the whole alkali solution would be neutralised, but not all of the acid solution would be — so the mixture would become slightly acidic overall. We can measure precise volumes of acids and alkalis needed to react with each other using **titrations**.

In the neutralisation reaction, the point at which the acid and the alkali have completely reacted is called the **end point**. We can show the end point using a chemical **indicator**. Indicators change colour over different pH ranges. We have to choose suitable indicators when carrying out titrations with different combinations of acids and alkalis:

- **strong acid + strong alkali** – use any indicator
- **weak acid + strong alkali** – use **phenolphthalein**
- **strong acid + weak alkali** – use **methyl orange**

These are the steps to carry out a titration to calculate how much acid is needing to react with an alkaline solution:

1. Measure an known volume of the alkali solution into a conical flask using a **pipette**
2. Add an indicator solution to the alkali in the flask
3. Now put the acidic solution into a **burette**. This long tube has measurements down the side, and a tap on one end and can accurately measure the amount entering the flask. So record the reading on the burette (i.e. starting volume)
4. Open the tap to release the acid solution. The solution from the burette is released one drop at a time, alongside swirling of the flask to ensure the solutions are mixed
5. Keep repeating Step 4 until the indicator changes colour to let you know the acid and the alkali have completely mixed
6. Record the amount of acid you entered by reading the measurement on the burette

Be sure to repeat the entire process two or three times at least to ensure accuracy.
### Calculations Involving Titrations

When talking about concentration, we tend to describe it as the amount of the solute (in terms of moles) dissolved in the solution (in one cubic decimetre), so the units are mol/dm$^3$ so if we know the amount of a substance dissolved in a known amount of solution we can calculate the concentration. For example, imagine we were making a sodium hydroxide solution in water by dissolving exactly 40g of sodium hydroxide to make 1dm$^3$ of solution:

> We know that the mass of one mole of NaOH is the sum of the atomic masses of sodium, oxygen and hydrogen:
> 23 + 16 + 1 = 40g
> Because 40g is in the solution, we know that there is exactly one mole of NaOH in the solution
> And we know that the solution is 1dm$^3$, so the concentration is 1 mol/dm$^3$

The worked examples below are more complicated calculations involving titrations:

<table>
<thead>
<tr>
<th>Worked Example</th>
<th>Worked Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>We use 40g of sodium hydroxide to make exactly 500cm$^3$ of solution</td>
<td>What mass of H$_2$SO$_4$ is there in 250cm$^3$ of 1 mol/dm$^3$ sulphuric acid (H$_2$SO$_4$)?</td>
</tr>
<tr>
<td>Solution</td>
<td>Solution</td>
</tr>
<tr>
<td>We need to calculate the amount of solvent going into 1dm$^3$ of solution</td>
<td>In 1 dm$^3$ of acid there would be 1 mole of H$_2$SO$_4$</td>
</tr>
<tr>
<td>$\frac{40}{500}$ g of NaOH would go in 1cm$^3$</td>
<td>The mass of 1 mole of H$_2$SO$_4$ is: $(2 \times 1) + 32 + (4 \times 16)g = 98g$, so</td>
</tr>
<tr>
<td>$\frac{40}{500} \times 1000g = 80g$ of NaOH would go in 1dm$^3$</td>
<td>in 1000cm$^3$ of solution there would be 98g of H$_2$SO$_4$ and $4 \times 250 = 1000$</td>
</tr>
<tr>
<td>The mass of 1 mole of NaOH is 40g so 80g of NaOH = $80 \div 40 = \text{2 moles}$</td>
<td>So we can do $\frac{98}{4} = \text{24.5g}$</td>
</tr>
<tr>
<td>2 moles of NaOH are dissolved in 1dm$^3$ of solution, so the concentration is 2 mol/dm$^3$</td>
<td></td>
</tr>
</tbody>
</table>
An **exothermic reaction** releases energy. We use exothermic reactions in burning fuels as a source of energy. However, some reactions give off more energy than others, so we can calculate how much energy is released in a given reaction. There is apparatus available to do this called a **calorimeter**. In a school lab, a **simple calorimeter** might be used – but a more accurate instrument is available, called a **bomb calorimeter**.

A bomb calorimeter works by measuring the temperature of the water inside it – because the energy produced in an exothermic reaction increases the temperature of its surroundings, in this case the water. The change in energy is calculated using the temperature change and amount of water (see later on for calculations using energy).

A simple calorimeter however involves very basic apparatus. We don’t use this to measure energy change necessarily because it isn’t very accurate – but we can use it to compare energy changes from different fuels.

When a reaction takes place, bonds are broken and new chemical bonds are made:

- **breaking bonds** is an **endothermic process**, because energy has to be taken in from the surroundings to break the bonds (remember energy is needed to break bonds)
- **making bonds** is an **exothermic process**, because energy is released in the formation of new chemical bonds

Because a reaction makes and breaks bonds, reactions are sort of both exo- and endothermic. For this reason, it is the balance between exo- and endothermic reactions which decides the overall reaction type; for example if more energy is released in the making of new bonds than is taken in to break the bonds, it is overall exothermic – because the exothermic > endothermic.

**Energy Level Diagrams**

We can draw **energy level diagrams** to show energy changes in a reaction. These diagrams show the relative amounts of energy stored in the products and reactants of a reaction, measured in **kJ/mol**.
This is the energy level diagram for an *exothermic* reaction. The products are at a lower energy level than the reactants, so energy has been released as the reactants form the products. In this release of energy, *temperature of the surroundings increases*. In such an exothermic reaction, we say that the change in energy is negative – which we write as $\Delta H$ -ve (see below). This is so because energy is released – so there is less energy in the products than the reactants.

So this is the energy level diagram for an *endothermic* reaction. With an endothermic reaction, more energy is needed to break to bonds of the reactants than is released in forming products. Here, *temperature of the surroundings decreases*. Because the change in energy this time is positive, we say $\Delta H$ +ve (see below).

The Greek letter “delta” (written as $\Delta$) is often used in the sciences and maths to represent change. In chemical energies, we use $\Delta H$ to abbreviate energy change. So $+\Delta H$ means energy increases, $-\Delta H$ means energy decreases.

The amount of energy needed to start a reaction is called the *activation energy* of the reaction. Adding a *catalyst* will significantly reduce this amount of energy (see Rates of Reaction, C2). This in turn increases the proportion of reacting particles which will have enough energy to react. This has many advantages, especially industrially, as it means reactions are more efficient – and the catalysts are economical also.

### Calculating Energy Changes

Looking back at the calorimeters above, when chemicals react and give off/take in energy, we can use calculations to work out exactly how much energy has changed. There is one vital piece of information we need to know to do this...
4.2 joules of energy raises 1g of water by 1°C

Hence the units involved in this energy change will be $\text{kJ/g/°C}$ (kilojoules per gram per degree). A simple calorimeter is used to measure energy change in a reaction $A + B \rightarrow C$. So let’s calculate an example of such a reaction:

Question: 60cm³ of a solution containing 0.1 moles of $A$ is mixed with 40cm³ of a solution containing 0.1 moles of $B$. Prior to mixing, their temperature was 19.6°C. After mixing, the maximum temperature reached was 26.1°C.

1. First, calculate the temperature change:
   
   $26.1°C - 19.6°C = 6.5°C$

2. Since 60cm³ of $A$ added to 40cm³ of $B$ makes 100cm³ overall, we are looking at 100g (assuming the density of the solution is the same as water density). And we know that 4.2J raises 1g by 1°C

3. So energy change = 100g x 6.5°C x 4.2J/g/°C = 2,730J = 2.73kJ

4. BUT – don’t forget the solutions are only 0.1 molar – so we have to multiply our value by 10 to find out a 1.0M solution

   $2.73kJ \times 10 = 27.3kJ$

5. So the final energy change was -27.3kJ

[We know that the temperature increased, so the reaction was exothermic - where energy gets released. That is how we know the energy change will be negative]

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**Bond Energies**

The energy required to break apart a bond between two particular atoms is known as bond energy. Bond energies are measured in kJ/mol and we can use them to work out $\Delta H$ in energy calculations. Some of the most common bond energies are displayed below:
To calculate energy change we need to know: a) the amount of energy needed to break the bonds between the atoms; and b) the amount of energy released in the formation of new chemical bonds.

For example, the bond energy for an H-H bond is 436kJ/mol. This means that the bond energy for forming a new H-H bond is -436kJ/mol.

These two energy level diagrams show H-H bonds being made and broken. The left diagram shows an already bonded H-H bond being broken. This has a bond energy of +436kJ/mol, so we write ΔH = +436kJ/mol on the diagram next to the change in height arrow. The right side is a diagram representing two separate hydrogen atoms bonding. Obviously, this is bond making – which releases energy – so the energy change is -436kJ/mol, written the same way as before, except with a minus sign.

To clarify, the left diagram is endothermic, the right is exothermic.

Making and breaking the same bond always involves the same amount of energy, just different + and – signs.

More Complicated Energy Level Diagram Calculations

Although bond making and breaking is always the same energy levels back and forward – different chemical reactions mix and match the type of bonds being made and broken. This is why energy levels can begin to look a bit more complicated. A good example of a chemical reaction where different bonds are involved is the Haber process. This is the making of ammonia from nitrogen and hydrogen.

Question: Ammonia is made from nitrogen and hydrogen in the Haber process. The balanced chemical equation for this reaction is:

\[ \text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3 \]

Calculate the overall energy change for this reaction.
Solution:

1. Notice we need to break 1 mole of nitrogen and 3 moles of hydrogen. Nitrogen molecules are held together by an extremely strong triple bond (N≡N) with a bond energy of 945kJ/mol. Hydrogen molecules are held together by a single bond with a bond energy of 436kJ/mol.

2. Since there are three moles of hydrogen, the energy needed to break one mole nitrogen and three hydrogen moles is:
   \[ 945\text{kJ} + (3 \times 436)\text{kJ} = +2253\text{kJ} \]

3. We can therefore draw on the first line taking us to our 2 nitrogen atoms and 6 hydrogen atoms, as shown

4. When these atoms form ammonia, NH₃, 6 N-H bonds are made as 2 moles of NH₃ are formed (N-H has a bond energy of 391kJ/mol)
   \[ 6 \times 391\text{kJ} = 2346\text{kJ} \]

5. Because we are making bonds, energy is released, so it becomes -2346kJ

6. We can draw on a second arrow taking us to our two moles of NH₃,

7. We can then calculate the difference
   \[ (+2253) - 2346 = -93\text{kJ} \]

8. So the ΔH = -93kJ
The Sun supplies rivers, lakes and oceans with energy, allowing the water there to **evaporate**. The water vapour formed rises into the atmosphere, where it cools and condenses to form droplets which clouds are made of. Eventually the water droplets fall as rain, replenishing the water sources they originally came from. This is the **water cycle**. Water covers around two thirds of the Earth’s surface, and is absolutely essential to **all** life.

### Solubility

We call the amount of **solute** which we can dissolve in a certain amount of **solvent** the **solubility** of that substance. This is usually measured in grams (of solute) per 100g (of solvent). The most common solvent used is water. Generally speaking, solubility of **solid** solutes increases with temperature. A **saturated solution** is a solution in which as much solute as possible has been dissolved. Heating the solution will allow more of the solvent to be dissolved until it becomes saturated again. Of course, this means when the hot saturated solution cools, some of the solute will have to come back – so it **crystallises** back out of the solution.

We can show the different amounts of solute which will dissolve into solution at different temperatures using special graphs called **solubility curves**. These can be used to a) predict how much solute will dissolve into a solvent at any given temperature, and b) predict how much solute will form again when we cool down a hot solution.

The solubility curves for potassium nitrate, sodium nitrate and sodium chloride are shown here. As you can see, the solubility of each one increases with temperature as the rule states – but the rate of increase differs between solutes. As you can see, sodium chloride barely increases in solubility between 0°C and 100°C, whereas potassium nitrate increases eightfold in the same period.

The thing that all of these solutes have in common is that they are all solid solutes. The solubility of gases works in exactly the opposite way – as temperature increases, the amount of solute which will dissolve into solution **decreases**. However, **pressure** is another
factor affecting the solubility of gases. So gas solubility only decreases with temperature as long as the pressure is kept constant – but if temperature is kept constant – solubility of gases increases as pressure increases.

The solubility curve here (note: solubility curves may be straight lines) shows the solubility of oxygen in water at 10°C. Temperature has to be kept the same here, because if temperature is not kept constant, it has a knock-on effect of the changing solubility. We measure the pressure in atmospheres.

The Importance of Solubility

There are a large number of reasons why we need to know about the solubility of solvents. For example, rivers, lakes and reservoirs are contaminated by chemical fertilisers on crops which are dissolved by rainwater – so the nitrate levels of the water supplied to our homes and schools have to be carefully monitored by the companies that supply the water.

Also, dissolved oxygen in water is what keeps the animals living in water alive. Whilst power stations which only pump hot water into rivers rather than waste are seen as environmentally friendly, is this necessarily true? Think about it – this increases temperature of the rivers and lakes, making less oxygen able to dissolve in it, so more animals living in the water die. We call this thermal pollution. It is especially bad for fish.

Water Hardness

Believe it or not there are different types of water coming through our taps in different areas of the country. In certain places, when we wash with soap, the water forms a rich lather easily, but in others it is more reluctant to – this is because the water is hard. Hard water makes it more difficult to wash, but also, more difficult to clean the bath or sink when finished, because hard water contains dissolved substances which react with soap to form scum. The scum floats on the water and will stick to the bath.

Generally speaking, hard water tends to contain calcium and magnesium compounds. These get dissolved into the water or riversstreams when they run over rocks containing these elements. For example, limestone which contains calcium carbonate, gets dissolved in water droplets – which even
makes the water slightly acidic. The water is then taken to reservoirs, and forwarded to our homes. It is the dissolved substances which react with hard water to form scum.

In terms of economic factors, hard water is more expensive because more soap is required for the same wash. The soap reacts with the magnesium and calcium ions in the water, forming salts called stearates (the chemical name for scum). Only after all of the calcium and magnesium have reacted can the soap begin to form a lather – and this is why so much more soap is needed per wash.

The below simple equation shows the ion exchanges in the reactions:

\[
\text{sodium stearate (soap)} + \text{Ca}^{2+}\text{ and Mg}^{2+}\text{ ions} \rightarrow \text{calcium stearate precipitate (scum)} + \text{Na}^{+}\text{ ions soluble in water}
\]

Scum is not the only problem hard water can potentially cause. Pipes can suffer from scale (also limescale). Scale is also common in heating elements and other parts of our hot water system. Pipes which have a lot of scale eventually block up and stop functioning. The same problem occurs in kettles. When it happens to them, they tend to become less energy efficient – by being more slow and heating to lower temperatures – because scale is such a poor conductor of heat.

The simple equation below shows limescale formation chemically:

\[
\text{Ca}^{2+}(aq) + 2\text{HCO}_3^-(aq) \xrightarrow{\text{heat}} \text{CaCO}_3(s) + \text{H}_2\text{O}(l) + \text{CO}_2(g)
\]

**Advantages of Hard Water**

There are certain attributes to hardness of water which are good. Whilst the dissolved substances in hard water make it harder for us to wash, form scum on our baths and sinks, and damage our pipe work, those same dissolved substances are very good for our health. Calcium ions in drinking water are good for making stronger bones and healthy teeth. It is also strongly probable that hard water reduces the risk of heart disease in the people who regularly drink it.

**Removing Hardness**

The other type of water of course does not contain the dissolved substances that cause scum and scale. This is soft water. We soften hard water by removing the calcium and magnesium ions in the hard water. This benefits us in terms of washing our bodies, our
clothes and heating our water, but of course, we are advised to continue drinking hard water for health purposes. Industry is another reason to soften water, there are a number of manufacturing concerns related to hard water. There are two main ways to soften water...

The first method is to use washing soda (chemical name sodium carbonate). All we do is add washing soda to the hard water. When added, it precipitates out calcium and magnesium ions as insoluble carbonates (see below). Once these ions which cause the hardness are no longer in the solution, they cannot react with soap (to form scum, etc). This means the water has become soft.

\[ \text{Ca}^{2+} \text{(aq)} + \text{CO}_3^{2-} \text{(aq)} \rightarrow \text{CaCO}_3 \text{(s)} \]

Method number two is to use an ion-exchange column. This also removes the calcium and magnesium ions from the hard water. They are columns containing sodium ions which are exchanged for the calcium and magnesium ions in hard water when it is passed through. This is good for homes, where the columns can be fitted in certain places to ensure all hard water used for washing ourselves, our clothes and providing showering water is softened (also, dishwashers contain their own water-softening system similar to this the majority of the time). Once the sodium ions have been exchanged for the calcium and magnesium ones, the column is washed with a salt solution to exchange the dumped calcium and magnesium ions with more sodium ions, ready for the next time. For this reason, water-softeners must be topped up with salt (sodium chloride) every so often to keep it functioning.

#### Health & Water

Because of the health benefits associated with drinking hard water, it is a requirement (for health reasons) for houses with water-softening systems to be fitted with at least one cold tap with a supply of hard water in the kitchen.

#### Water Treatment

There are so many uses for water for everyone all around the world. In a developed country such as our own, we always need to treat the water to make sure it’s clean and safe. A simple way of treating water is to use a filter jug around the home. This is fitted with a filter, so that as you pour water in, it runs through the filter, “purifying” it. This normally contains activated carbon, an ion-exchange column and some silver. Firstly, the carbon in the filter reduces the levels of chlorine, pesticides and other organic impurities. Then the ion-exchange column removes the dissolved substances as described earlier, including calcium, magnesium, lead, aluminium and copper ions. Finally, the silver (fitted in most, but not all filters) discourages bacterial growth in the filter, keeping it clean and functioning. However, the filter usually needs changing every few weeks.
Even after passing through a filter, water does not become pure. There are still dissolved impurities in it. However, the water is still definitely cleaner after going through a filter – and it is always sensible to filter water you are putting into a kettle to be boiled, as it reduces the limescale build up.
The Earlier Tables of the Elements

One of the first proper stabs at developing a table of the elements was by John Dalton. By testing the known elements’ reactions he arranged them all in order of mass. The image shows Dalton’s finalised table. As different as it looks to the modern table, it actually bears some resemblance! John Newlands built on Dalton’s table using the law of octaves, suggesting that every eight element had similar properties (now you can see the similarity).

The problem with Newlands’ table was that he was too determined to get it done and working that he made some mistakes. What he didn’t know was that there were still many elements to be found, so he filled in octaves regardless of their properties, and some of them ended out not being similar at all. He did this to make everything fit in, so as a result, his ideas were not accepted.

On the bright side, a year prior to Newlands’ attempt, the French chemist Alexandre-Emile Beguyer de Chancourtois had a better go at arranging the elements by ordering them by properties (in sections of eight again). He successfully made a very clever diagram table of the elements. Unfortunately, the diagram was missed out when his work was published!

Towards the late 1860s, things started to look better though, because the Russian scientist Dmitri Mendeleev created a very promising table. By this time, fifty elements had been discovered, and Mendeleev arranged them all in order of atomic mass. After that, he arranged them into further groups based on their behaviour and properties, to arrive with a periodic table. He is considered to be “the father of the modern periodic table.” A minor problem with this table was that there had to be a few blank spaces – although to be fair at least he recognised that certain elements weren’t discovered by looking at his table – he didn’t try to cram them all in like Newlands.

The Modern Periodic Table of the Elements

A problem faced by Mendeleev was that although arranging elements in order of atomic mass produced groups of elements which behaved the same way, not all elements did do this. For example (look at a modern table for help here), argon (Ar) has a higher atomic mass...
than potassium (K) and so would be placed after potassium with the reactive metals – but argon is a noble gas! Therefore, argon was put before potassium, even if the argon atom was heavier.

In the early 20th century, scientists began to look more closely at the atomic structure and decided that the way to solve problems like the one described above was to arrange all the elements in order of atomic number (or proton number). This put them in exactly the right order, and they all are in their correct groups and periods. You can tell this nowadays because of the way you know how certain elements behave based on electrons (depending on what group they are in).

So the modern periodic table today is now a reliable model which we can use for many functions, split into groups and periods.

### Reactivity Groups

You should know that an element’s placing in a group tells you something about its reactivity, but the rules differ between the groups. The reason they all have similar properties is because atoms in one group all share the same number of electrons in the outer shell (energy level).

Within a group, the properties of elements are affected by the number of energy levels beneath the outer one. As you go down a group, the size of the atoms gets bigger – this is because the number of energy levels from the nucleus has increased. Because the energy levels are further away from the positive nucleus as you go down the group:

- the larger atoms GAIN electrons LESS easily
- the larger atoms LOSE electrons MORE easily

### Group 1 – The Alkali Metals

We call the first group (Group 1) the alkali metals. These are lithium, sodium, potassium, rubidium, caesium and francium. The first three of those elements being reactive, you may experiment with their reactions in class, but the last three of those elements are extremely reactive, you won’t get to play about with them in class. Francium is even radioactive.

All of the alkali metals are reactive, though, and have to be stored in oil to stop them reacting with oxygen in the air. Reactivity increases as you go down the group. Lithium is therefore the least reactive, and francium the most. This is because, as the atoms get
bigger, the single atom on the outer energy level (see the diagrams on the right) gets further away from the nucleus and becomes less attracted to the positive nucleus. Therefore, it is easier for that electron to be lost to another atom.

Alkali metals have quite a low density, in fact, lithium, potassium and sodium will float on water. They are all also very soft, so can be cut with a knife. As for appearance, they are shiny, metallic, however, when they react with oxygen, they turn a dull colour as they form a layer of oxide.

They also melt and boil at fairly low temperatures compared to other metals. **Melting and boiling points decrease as you go down the group.**

Alkali metals can only form ionic compounds. This means that they can only bond with non-metals (see Atoms and Bonding, C2). They always react by losing their sole outer electron and formed +1 ions.

These metals react with water, some more violently than others. Adding sodium, lithium or potassium will cause the metal to float on the water and move around — fizzing, because the metal reacts with water to form hydrogen gas. Potassium reacts more violently than the other two, because it is so vigorous the hydrogen actually catches fire and burns with a lilac flame. The other three metals react insanely with water. Whenever an alkali metal reacts with water, it also forms a metal hydroxide. These are soluble in water and give the solution a high pH — alkaline, hence the name “alkali metals.”

It is not just water that these alkali metals will react with. Other non-metals such as chlorine will react with them. When they react with chlorine, chlorides are produced (white solids). They dissolve into the solution and make a colourless solution. Again, the reactions get more and more vigorous as they go down the group. You will already know the equation:

\[
\text{sodium + chlorine} \rightarrow \text{sodium chloride}
\]

The alkali metals react similarly with fluorine, bromine and iodine.
Group seven in the periodic table contains the **halogens** - poisonous non-metals which all have coloured vapours. They are relatively typical non-metals, in such that they are poor conductors of heat and electricity, and that they have low melting and boiling points. At room temperature, fluorine is a poisonous, yellow gas, whilst chlorine is a green poisonous gas. All of them are reactive.

All of the halogens exist in molecules, pairs of atoms. These molecules are made from a **covalent bond** (see Atoms and Bonding, C2), and so we call the type of structure a **diatomic molecule**. Halogens are the opposite to alkali metals, they don’t have one spare electron – they have one **missing** electron. This is because their outer energy level contains seven electrons, and they need one more to complete the shell. This means halogens do both ionic and covalent bonding. Again, in contrast to the alkali metals – electrons will be **less** attracted to halogens further down the group because they are further away from the nucleus, so there is less attraction between the negative electron and the positive nucleus. For this reason, **reactivity decreases as you go down the group**. Look at the diagram on the right to see this.

Halogens all react with metals. They gain a single electron to complete their outer energy level, so their ions are of a -1 charge. In these reactions, ionic salts (called **metal halides**) are formed. Examples of these include sodium chloride, iron(III) bromide and magnesium iodide.

When halogens react with other non-metals, they have to share electrons to gain a stable electronic configuration – so this is covalent bonding. An example of a covalent compound formed this way is hydrogen chloride (not to be confused with hydrochloric acid), HCl.

This **dot cross diagram** shows how the electron on the outer shell of a hydrogen atom shares to fill empty spaces on both atoms. This is covalent bonding.

This dot cross diagram shows an ionic bond between a halogen (chlorine) and an earth metal, calcium. Calcium, being in Group 2, has two electrons free on its outer shell, which can be lost to other atoms to
complete their outer shell in ionic bonding. In this case, the two electrons have gone to the chlorine – one electron to each chlorine atom, because chlorine atoms only need one extra electron each to complete their outer energy level.

The halogens’ melting and boiling points vary drastically. Melting and boiling points increase as you go down the group.

We can use a more reactive halogen to displace a less reactive halogen from a solution of its salt. For example, bromine displaces iodine from its solution because bromine is more reactive, and chlorine will displace both iodine and bromine. Chlorine will displace bromine if we bubble the gas through a solution of potassium bromide:

\[
\text{Cl}_2 + 2\text{KBr} \rightarrow 2\text{KCl} + \text{Br}_2
\]

Obviously, fluorine, the most reactive halogen, would displace all of the other halogens, but since its reaction with water is so incredibly violent, we cannot do displacement reactions involving fluorine.

## The Transition Elements

The large block of metallic elements in the centre of the modern periodic table is what we call the transition elements, or sometimes the transition metals.

Most of these elements have similar physical properties to each other, yet their properties differ to other elements around the table. Their typical metallic structure explains the majority of their properties. They have giant structures held together by metallic bonds, with free roaming electrons (see Atoms and Bonding, C2). They are good conductors of heat and electricity, like all metals, this is because delocalised electrons carry the current/heat energy around the metal. They are also quite strong metals, but malleable. With the exception of mercury (which is liquid at room temperature), they all have very high melting and boiling points (usually ranging between 1200°C and 2000°C).
In the transition elements, a lower energy level (or inner shell) is filled up between Group 2 and Group 3. This partly-filled lower energy level explains why transition metals form brightly coloured compounds and results in their use as catalysts.

These metals are far less reactive than the alkali metals (Group 1). They do not react as readily with oxygen or water, i.e. they corrode more slowly. Having these chemical properties combined with the physical properties described above makes the transition metals very suitable for structural materials. They are especially useful when mixed together with each other or other elements to make alloys (see Rocks and Metals, C1). Iron mixed with carbon in steels is a good example of an alloy. Other useful alloys of transition elements include brass (a combination of copper and zinc) and cupro-nickel - a very hard alloy of copper and nickel which is used to make coins in British currency.

Many of the transition metals form coloured compounds. For example:

- potassium dichromate(VI) is orange, where the orange colour is there due to the presence of chromium ions
- copper(II) sulphate is blue, from copper ions
- and potassium manganate(VII) is purple, from manganese ions

The colours produced play a part in our lives. For example, the colours of rocks, minerals and gemstones are the direct result of transition elements’ ions. A red-brown colour in a rock usually means there are iron ions present. Likewise, the blue in sapphire and the green in emerald is due to the ions in the crystal-structure.
Flame Tests

Identifying Group 1 and Group 2 metals is a piece of cake when we burn them, as they tend to have unique flames which we can associate with the elements. We call these tests **flame tests**. We perform a flame test by putting a small amount of a compound to be tested in a **platinum wire loop** which has been dipped in hydrochloric acid, and then we hold the substance over a blue Bunsen flame. The flame should show a particular colour which can be used to identify the unknown substance.

- lithium will burn a bright red flame
- sodium will burn a golden yellow flame
- potassium will burn a lilac flame
- calcium will burn a brick red flame
- barium will burn a green flame

Testing for Positive Ions

Another test for unknown substances is to test the reactions with **sodium hydroxide solution**. Aluminium, calcium and magnesium ions all form a **white precipitate** when they react with sodium hydroxide (NaOH). Hence adding it, if a white precipitate forms, we know it’s one of those three. To find out which one, we can add more and more NaOH – because eventually aluminium ions dissolve in it. If it dissolves, it’s aluminium, otherwise it’s either calcium or magnesium. To find out which of those two it is, we can use their flame tests – calcium burns with a **brick red** flame, whereas magnesium produces no special flame.

Other metal ions produce coloured precipitates when they have NaOH added to them. Adding sodium hydroxide solution to:

- a substance with copper(II) ions produces a light blue precipitate
- a substance with iron(II) ions produces a “dirty” green precipitate
- a substance with iron(III) ions produces a red-brown precipitate

As well as this, NaOH can be used to detect if ammonium ions ($\text{NH}_4^+$) are present in an unknown substance. Ammonium ions react with NaOH to form ammonia and water:

$$\text{NH}_4^+ (aq) + \text{OH}^- (aq) \rightarrow \text{NH}_3 (aq) + \text{H}_2\text{O} (l)$$
To test for ammonium ions, we add NaOH to a solution of an unknown substance. If ammonium ions are present, ammonia (as well as water) forms. When we warm the solution, ammonia is then given off as a gas. We can detect ammonia gas using damp red litmus which should turn blue as ammonia is an alkaline gas.

**Testing for Negative Ions**

There are a number of different tests we can use to detect negative ions. Each type of negative ion has its own test...

***Testing for carbonates***: If we add a dilute acid (e.g. hydrochloric) to a carbonate, it fizzes and produces carbon dioxide gas. We can test for carbon dioxide gas using limewater, and if it fizzes and produces the gas, we know it’s a carbonate. There are two particular metal carbonates which are giveaways, however, which makes it slightly easier to detect them. For example, copper carbonate is a green substance which when heated decomposes to give black copper oxide and carbon dioxide. Also, when zinc carbonate, a white substance, is heated, it forms the lemon-yellow zinc oxide and carbon dioxide. These two reactions are shown below:

\[
\text{CuCO}_3 (s) \rightarrow \text{CuO} (s) + \text{CO}_2 (g)
\]

\[
\text{ZnCO}_3 (s) \rightarrow \text{ZnO} (s) + \text{CO}_2 (g)
\]

***Testing for halides***: When we add dilute nitric acid and silver nitrate solution to an unknown solution, the appearance of a precipitate tells us what halide ion is present. Chloride ions give a white precipitate; bromide ions give a cream precipitate and iodide ions give a pale yellow precipitate. The ionic equation for this is, where \(X^-\) is the halide ion:

\[
\text{Ag}^+ (aq) + X^- (aq) \rightarrow \text{AgX} (s)
\]

***Testing for sulphates***: Adding hydrochloric acid followed by a barium chloride solution to sulphate ions in solution produces a white precipitate (barium sulphate, an insoluble salt). The ionic equation for this is shown:

\[
\text{Ba}^{2+} (aq) + \text{SO}_4^{2-} (aq) \rightarrow \text{BaSO}_4 (s)
\]

***Testing for nitrates***: The test for ammonia (see Testing for Positive Ions above) is used again here. We add sodium hydroxide to a solution of the unknown substance and gently warm it. If no ammonia is detected, we add some aluminium powder. This reduces the nitrate ions to ammonium ions. These react with the sodium hydroxide to produce ammonia gas, which is given off. This is detected using damp red litmus which will turn blue.
Testing for Carbon=Carbon Double Bonds

200 years ago, the Swedish scientist Jöns Jakob Berzelius decided to categorise all substances depending on their behaviour when heated. Chemicals which burned or charred when heated came mainly from living things, so were called organic substances. Other substances melted or vaporised when heated, and returned to their original state when cooled – these were inorganic substances. However, nowadays we refer to anything “organic” as based on the element carbon.

Unsaturated hydrocarbons will react with bromine water to give a colourless compound (see Crude Oil, C1). Unsaturated hydrocarbons contain carbon-carbon double bonds (C=C), so this is a good test to detect them. It is the basis for detecting C=C bonds in unsaturated oils and fats. If we want to do this, the oil is titrated against an iodine solution (iodine solution reacts in the same way with C=C bonds as bromine water). The iodine number is then calculated – based on the number of iodine molecules needed to react with all of the C=C bonds in one molecule of fat.

Combustion Analysis

We can work out the empirical formula (see Chemical Calculations, C2) of an organic compound by burning it and measuring the amounts of the formed products. For example, an organic substance A contains hydrogen and carbon. A sample of A is burnt in an excess of oxygen, producing 1.80g of water and 3.52g of carbon dioxide. To work out the empirical formula for A we would:

1. Firstly, calculate the moles of carbon dioxide:
   The relative atomic mass of carbon dioxide is 12 + (2 x 16) = 44g
   Amount of carbon dioxide = 3.52 ÷ 44 = 0.08 moles
2. Then calculate the moles of water:
   The relative atomic mass of water is (2 x 1) + 16 = 18g
   Amount of water = 1.80 ÷ 18 = 0.1 moles
3. Each molecule of carbon dioxide formed requires one carbon atom from a molecule of A. So for every mole of carbon dioxide formed, A must contain one mole of carbon atoms
   Amount of carbon atoms in sample of A = 0.08 moles
4. Similarly, every molecule of water formed requires two hydrogen atoms from A. So for every mole of water formed, A must contain two moles of hydrogen atoms
   Amount of hydrogen atoms in sample of A = 0.1 x 2 = 0.2 moles
5. So A contains carbon atoms and hydrogen atoms in the ratio: 0.08 : 0.20 = 2:5
6. The empirical formula of A is then C₂H₅
Instrumental Analysis

There are a large number of reasons why instrumental analysis is a preferred method of analysis. The ability to quickly, and still accurately, check products in industry to make sure there are no contaminated substances or by-products in manufacturing is a particularly beneficial advantage. The analysis of compounds is especially significant in terms of health (e.g. kidney dialysis machines can build up dangerous levels of aluminium in the water, so the water is often monitored to check the aluminium content is low).

The recent boom in terms of electronics and computing has aided the progression of analysis, and new industrial methods have been developed. Instrumental methods are preferred in general because they are highly accurate, quicker, and enable tiny quantities of chemicals to be monitored.

However, the main disadvantages of these methods include it being very expensive, people with special skills and a lot of training need to do it and it only gives results which can be compared to those we already have.

Detecting Elements Using Instrumental Analysis

There are two main ways we can instrumentally detect and identify elements. The first is atomic absorption spectroscopy (AAS), which is a technique used to record the concentration of a particular metal in a liquid. The liquid is fed into a flame so it vapourises, and light passing through the flame passes through a monochromator which chooses the wavelength to be studied. The light then falls on a detector where an electric current is produced, whose strength depends on the light intensity. This current is then analysed in the system’s electrical circuits – and a measurement of the metal’s concentration in the liquid is provided.

liquid → flame → vapour → monochromator → detector → current → circuit

The other main method is using a mass spectrometer, to compare the mass of different atoms. This provides a way of determining relative atomic masses as well as identifying particular elements in a sample. The diagram below shows how this works:
Also see UV-visible spectroscopy and nuclear magnetic resonance spectroscopy]

**Chromatography**

An example of a simple technique used to separate compounds within a mixture is chromatography. This process determines the separate compounds based on how well they dissolve in a particular solvent. We can tell their solubility based on how far they travel up a piece of chromatography paper. Here are the different methods of chromatography available:

- gas-liquid chromatography to separate compounds which are easily vapourised
- high performance liquid chromatography to separate compounds in solution
- gel permeation chromatography to separate compounds according to the size of their molecules
- ion-exchange chromatography to separate compounds containing differently-charged particles

After the substance has been tested, the results are compared to already known chemicals to assess what they might be. If those are not available, an alternative is to use more technical instruments to analyse the data.

**Gel Electrophoresis**

A technique called gel electrophoresis is used to analyse DNA. This technique is used for commonly called “DNA tests.” They can be used to find out if people are related or to identify if a suspect was at a crime scene by the police.
Chemical Analysis

Chemical analysis can be used in society for a number of reasons. One of these reasons is to stop doping (performance-enhancing drugs being used in sport). Chemical analysis can be done to test if any of these have been taken by athletes. Athletes are advised to ask a doctor about any medication, no matter how normal it looks, in case it contains these agents and may be accused of cheating.
End of Unit Questions

C3-1 : Acids & Bases

1. Explain the term “hydrated” in terms of hydrogen ions
2. What are proton donors and proton acceptors?
3. Which acid is stronger: pH 6.0 or pH 3.4?
4. What is the point of a titration at which the acid and the alkali have completely reacted?
5. When would you use a phenolphthalein indicator?
6. When would you use a methyl orange indicator?
7. Describe the titration process, including the specialist equipment required

C3-2 : Energy Calculations

1. What is the function of a bomb calorimeter?
2. Explain the difference between an endo- and exothermic reaction
3. Describe the energy changes in an endo- and exothermic reaction
4. What is energy measured in?
5. What is the standard denotation for change in energy?
6. Define activation energy
7. Draw energy level diagrams to show a H-H bond breaking and making

C3-3 : Water & Solubility

1. Describe the processes involved in the water cycle
2. What are: solute, solvent, solution and solubility?
3. Explain what makes a solution saturated
4. What does a solubility curve show?
5. Name two factors affecting solubility
6. Explain how we can use solubility curves
7. How does the solubility of gases differ to solubility of solids
8. Why do we need to know about solubility?
9. What makes water hard?
10. Explain the difference between scum and scale
11. Describe the formation of scum and scale
12. Describe one advantage of drinking hard water
13. How can we make water soft using an ion-exchange column?
14 What can washing soda be used for?
15 Describe one method of water treatment

C3-4 : The Development of the Periodic Table

1 Describe Dalton, Beguyer and Mendeleev’s tables
2 Why was Newlands’ table useless?
3 How was the table developed into the modern periodic table?
4 Why are Group 1 metals called the alkali metals?
5 Why do they have to be stored in barrels under oil?
6 How is reactivity affected as you go up/down the group?
7 Describe some of the halogens’ chemical properties?
8 What is a metal halide?
9 What are some of the physical properties of the transition metals?
10 How is mercury different to other elements in this group?
11 What is an alloy?
12 Describe brass and cupro-nickel

C3-5 : Chemical Analysis

1 Describe the flame tests for lithium, sodium and potassium
2 Explain in detail the test for ammonia
3 Describe the test for: carbonates, halides, sulphates and nitrates
4 What is an organic substance?
5 How would you test for C=C bonds?
6 Name and explain two forms of instrumental analysis