

GCSE Chemistry

2020 Exam



Ammonium cerium (IV) sulfate
 $(\text{NH}_4)_4\text{Ce}(\text{SO}_4)_4 \cdot 2\text{H}_2\text{O}$

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Kinetic Theory

Kinetic theory is the theory that says “everything is made up of particles, and, unless at absolute zero the particles are constantly moving.”

Kinetic theory helps us to explain why diluting a coloured solution makes the solution lighter in colour, and also that gas or liquid particles move by diffusion.

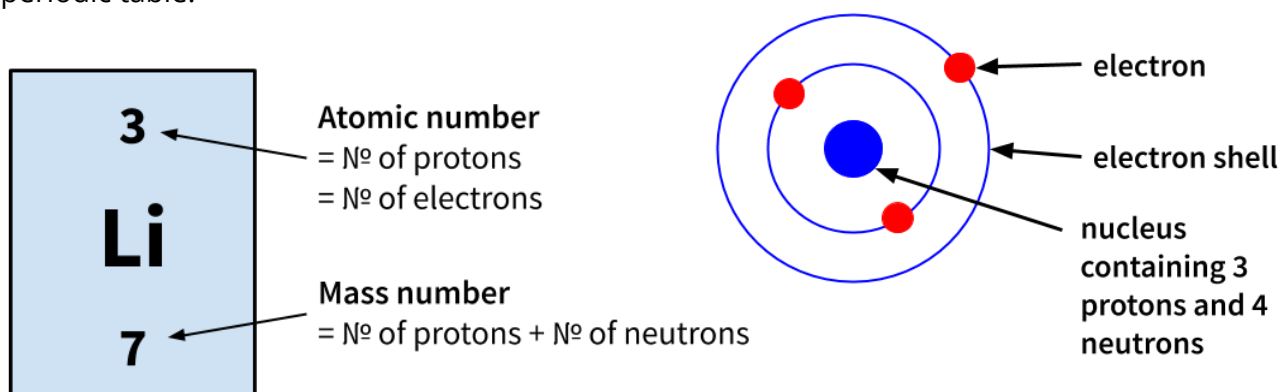
In chemistry we are particularly concerned with two types of particles:

1. Atoms the smallest particles that can be identified as a particular element
2. Molecules two or more atoms chemically bonded together

Atoms

The idea of atoms is based on a model proposed by Rutherford. He fired positively charged particles at gold foil. Most of the particles passed straight through, showing that most of a gold atom is empty space. A few particles turned around and came back, showing that they had hit something small and positive.

We now have a well established model for atomic structure. This is a drawing of a lithium atom. It has three protons, four neutrons and three electrons. This information can be found on the periodic table.



Protons and neutrons are found inside the nucleus and make up most of the atom's mass. Protons and neutrons have approximately the same mass, while electrons are lightweight. Electrons are arranged in rings; they orbit the nucleus and are negatively charged. Protons are positively charged. Neutrons, as per the name, are neutral and have no charge.

Molecules

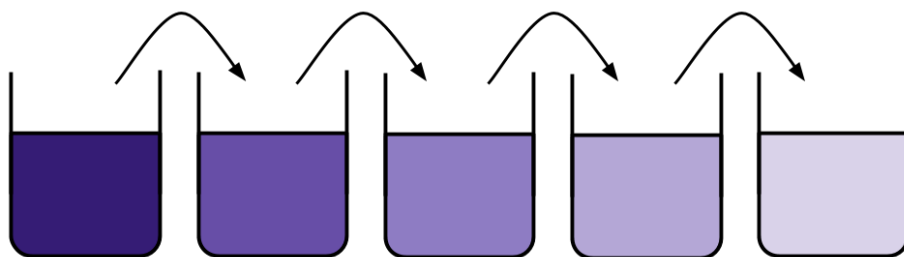
A molecule is two or more non-metal atoms that are chemically bonded together, not necessarily the same elements. For example, liquid water from your tap is made of lots and lots of water molecules. Each water molecule is made of two hydrogen *atoms* and one oxygen *atom*. A compound is a type of molecule, but not all molecules are compounds.

Proving the existence of atoms

Atoms and molecules are generally too small to be seen. We can show that they are there and exist through a series of experiments.

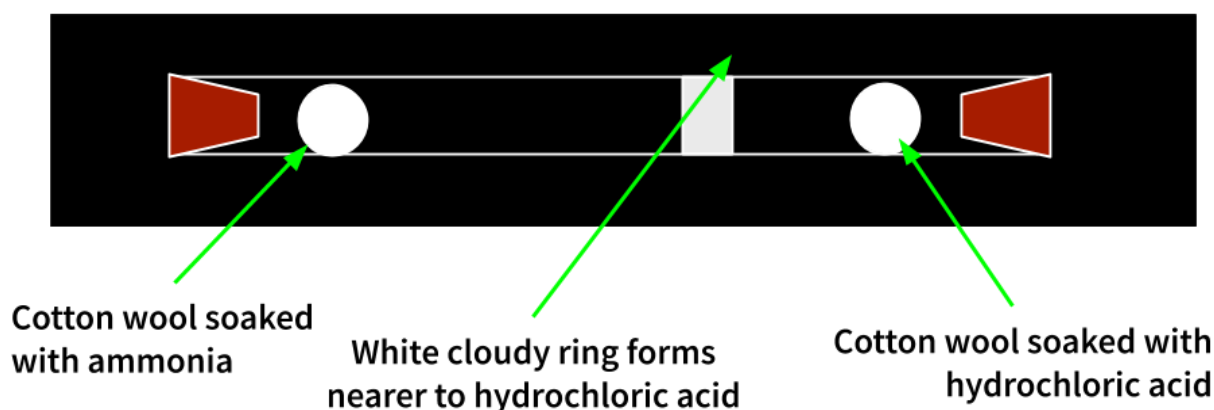
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Dilution of potassium permanganate



The potassium permanganate has invisible purple particles. We cannot see them but we can see their colour. A small amount of potassium permanganate crystals are able to give colour to a large amount of water, because the potassium permanganate breaks down to individual particles when dissolved in water. Since even the last beaker is somewhat coloured, we can say that these individual particles must be very small.

Diffusion of gases



The cotton wool soaked in ammonia provides ammonia gas. The cotton wool soaked in hydrochloric acid provides hydrogen chloride gas. When we place the gases at opposite ends of the tube the gases spread out. A white ring, marking where the diffusing gases meet. The white solid (ammonium chloride) forms closer to the hydrochloric acid end, since ammonia gas particles are lighter and travel faster than the hydrogen chloride.

Compounds

A compound is formed when atoms of different elements react together, forming a molecule. All chemical compounds can be represented by chemical formulae. A chemical formula tells us which atoms are present in a compound. For example:

CO_2 1 carbon & 2 oxygen atoms
 H_2O 2 hydrogen & 1 oxygen atom
 NaCl 1 atom of sodium & 1 atom of chlorine

How do we know what the formula is? The answer lies in something called valencies. Valencies are on the back of your periodic table. Valencies need to match in a formula.

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Writing Chemical Equations

All chemical reactions can be represented by chemical equations. Word equations are common in prep school. For example,

Aluminium + oxygen → aluminium oxide

At GCSE we need to write this as a balanced symbol equation. This happens in two steps, as follows.

1. Replace words with symbols using valencies on the back of your periodic table
 - a. $\text{Al} + \text{O}_2 \rightarrow \text{Al}_2\text{O}_3$
2. Put BIG numbers to balance all the atoms
 - a. $4\text{Al} + 3\text{O}_2 \rightarrow 2\text{Al}_2\text{O}_3$

Any chemical that ends in -ide USUALLY has only two elements.

e.g. sodium iodide



e.g. magnesium oxide



e.g. aluminium sulfide

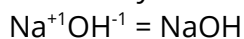


The only exception to this rule is hydroxide.

Formula OH

Valency 1

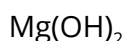
e.g. sodium hydroxide



BUT magnesium hydroxide

MgOH_2 is **incorrect**

We need brackets, like so



Many other compound names end in “-ate”. This ending means oxygen is present as well as another element. For example:

Sodium sulfate = Na_2SO_4

Magnesium carbonate = MgCO_3

Potassium nitrate = KNO_3

Aluminium phosphate = AlPO_4

Highlighted in green are the “-ate”s. We need to learn the following ones:

NO_3 nitrate valency 1

CO_3 carbonate valency 2

SO_4 sulfate valency 2

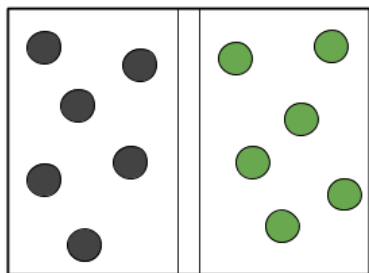
PO_4 phosphate valency 3

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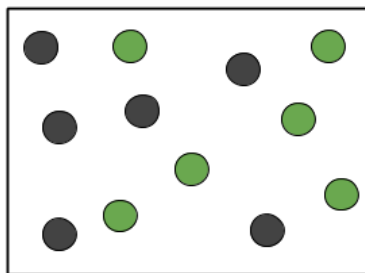
Diffusion

Diffusion is the net movement of particles in **fluid** from one region to another. It occurs because of two properties of fluids: the particles have space between them and can move about.

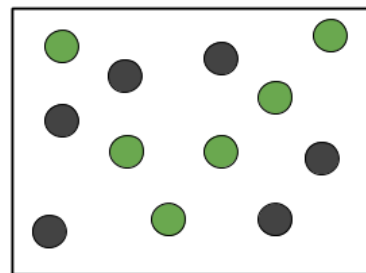
If we start off with two types of fluid which are separated, and then remove the barrier which is separating them then diffusion will occur. The particles from both regions will spread out and move in the spaces between other particles. The particles in each fluid will continue to mix until they are evenly distributed. Diffusion cannot happen in solids because their particles cannot move around.



1. The two types of particle are separated by a barrier.

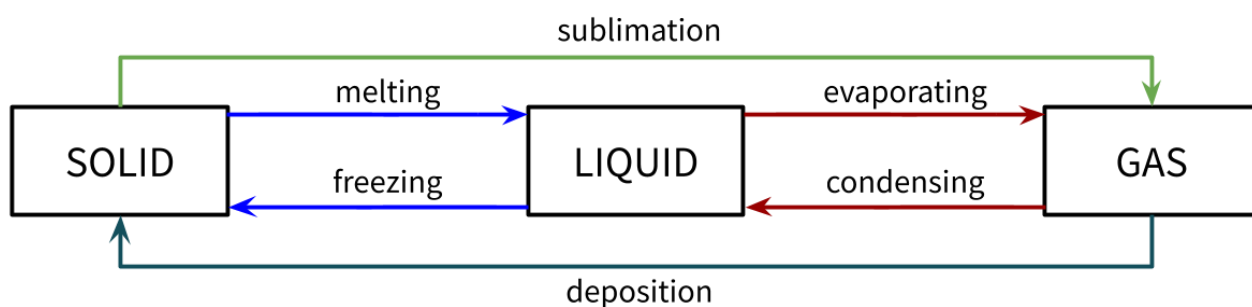


2. The barrier is removed and the particles have begun to mix.



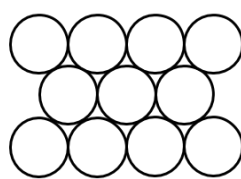
3. The particles are evenly distributed

Changes of State



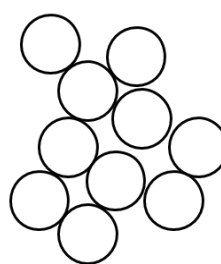
What is the difference between boiling and evaporation?

A liquid only boils if its temperature is equal to or is greater than its boiling point; bubbles form in the body of the liquid as the gas is produced. In evaporation, only surface particles that have sufficient energy to break free from the neighbours become a gas; no bubbles are seen.



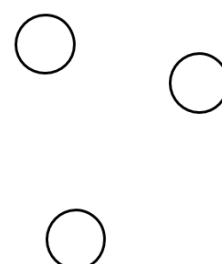
Solid

Tightly packed
Vibrate about fixed positions
Not much kinetic energy



Liquid

Particles fairly close
Can move and swap places
More kinetic energy



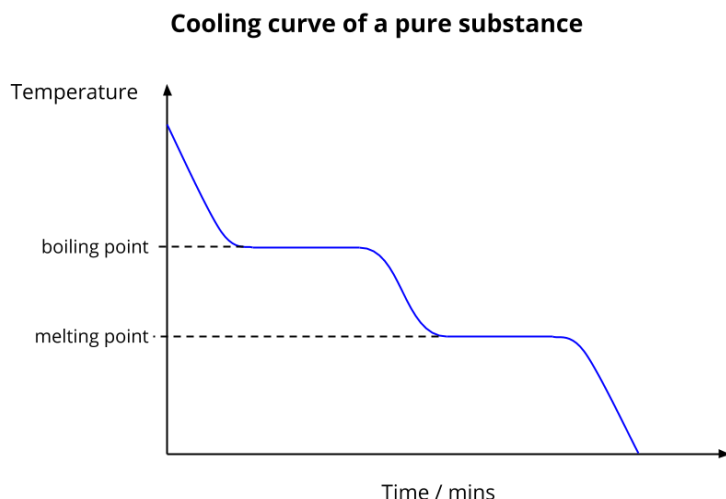
Gas

Particles far apart
Can move large distances
Lots of kinetic energy

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Cooling Curves

A cooling curve shows the temperature of a pure substance as it cools. These curves have a special shape.



The flat bit happens at the melting or boiling point. As the liquid cools, the liquid particles lose energy. This means the temperature falls.

When the liquid freezes (or gas condenses), new bonds are formed between the particles and the particles stick together in a solid.

When these bonds are formed, energy is released. Because this energy is released, it stops the substance cooling.

Once the liquid or solid is formed, no more bonds are made and the cooling process continues.

Mixtures

A mixture is a material made up of two or more different substances which have been mixed and have not reacted. Air is a mixture of gases, as explained on page 14. When we mix two substances, there is no chemical change, so we can (in most cases) obtain the previous materials. When a solid is added to a liquid and the solid dissolves (like sugar in tea), a solution is formed. You need to know the following terms:

Solvent the liquid in which a solute is dissolved to form a solution. For example, some tea.

Solute the minor component in a solution, dissolved in the solvent. For example, the sugar which you stir into the tea.

Solution a liquid mixture in which the solute is uniformly distributed within the solvent. For example, the sweet tea mixture.

Saturated solution a saturated solution is a chemical solution containing the maximum concentration of a solute dissolved in the solvent. Additional solute will not dissolve in a saturated solution. For example, when you put a whole 500g pack of sugar in your cup of tea and it does not dissolve any more.

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Simple distillation

Simple distillation is a method for separating the solvent from a solution. For example, water can be separated from a salt solution by simple distillation. This method works because water has a much lower boiling point than salt. When the solution is heated, the water evaporates. It is then cooled and condensed into a separate container. The salt does not evaporate and so it stays behind. In simple distillation we often use a Liebig condenser. Every pure substance has its own particular melting point and boiling point. One way to check the purity of the separated liquid is to measure its boiling point. For example, pure water boils at 100°C. If it contains any dissolved solids, its boiling point will be higher than this, thus we can prove whether water is pure. We rely on a similar idea of dissolving solids to alter the freezing point when we throw salt on ice and watch the ice melt, since salty ice has a lower freezing point.

Fractional distillation

Fractional distillation is a method for separating a liquid from a mixture of two or more liquids by identifying their different boiling points. For example, we can separate a mixture of liquid water and ethanol by fractional distillation. This method works because ethanol has a boiling point of 78.37 °C, while water boils at 100 °C. When we heat the mixture, the ethanol boils and evaporates before the water.

Filtration

If a substance does not dissolve in a solvent, we say that it is insoluble. For example, sand does not dissolve in water – it is insoluble. Filtration is a method for separating an insoluble solid from a liquid. When a mixture of sand and water is filtered:

- the sand remains in the filter paper (it becomes the **residue**)
- the water passes through the filter paper (it becomes the **filtrate**)

This method works because the tiny particles of water can easily pass through the minute gaps in the filter paper, however the much larger particles of sand cannot pass.

Crystallisation through evaporation

Evaporation is used to separate a soluble solid from a liquid. Just like in simple distillation, when the water evaporates, it leaves the crystallised solute behind. If we do not heat the mixture to make it evaporate, but instead leave it to evaporate slowly at room temperature, then larger crystals will form.

For example, copper sulfate is soluble in water and its crystals dissolve in water to make the water have a blue colour. During evaporation, water leaves the mixture and leaves copper sulfate crystals behind. If we leave it to evaporate slowly, large blue crystals of *hydrous* copper sulfate will be formed. On the other hand, if we boil the mixture, small white crystals of *anhydrous* copper sulfate will be formed. We can use white anhydrous copper sulfate crystals as a test for water, since they will turn blue when water is present.

Paper Chromatography

Chromatography is one of the ways of separating **a mixture** of substances. A mixture is two or more substances mixed together but without being chemically bonded.

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We separate a mixture by identifying a property of the substances which differs for each individual substance. Indicators differ by solubility level and by colour. In order to identify the specific substances we have separated using chromatography we can either compare them with known substances or we can use a numerical factor called the retention factor or R_f value to compare our substances with data from a database.

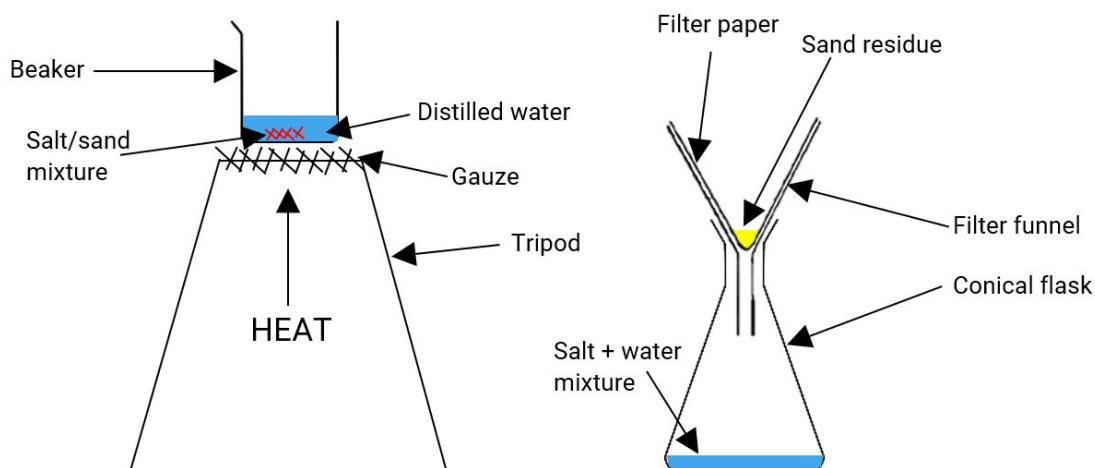
The R_f value is specific to a specific substance in a given solvent.

$$R_f \text{ value} = \frac{\text{Distance the chemical travelled up the paper}}{\text{Distance the solvent travelled up the paper}}$$

R_f values vary with temperature, because solubility varies with temperature.

Definition	Word
The part of the apparatus that the solvent and dyes move up	Stationary phase
A ratio that compares the distance moved by a dye to the distance Moved by The solvent	R_f value
The solvent that moves up the paper in chromatography	Mobile phase
The pattern of spots produced when an ink separates into its different dyes	Chromatogram

Separating Salt and Sand



We added the salt/sand mixture to water and stirred until all the salt had dissolved; we heated it to speed up the dissolving. We then poured the water/salt mixture and sand into a filter funnel with filter paper and waited until all the water had drained along with the salt. Only sand remained. The remaining sand was tested with silver nitrate solution, if it went cloudy, salt still remained in the sand residue.

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Isotopes

Isotopes are atoms of the same element with different numbers of neutrons. They have the same atomic number, however, they have different mass numbers. Most elements have isotopes, the most famous being chlorine, which has two isotopes.

	$^{35}_{17}\text{Cl}$	$^{37}_{17}\text{Cl}$
Protons	17	17
Neutrons	18	20
Electrons	17	17

Calculating relative atomic mass

Most periodic tables do not have a mass number; instead, they have the relative atomic mass (this is also signified as **A_r** in some cases), which is calculated from the masses of the different isotopes and their **abundances**.

- **75%** of Chlorine is ^{35}Cl
- **25%** of Chlorine is ^{37}Cl

RAM (Relative Atomic Mass) of Cl

$$\begin{aligned} &= \left(\frac{75}{100} \times 35\right) + \left(\frac{25}{100} \times 37\right) \\ &= 26.25 + 9.25 \\ &= 35.5 \end{aligned}$$

Another example: **RAM** of Sb

- **57%** of antimony is ^{121}Sb
- **43%** of antimony is ^{123}Sb

$$\begin{aligned} &= \left(\frac{57}{100} \times 121\right) + \left(\frac{43}{100} \times 123\right) \\ &= 121.86 \end{aligned}$$

Diatomic elements

The following elements are diatomic: **[I Have No Clever Or Bright Friends]**

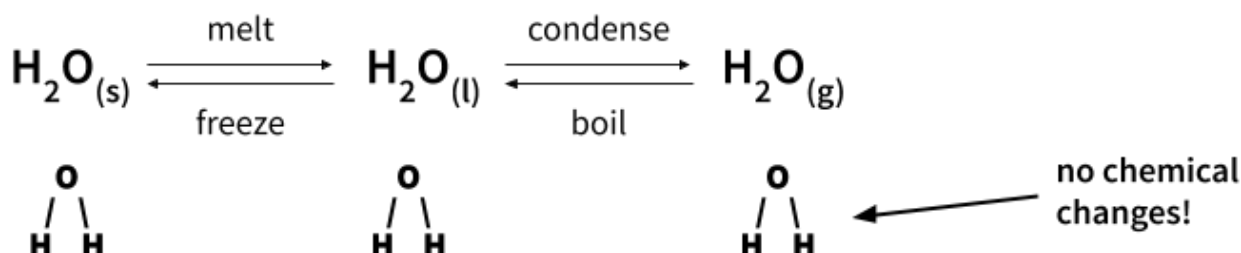
Bromine (Br_2)	Hydrogen (H_2)
Iodine (I_2)	Oxygen (O_2)
Nitrogen (N_2)	Fluorine (F_2)
Chlorine (Cl_2)	

These elements being diatomic means that we **never** find just one of these atoms on its own, but instead these atoms join up to form a molecule of two atoms of that element.

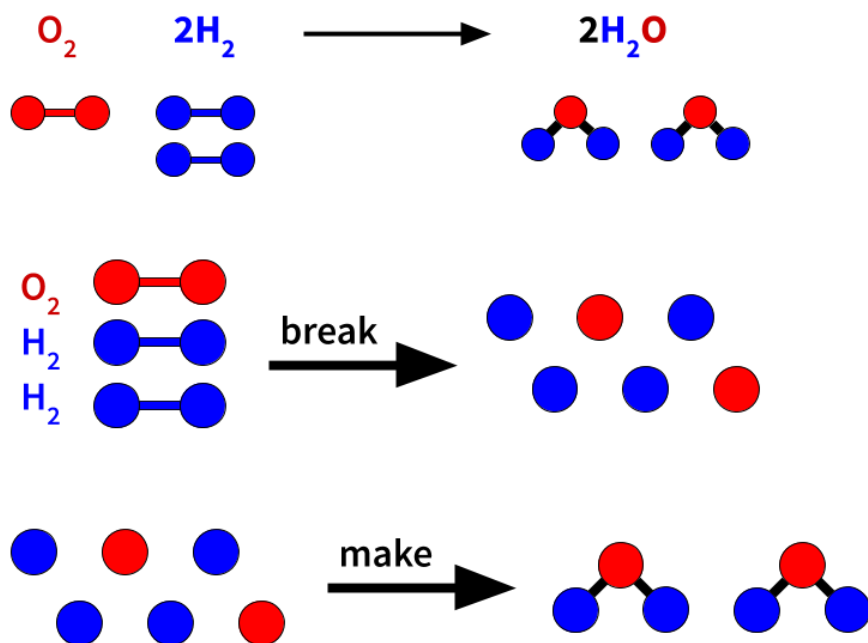
Chemical and Physical Change

Changes of state are **physical** changes. Particles gain or lose energy. They can get closer together or further apart. Bonds are formed or broken between the particles, but no **chemical bonds** are broken or made.

Consider water:



For a chemical change to happen, we need to break and make chemical bonds. Consider making water from hydrogen and oxygen.



When this reaction occurs, we need to break some bonds. Bond-breaking requires energy. It is an **endothermic** process. The product (water) is formed when new chemical bonds are made. **Bond making** releases energy. It is **exothermic**!

To sum up:

- Breaking bonds *requires* energy
- Making bonds *releases* energy

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Electrons and their arrangement

Electrons are found in shells or orbitals which surround the nucleus. As you get further from the nucleus, the shells get bigger. Bigger shells can hold more electrons.

Shell number	Maximum number of electrons
1	2
2	8
3	18 (8 at GCSE)
4	32 (8 at GCSE)

Electron shells always fill from the first shell outwards. The arrangement of electrons in an atom has many different names. It is called electron(ic) arrangement/configuration

It looks like this: ^{12}Mg has an electron arrangement of 2, 8, 2.

[12 is the atomic number, meaning it has 12 protons and 12 electrons]

We need to know the electron arrangement for the first 20 elements.

^1H	1	^6C	2, 4	^{11}Na	2, 8, 1	^{16}S	2, 8, 6
^2He	2	^7N	2, 5	^{12}Mg	2, 8, 2	^{17}Cl	2, 8, 7
^3Li	2, 1	^8O	2, 6	^{13}Al	2, 8, 3	^{18}Ar	2, 8, 8
^4Be	2, 2	^9F	2, 7	^{14}Si	2, 8, 4	^{19}K	2, 8, 8, 1
^5B	2, 3	^{10}Ne	2, 8	^{15}P	2, 8, 5	^{20}Ca	2, 8, 8, 2

All chemical reactions occur because electrons move from one atom to another. Sometimes they are lost or gained. Sometimes they are shared. If electrons are lost from an atom, a positive **ion** is formed.

Acids, Bases & Salts

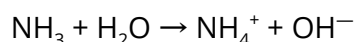
An **acid** is defined as a **substance** that releases H^+ ions in aqueous solution.

A **base** is a substance that reacts with (neutralises) an acid to form a salt.

An **alkali** is a substance that releases OH^- ions in aqueous solution.

A **salt** is a substance formed when the H^+ ions in an acid have been replaced by metal or ammonium ions.

Ammonia is an **alkali** because it reacts with water.



It is an alkali because it releases OH^- ions.

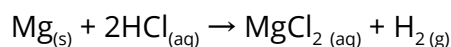
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H⁺ is a hydrogen atom without the electron, therefore just a single proton. Thus an acid can be described as a **proton donor**, while a base is a **proton acceptor**.

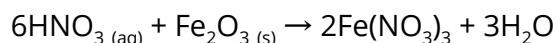
Reactions of acids

When acids react, they form compounds called salts. Different acids form different types of salt, for example hydrochloric acid forms chlorides, sulfuric acid forms sulfates and nitric acid forms nitrates.

1. Acid + metal → salt + hydrogen



2. Acid + base → salt + water



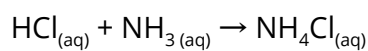
3. a. Acid + carbonate → salt + carbon dioxide + water



- b. Acid + hydrogencarbonate → salt + carbon dioxide + water



4. Acid + ammonia → ammonium salt (only one product!)



As well as learning the equations, you need to know the state symbols. To get these right, you need to know your solubilities.

Solubilities

Chemical compounds vary in their solubility. Some compounds dissolve in water and can form solutions: these compounds are soluble. Other compounds are insoluble, they do not form solutions and so can never have the (aq) sign.

Luck does not work on solubilities! Solubilities need to be learnt.

The first three rules are **always** valid

1. All nitrates are soluble.
2. All Group I compounds are soluble.
3. All ammonium compounds are soluble.

The remaining rules are always valid except when one of the first three rules takes priority.

4. Most carbonates are insoluble
5. Most oxides are insoluble
6. Most hydroxides are insoluble (except Ba(OH)₂ & Ca(OH)₂ which dissolve a bit)
7. Most halides are soluble (except Ag and Pb compounds)
8. Most sulfates are soluble (except Ag, Pb, Ba & Ca compounds)

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Solubility curves

Solubility changes with temperature. By measuring how much solid can be dissolved in the solvent at a given temperature until it no longer dissolves, we can record a solubility value. Plotting these values on a graph of solubility over temperature will produce a curved line of best fit. The solubility of a particular substance increases with temperature, as with most soluble solids that you will meet at GCSE. If a solution of this substance is saturated at a high temperature and then cooled, some of this substance will crystallise / precipitate out of the solution.

Salt Preparation

A salt is formed when the hydrogen in an acid is replaced by a metal or ammonium. Salts can be prepared in multiple different ways. The simplest method is to prepare a soluble salt from an insoluble base or carbonate.

Titration (sodium sulfate)

1. Using a volumetric pipette place 25cm^3 $\text{H}_2\text{SO}_{4(aq)}$ into a conical flask
2. Add a few drops of phenolphthalein
3. Fill a burette with $\text{NaOH}_{(aq)}$ and note the volume
4. Slowly add NaOH to the flask until a permanent pink colour is seen. Record the volume.
5. Repeat without the indicator but use the same volumes of $\text{H}_2\text{SO}_{4(aq)}$ and NaOH
6. Pour mixture into an evaporating basin and heat until about $\frac{2}{3}$ of the water has evaporated
7. Leave to crystallise
8. *Filter to obtain crystals and dry in a warm oven (new on the specification!)*

Insoluble salt (silver chloride)

1. Measure out 25cm^3 $\text{AgNO}_{3(aq)}$ into a beaker
2. Add 25cm^3 $\text{NaCl}_{(aq)}$
3. Stir
4. Filter
5. Wash the precipitate with distilled water
6. Leave to dry

Soluble salt (copper sulfate)

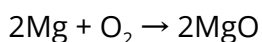
1. Measure out 25 ml of hydrochloric acid into a small beaker and warm it
2. Place a spatulaful of CuO into the acid and stir
3. Keep adding the CuO until it is in excess
4. Filter the mixture to remove excess CuO
5. Heat the filtrate in an evaporating basin until $\frac{2}{3}$ of the water has evaporated
6. Leave to crystallise
7. *Filter to obtain crystals and dry in a warm oven (new on the specification!)*

The acid is heated to speed up the reaction, and excess copper oxide is added to make sure that the reaction is complete and all of the acid has been reacted. Only half of the water is boiled away because crystal creation requires water. If more is boiled, anhydrous copper sulfate will be produced.

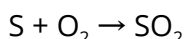
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Combustion

Combustion means burning or exploding. A substance reacts with oxygen, releasing energy (light+heat). For example, magnesium, which burns very well in oxygen with a bright white flame.



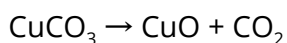
Non-metals can also undergo combustion. Sulfur burns well in oxygen with a bright blue flame.



Melting and decomposition

When heated, thermoplastics and materials with crystalline structures such as metals will melt. On the other hand, some compounds will break down by **thermal decomposition** when heated, forming two or more products from one reactant.

Metal carbonates can often decompose, for example green copper carbonate can thermally decompose to form black copper oxide and carbon dioxide gas.



Thermal decomposition is an endothermic reaction, since breaking bonds requires energy.

Indicators and pH

A substance with a pH lower than 7 can be considered acidic. It will turn universal indicator paper red/orange/yellow and turns litmus paper red.

An acid is defined as a substance that releases H^+ ions in water.

Acids are important laboratory chemicals and we need to know five of them.

Hydrochloric acid	HCl	
Ethanoic acid	CH_3COOH	(vinegar is primarily dilute ethanoic acid)
Sulfuric acid	H_2SO_4	
Nitric acid	HNO_3	
Phosphoric acid	H_3PO_4	

How do we know if a solution is acidic?

We can use a substance called an indicator, which changes colour in the presence of an acid. There are three common indicators.

Name	Color in Acid	Color when Neutral	Color in Alkali
Litmus	Red	Purple	Blue
Methyl Orange	Pink	Orange	Yellow
Phenolphthalein	Colourless	Pale Pink	Bright Pink

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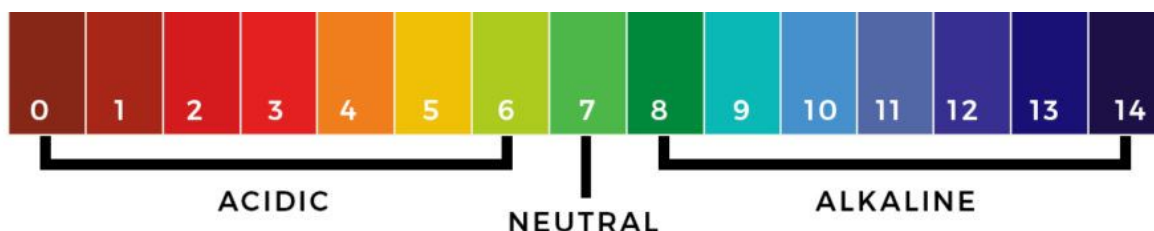
Universal indicator is a mixture of many different indicators. This means that it can have different colours and indicate different strengths of acid and alkali.

The different colours correspond to different pH values and at GCSE, the pH scale goes from 1 through to 14.

pH 1 very strong acid

pH 7 neutral

pH 14 very strong alkali



Whilst universal indicator is useful in telling us whether something is a strong or weak acid or alkali, the actual distinction between different pH values is often difficult to see.

Acids react with bases and alkalis. What is the difference between a base and an alkali?

ALL bases will react with an acid to neutralise it.

ALL metal oxides and metal hydroxides are bases.

Composition of the atmosphere

The atmosphere has evolved over millions of years. The first atmosphere contained ammonia, carbon dioxide and water vapour. Over millions of years chemical reactions occurred and our current atmosphere came into being:

Nitrogen	78%
Oxygen	21%
Argon	0.9%
Carbon dioxide	0.04%

Other gases make up only a small percentage and vary by location.

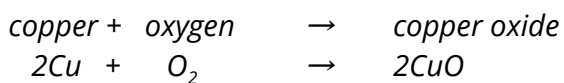
An experiment to show the percentage of oxygen in the air

The percentage of oxygen in the air can be measured by passing a known volume of air over hot copper, and measuring the decrease in volume as the oxygen reacts with it. There is an excess of copper turnings so that all the possible oxygen can react. The volume of air will stop decreasing once all the oxygen has reacted.

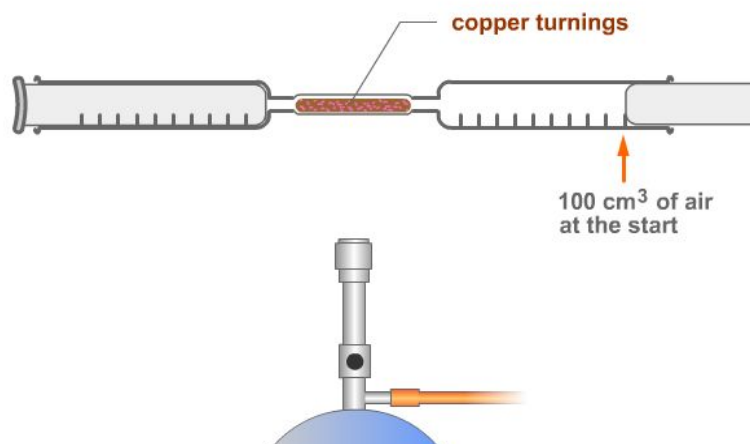
Note that there is some air in the tube with the copper turnings. The oxygen in this air will also react with the hot copper, causing a small error in the final volume recorded. It is also important to let the apparatus cool down at the end of the experiment, otherwise the final reading will be too high. There could also be a leakage and some air has escaped from the apparatus.

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Here are the word and symbol equations for this experiment:



Gas syringes are used to measure the volume of gas in the experiment.



Results

Volume of air at start 100cm³

Volume of air at end 75cm³

$$\frac{100-75}{100} \times 100 = 25\% \text{ oxygen}$$

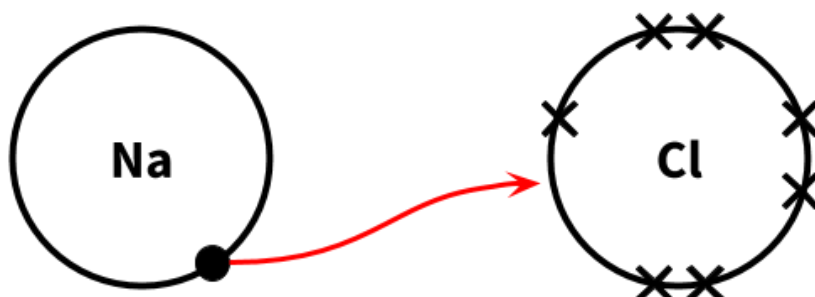
Some air escaped from the apparatus in our experiment. This means the final volume ended up being *less* than it should be, so the resultant calculated concentration of O₂ is higher than it really is.

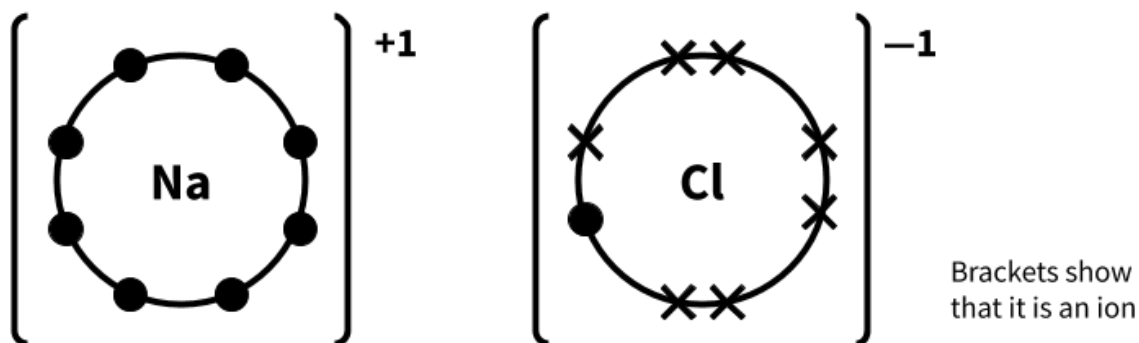
Chemical Bonding

Whenever a chemical reaction takes place, chemical bonds are broken and made. Different elements form different types of bonds. Whenever a metal reacts with a non-metal an **ionic bond** is formed. All atoms react in order to get a full outer shell of electrons. Consider sodium reacting with chlorine, forming sodium chloride, for example:

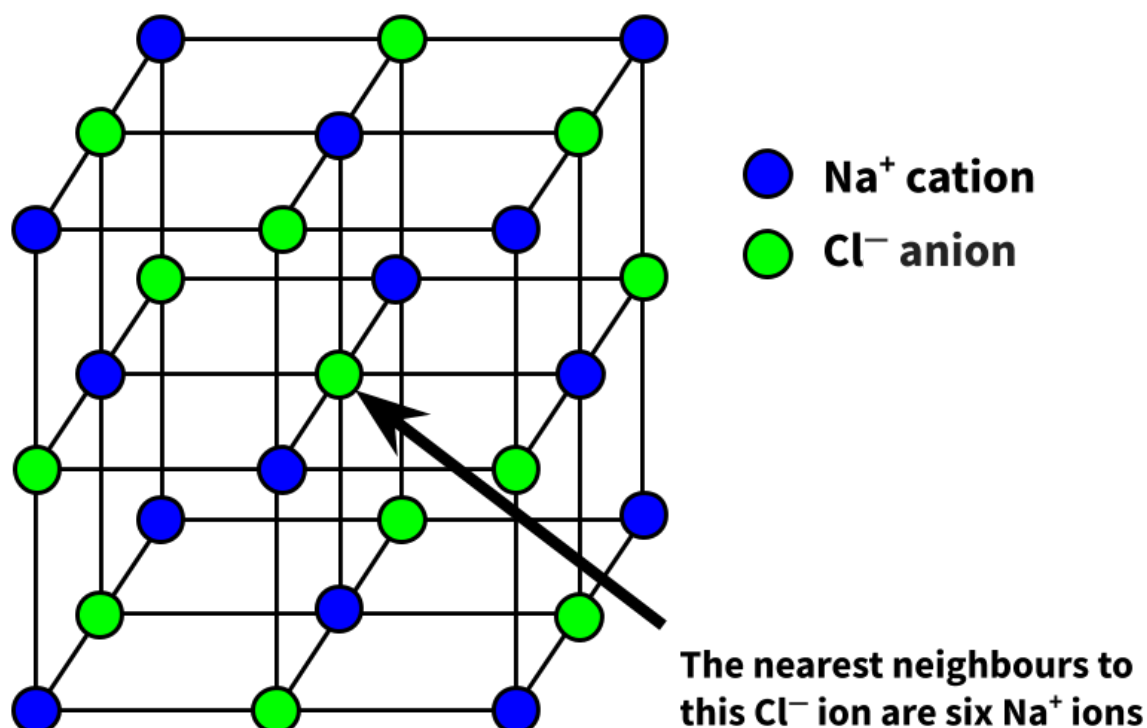
¹¹Na has 11e⁻ & 11p⁺
2, 8, 1 arrangement
1 in outer shell

¹⁷Cl has 17e⁻ & 17p⁺
2, 8, 7 arrangement
7 in outer shell





Ionic Structures



Metal atoms lose electrons to form positively charged ions or **cations**. (Remember that **cations** are **pussy**itive) Non-metal atoms gain electrons to form negatively charged ions or **anions**.

The oppositely charged ions that are formed are held in a regular three-dimensional lattice by electrostatic attractions between the ions. The ions pack together in the most efficient way so there is little wasted space.

In NaCl, each Na^+ ion is surrounded by six Cl^- ions, and each Cl^- ion is surrounded by six Na^+ ions. This means that there are many strong electrostatic attractions between the ions within the giant lattice structure.

Examples of ionic compounds include:

- Sodium chloride
- Copper (II) sulfate

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The giant ionic lattices that are formed by ionic compounds account for their properties or characteristics:

1. Melting and boiling points

Each ion is strongly bonded by electrostatic attraction to several others. These strong bonds are hard to break. A lot of **heat energy** is required and so the melting and boiling points are high.

2. Crystalline structure

Salt has a hard crystalline structure. Crystals require a very regular arrangement with straight lines, meaning that the ions line up in straight lines, forming crystals with precise and straight edges and corners.

3. Electrical conductivity

The flow of charged particles creates an electrical current. If ions are able to move, they can conduct electricity. Ionic compounds such as NaCl contain ions. When they are solid, they cannot conduct electricity because the *ions* are unable to move. If the crystal is melted to become a liquid or dissolved in water to become a solution, the *ions* are free to move and they can carry an electrical charge.

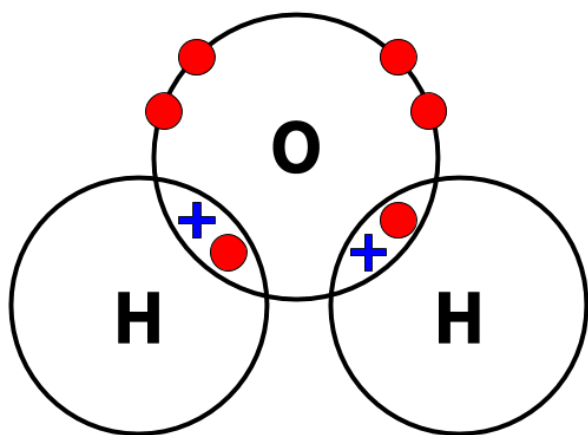
4. Brittleness

Ionic crystals are **brittle**; they shatter when hit. This is because the hitting can displace a row of ions which are then forced next to ions of a similar charge. Ions with the same charge repel, so the crystal breaks apart.

Covalent Bonding

Covalent bonding occurs between two or more non-metal atoms. The atoms all need to gain electrons, so they share. The sharing is *always* fair!

How to draw a dot and cross diagram

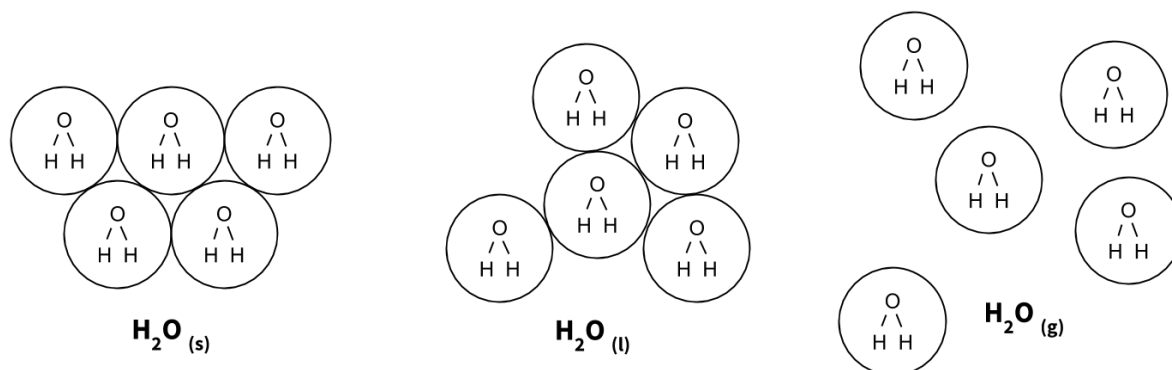


e.g. H_2O

1. Draw atom circles first and label them.
2. Concentrate on the outside atoms and what they need to gain.
3. Fill in the shared areas fairly
4. Put in the rest of the outer shell electrons

Simple Molecular Structures

The vast majority of covalently bonded compounds exist as simple molecules. Water, H_2O , is an example. Consider the three states of water.



The water molecule H_2O is the same in each state. The covalent bonds **do not change**.

Simple molecular structures (e.g. H_2O , CO_2 , O_2 , NH_3 , etc.) tend to have low melting and boiling points. The covalent bonds are very strong, but the **intermolecular forces** that exist between different molecules are weak. These forces of attraction need only a small amount of heat energy to be broken, and so simple molecular substances are easily melted and boiled.

Electrical conductivity

Simple molecular structures do not conduct electricity under **any** circumstances, as they have no ions, and all the electrons are bonded to the atoms and are not free to move. Water conducts electricity only when it contains impurities such as salt. Pure (distilled) water does **not** conduct.

Giant Covalent Structures

The vast majority of non-metals bind together to form simple molecules. In a few cases (3 at IGCSE) they form giant covalent lattices.

Diamond

Diamond is an **allotrope** of carbon. An allotrope is a different physical form of the same element. Each carbon atom is covalently bonded to 4 others in a tetrahedral arrangement. These covalent bonds are very strong and its structure is hard and rigid.

Properties of diamond

a. Melting point / boiling point

Diamond has a **very high** melting and boiling point. The whole structure is held together by many strong covalent bonds. These need a large amount of heat energy to break them.

Fact: Diamond sublimes at roughly 3600°C

b. Electrical conductivity

Diamond has no ions and no mobile electrons, so there is no way it can conduct.

c. Crystalline structure

The atoms are all lined up in a regular lattice. This makes diamond crystalline.

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Graphite

Each carbon atom is covalently bonded to 3 others in the same layer. Loose electrons are found between the layers and hold the layers together as an intermolecular force.

Properties of Graphite

a. Melting point / boiling point

Very high because you need to break many strong covalent bonds and this requires a lot of heat energy.

b. Electrical conductivity

Graphite *does* conduct electricity. The electrons between the layers are free to move, and therefore because they are charged *and* mobile, they can conduct. Graphite is the **only** covalent conductor.

Graphite layers are held very loosely in place. They slide off very easily, making graphite slippery to the touch. This means it can be used in “lead” pencils. Since graphite is a **lubricant** it is used inside locks. It is also great for unsticking zippers.

Silicon dioxide

Silicon dioxide (SiO_2), also called silica, is a giant lattice in which each silicon atom is covalently bonded to four oxygen atoms and each oxygen atom is bonded to two silicon atoms.

Note: Do not confuse it with silicone. Silicone is used for implants.

Properties of silicon dioxide

a. Melting point / boiling point

Silicon dioxide is a giant structure. There are many strong covalent bonds between the atoms, so lots of heat energy is needed to break them. Therefore the melting and boiling points of silicon dioxide are **very high**.

b. Electrical conductivity

Silicon dioxide does not conduct electricity, as there are no ions and no free electrons, so nothing can carry a charge.

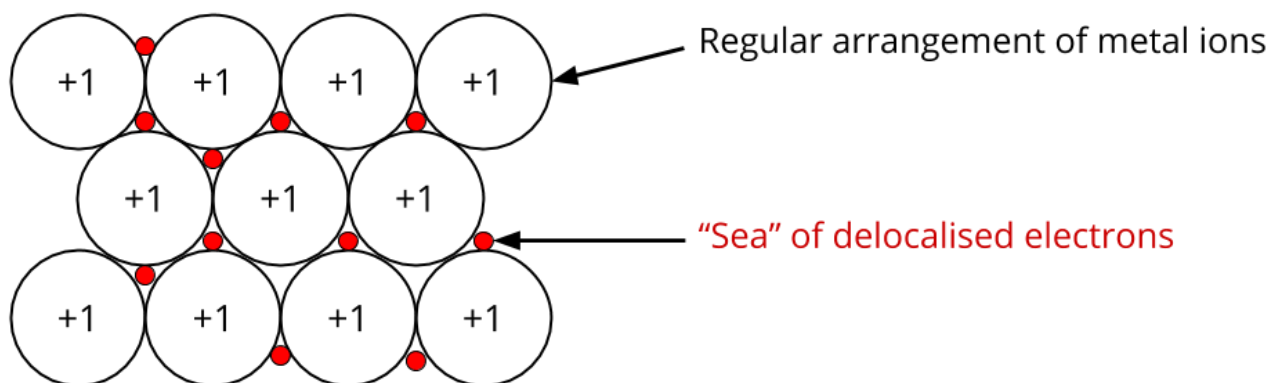
Buckminsterfullerene

Colloquially known as footballane due to its shape.

Buckminsterfullerene is a simple molecular structure. It has a low melting and boiling point due to weak forces of attractions between the molecules. Each carbon atom is covalently bonded to 3 others. This is the same arrangement as in graphite. The 4th outer shell electron of each atom is delocalised and is free to move around the ball. However, these electrons cannot move from one ball to another, meaning that there is no electrical conductivity because there are no mobile charged particles. Buckminsterfullerene is used as miniature ball bearings in machinery (just like ball bearings in a fidget spinner) and joints as it is a suitable lubricant.

Metallic Structures

All metals have a giant metallic structure. Metal atoms line up in a regular crystalline fashion. The outer shell electrons leave the atoms and move freely around the structure. These electrons are said to be **delocalised**. Consider sodium, for example:



The structure is held together by numerous electrostatic attractions between the metal ions and the sea of delocalised electrons. These attractions need a lot of heat energy to be overcome, therefore metals have high melting points. (at GCSE)

Metals are **good electrical conductors** as the electrons are free to move throughout the structure and thus they can carry a charge.

The Periodic Table

The Periodic Table is a list of all known elements arranged in a very useful way. The elements occur in order of their **atomic number**. Rows going across the table are called **periods**. All elements in the same period have the same number of electron shells. Columns going down the table are called **groups**. Elements in the same group have the same number of electrons in the outer shell.

At GCSE we study 3 chemical groups:

- Group 0 Noble gases
- Group I Alkali Metals
- Group VII Halogens

Group 0 — Noble gases

This group is not very interesting. All members of the group have a full outer shell of electrons so they do not form any chemical bonds. This makes them exceedingly unreactive. However, this lack of reactivity means they are useful as such:

He	Helium	party balloons, airships, blimps
Ne	Neon	advertising signs, "neon lights"
Ar	Argon	incandescent light bulbs (stops filament from burning)
Kr	Krypton	used as a "flash" in photography (high-speed flash)
Xe	Xenon	used in "sunbed" tubes

Group I — Alkali Metals

This group of elements is *much* more interesting.

Li	Lithium	
Na	Sodium	Available in school
K	Potassium	
Rb	Rubidium	
Cs	Caesium	Available, but the school does not buy them
Fr	Francium	Radioactive and only exists for fractions of a second

Group I elements react by losing an electron. As you go down the group, the atoms get bigger. The outer shell electrons gets further from the nucleus. The electron shells between the outer electrons and the nucleus act as “shields” and deflect some of the **nuclear attraction**. The outermost electron feels less attraction from the nucleus, and so it is lost more easily and reactions happen more readily.

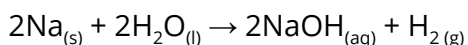
- **Reactivity increases** as you move down the group
- **Melting point decreases** as you move down the group
- **Hardness decreases** as you move down the group
- **Density increases** as you move down the group

The alkali metals are all very reactive. In the presence of air, the shiny metals will quickly tarnish as they react with the oxygen present. To prevent this from happening, lithium, sodium and potassium are stored in oil to prevent them from reacting. Rubidium and caesium are so reactive that they have to be stored in sealed glass tubes.

They will also react rapidly with water to form strongly alkaline solutions of the metal hydroxides. This is why they are known as the **alkali metals**.

Reaction with water

A piece of sodium is dropped into a basin of water.



Observations include:

- Sodium metal floats because it is less dense than water.
- Sodium melts into a ball because a lot of heat is produced in the reaction.
- There is fizzing/effervescence because hydrogen gas is produced.
- The sodium moves around on the surface of the water.
- The piece of sodium becomes smaller and smaller until it disappears.
- Addition of universal indicator after the reaction has finished turns the solution blue, indicating an alkaline solution with OH^- ions present.

If the experiment was repeated, but using potassium, it would fizz *more* vigorously, move around *more* vigorously and disappear *more* quickly. The potassium metal would also produce enough heat energy to ignite the hydrogen gas produced and thus burn with a lilac flame.

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Group VII — The Halogens

All the Halogens are toxic non-metals and have 7 electrons in their outer shell.

Fluorine	F ₂	Pale yellow	gas
Chlorine	Cl ₂	Green	gas
Bromine	Br ₂	Red / Brown	liquid
Iodine	I ₂	Grey / Black	solid
Astatine	At ₂	Black*	solid*

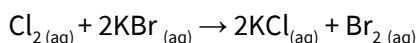
* these are predictions since Astatine is radioactive and decays too quickly for us to see

Halogen Reactivity

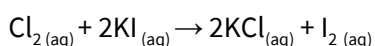
Halogens react in a series of **displacement reactions**.

	KCl _(aq)	KBr _(aq)	KI _(aq)
Cl _{2 (aq)}		Pale green → orange. Addition of cyclohexane turned it orange.	Pale green → brown. Addition of cyclohexane turned it pink.
Br _{2 (aq)}	No colour change		Orange → brown. Addition of cyclohexane turned it pink.
I _{2 (aq)}	No colour change	No colour change	

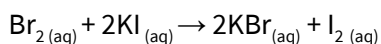
In the first reaction below, the more reactive chlorine displaces the less reactive bromine from the compound.



colour in cyclohexane: **orange**



colour in cyclohexane: **pink**



colour in cyclohexane: **pink**

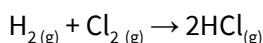
Conclusion: Chlorine is more reactive than bromine, which is more reactive than iodine.

In each case, the more reactive halogen is able to displace the less reactive one from its compound.

Halogens have **seven** electrons in their outer shell. When they react, they need to gain **one** electron. Fluorine is **2,7** so the incoming electron goes into the second shell. This is close to the nucleus and there is only **one** shield and so there is a stronger nuclear attraction. As you go down the group, the atoms get larger, so the outer shell gets further from the nucleus and the number of shields increases. The attraction between the incoming electron and the nucleus decreases. It becomes more difficult for the new electron to join the atom. **Reactivity decreases** as you move down the group.

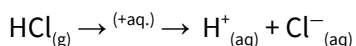
Hydrogen Halides

All halogens react with hydrogen gas to form hydrogen halides. For example:



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All hydrogen halides are covalently bonded molecules and are gases at RTP. When any hydrogen halide dissolves in water (*only* water; the solvents topic will be covered in 6th form) it **dissociates** (splits into two) and turns into ions. For example:

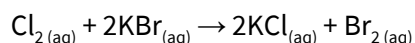


It is the H^+ ions which makes the solutions acidic, therefore $\text{HCl}_{(\text{aq})}$ is hydrochloric *acid*.

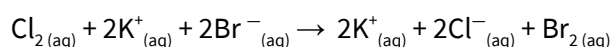
Redox and ionic equations

[All high-level stuff] Halogens react with halide ions in a series of displacement reactions. By looking at these in more detail, we can see that they are actually **redox** reactions, in which reduction and oxidation are both taking place. To look at the reactions more closely, we need to convert the balanced symbol equations into **ionic equations**. To do this, we follow a series of steps:

1. Write out the full balanced chemical equation with state symbols.



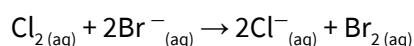
2. For all **compounds** that are *both aqueous and ionic*, split them into their ions.



3. Cancel out "spectator ions"

In this case, remove $2\text{K}^+_{(\text{aq})}$ from both sides as it does not change nor react.

4. Write out the ionic equation *with state symbols*.



How is this redox?

"Redox" is a reaction that involves simultaneous oxidation **and** reduction.

What is oxidation?

Oxidation can be defined as

- When something reacts with and combines with oxygen
- When something loses hydrogen
- When something loses electrons

For example: $2\text{Br}^- \rightarrow \text{Br}_2 + 2\text{e}^-$

Bromide has lost electrons. It has been oxidised.

What is reduction?

Reduction can be defined as

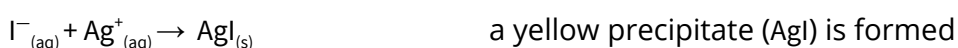
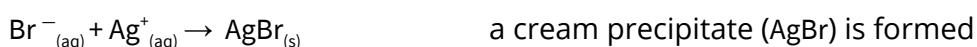
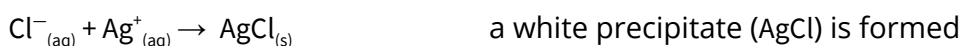
- When something reacts with and gains hydrogen
- When something loses oxygen
- When something gains electrons

For example: $\text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Cl}^-$

Chlorine has gained electrons. It has been reduced.

Silver Halides

All aqueous halide ions react with silver nitrate solution to form **precipitates**. These precipitates are of characteristic colour, and so they can be used to identify the halide ions.



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Metals and non-metals

In general, metals conduct electricity and have oxides which are basic, reacting with acids to give a salt and water. Non-metals do not conduct electricity (except for graphite) and have oxides which are acidic or neutral. The metals are on the left hand side of the Periodic Table and non-metals are on the right hand side. There is some uncertainty on where exactly the dividing line lies. For example, arsenic (As) has properties of both metals and non-metals.

Organic Chemistry

Organic chemistry is the study of covalently bonded carbon compounds. The subject is divided into studying different types of organic compounds, which range from very simple (e.g. CH_4 - methane) to very complicated (e.g. haemoglobin)

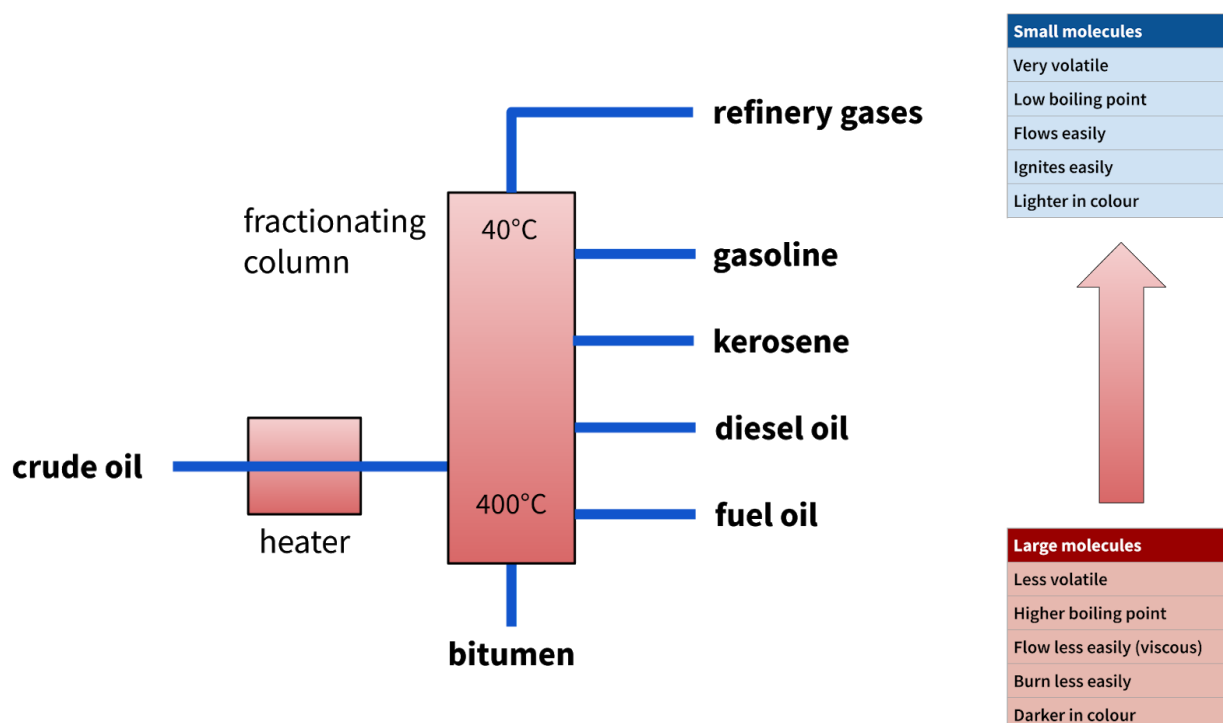
The simplest organic compounds are **hydrocarbons**. These are compounds that contain *only* hydrogen and carbon atoms. Many hydrocarbons are sourced from **crude oil**.

Crude oil is a fossil fuel. It has taken tens of millions of years to form. It is formed from dead plankton. These dead organisms lived in the prehistoric seas. When they died, they sank to the bottom and were covered in layers of sediment. Neither oxygen nor water was present: only the dead material and the sediment. (If there had been oxygen, it would have decomposed)

At high pressure and temperature between 90°C and 120°C a series of very slow reactions turn the dead material into crude oil. Crude oil is extracted from rocks and processed at oil refineries.

Refining crude oil

Crude oil is pumped into the refinery where it can undergo a variety of treatments. **Fractional distillation** occurs first. This occurs in a giant fractionating column which is between 80 and 100 metres tall.



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Crude oil is heated until it boils and the vapours pass into a fractionating column, which is cooler at the top and hotter at the bottom. The vapours rise up the column. How far up the column a particular hydrocarbon moves depends on its boiling point. Smaller hydrocarbons have lower boiling points, while longer chain hydrocarbons have higher boiling points.

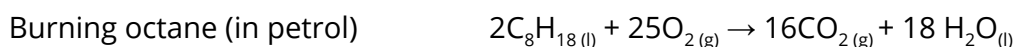
Crude oil and its uses

Crude oil itself has no uses and it has to be separated into fractions before it can be used. These fractions are all mixtures, but each one contains a narrow range of sizes of hydrocarbons with similar boiling points.

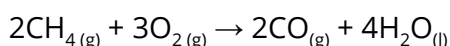
Refinery gases	Refinery gases are a mixture of methane, ethane, propane and butane, which can be separated into individual gases if required. These gases are commonly used as liquefied petroleum gas (LPG) for domestic heating and cooking.
Gasoline (petrol)	As with all other fractions, petrol is a mixture of hydrocarbons with similar boiling points. It is used as fuel in cars.
Kerosene	Kerosene is used as a fuel for jet aircraft, as domestic heating oil and as “paraffin” for small heaters and lamps.
Diesel	Used as fuel for buses, lorries, some cars and certain railway engines.
Fuel oil	Fuel oil is used as fuel for ships and for industrial heating.
Bitumen	Bitumen is a thick sticky black material (also known as <i>residue</i>) which is used to make the top surface of roads and roofs.

Hydrocarbons

All hydrocarbons burn in the air (oxygen) to form carbon dioxide and water and release heat in the process. The various fractions obtained from the column can, therefore, be used as **fuels**. A fuel is a substance which, when burned, releases heat energy. Examples:



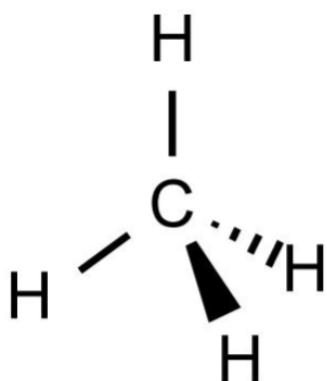
If there is not enough air (or oxygen), incomplete combustion occurs. This leads to the formation of carbon (soot) or carbon monoxide instead of carbon dioxide. For example, if methane burns in a badly maintained gas appliance, there may not be enough oxygen available to produce carbon dioxide, and so you get toxic carbon monoxide instead:






The formation of carbon monoxide from the incomplete combustion of hydrocarbons is very dangerous. Carbon monoxide is colourless and odourless and is very poisonous. Carbon monoxide is poisonous because it reduces the ability of the blood to carry oxygen around the body.

Fact: Carbon monoxide has a higher affinity for haemoglobin than oxygen does. Once carbon monoxide has bonded to haemoglobin it stays there until the red blood cell dies. (Up to 3 months) This means that the supply of oxygen to the body's cells is reduced and cells cannot respire. Cells die and, depending on how much carbon monoxide has been inhaled, the person can die as well.

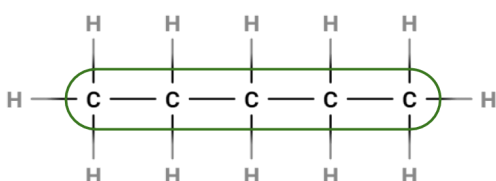
Chemical diagrams



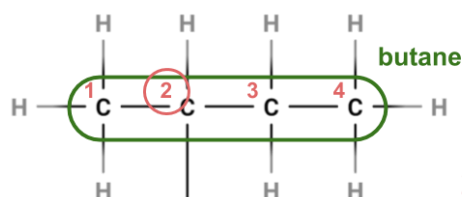
-  In the plane of the page
-  Comes *forwards* out of the plane of the page **∴ in front**
-  Goes *backwards* out of the plane of the page **∴ behind**

Left: wedge-and-dash projection of methane. Methane has a tetrahedral arrangement.

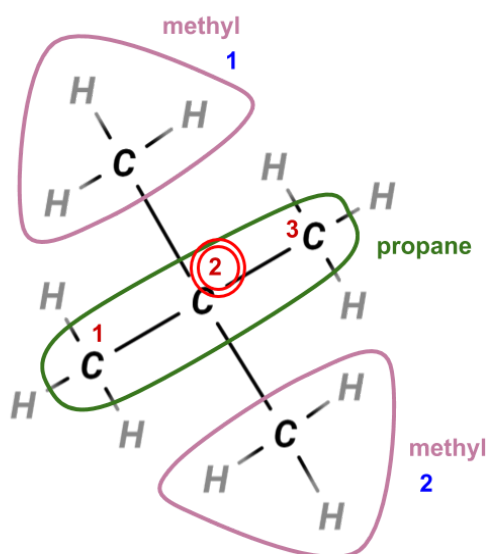
Naming compounds



pentane
Longest chain
of carbons is 5
∴ **pent-**



2-methylbutane



2,2-dimethylpropane

To name a compound:

1. Decide what homologous series it is and do the ending. If it is an alkane, it will end in "ane". If it is an alkene, it will end in "ene".
2. Find the longest chain (it may go around corners) and put it at the end.
3. Identify the "dangly bits" and list them in alphabetical order.
4. Number the carbon atoms in the chain so that the dangly bits get the lowest numbers.
5. Complete the name using commas to separate numbers and using hyphens to separate numbers and letters.

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Alkanes

Organic chemicals are arranged in families called **homologous series**. The alkanes are an example of a homologous series. The alkanes are described as being **saturated** (meaning they *only* contain single bonds) **hydrocarbons** (meaning they *only* contain hydrogen and carbon atoms).

When dealing with any organic compound, certain bonding rules are *always* obeyed.

C	4 bonds
H	1 bond
Halogen	1 bond
Oxygen	2 bonds

The alkanes, like all homologous series, share a **general formula**, have **similar chemical properties** and **trends in physical properties**.

General formula C_nH_{2n+2} Where n is any whole number ≥ 1

n	Formula	Structure	Name
1	CH_4	<pre> H H - C - H H</pre>	methane
2	C_2H_6	<pre> H H H - C - C - H H H</pre>	ethane
3	C_3H_8	<pre> H H H H - C - C - C - H H H H</pre>	propane
4	C_4H_{10}	<pre> H H H H H - C - C - C - C - H H H H H</pre>	butane
5	C_5H_{12}	<pre> H H H H H H - C - C - C - C - C - H H H H H H</pre>	pentane
6	C_6H_{14}	<pre> H H H H H H H - C - C - C - C - C - C - H H H H H H H</pre>	hexane

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Remember the first four alkanes!

- Monkeys Eat Peanut Butter
- My Elephant Prefers Bananas

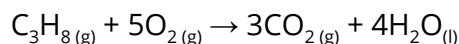
Reactions of Alkanes

Alkanes are not very reactive when compared to other organic chemicals. They only do two types of reaction.

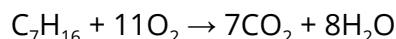
Combustion

All alkanes burn very well and produce a lot of heat energy. The vast majority of them are used as fuels, and, provided the oxygen supply is sufficient, they undergo **complete combustion**.

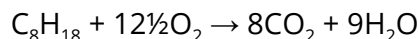
1. **Propane C₃H₈**



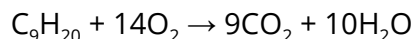
2. **Heptane C₇H₁₆**



3. **Octane C₈H₁₈**

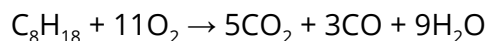


4. **Nonane C₉H₂₀**



Sometimes the oxygen supply is limited. If this happens, the alkane will still burn, but **incomplete combustion** takes place. Many different products can be formed depending on both the amount of oxygen and pure luck.

Octane C₈H₁₈

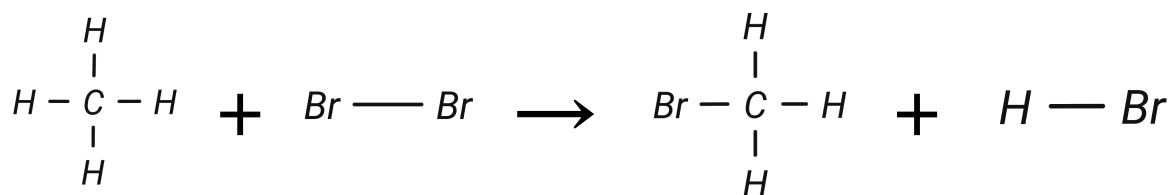
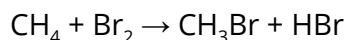


Both are correct. Both can occur.

Carbon monoxide has a higher affinity for haemoglobin than oxygen does. *For more information about carbon monoxide poisoning, see the section about hydrocarbons.*

Substitution

When an alkane is mixed with a halogen (an element in Group VII) *in the presence of UV light* hydrogen atoms can be swapped for halogen atoms to form a product called a **halogenoalkane**. There is only one such reaction on the syllabus, however:

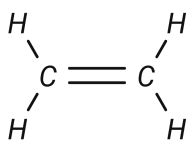
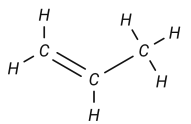
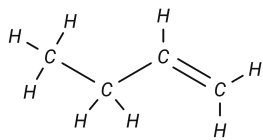
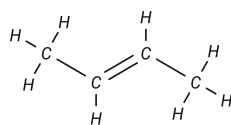
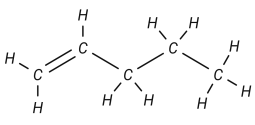
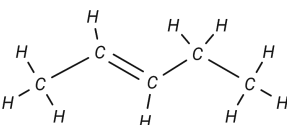


methane + bromine → bromomethane + hydrogen bromide

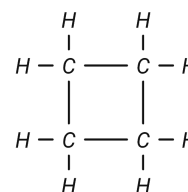
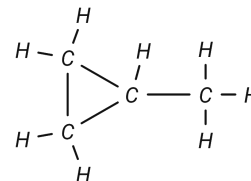
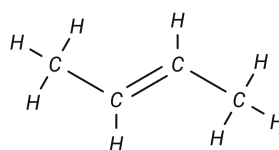
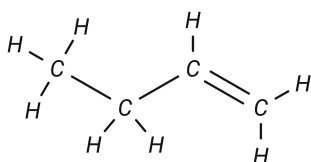
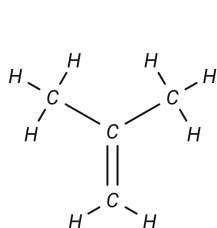
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Alkenes

The alkenes are another homologous series, and they too are hydrocarbons. Unlike alkanes, which are **saturated**, alkenes are **unsaturated**. A saturated compound contains *only* single bonds. An unsaturated compound contains (at least) one C=C, i.e. a double bond. Alkenes have a general formula C_nH_{2n}

n	name	formula	structure
1	does not exist, else there would not be a double bond.		
2	ethene	C_2H_4	
3	propene	C_3H_6	
4	but-1-ene	C_4H_8	
	but-2-ene		
5	pent-1-ene	C_5H_{10}	
	pent-2-ene		

One of the fun things about alkenes is that they have loads of isomers. Another homologous series, the cycloalkanes, also have a general formula C_nH_{2n} . This means that C_4H_8 has 5 isomers (7 at A-level)



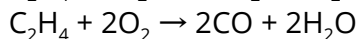
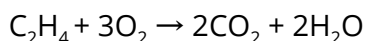
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Reactions of Alkenes

Combustion

Like alkanes, alkenes can undergo complete or incomplete combustion.

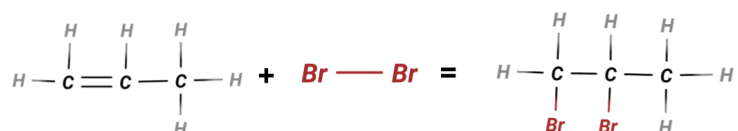
ethene C_2H_4



However, unlike alkanes, alkenes are not used as fuels. They are too useful in other reactions and therefore they are too valuable to burn.

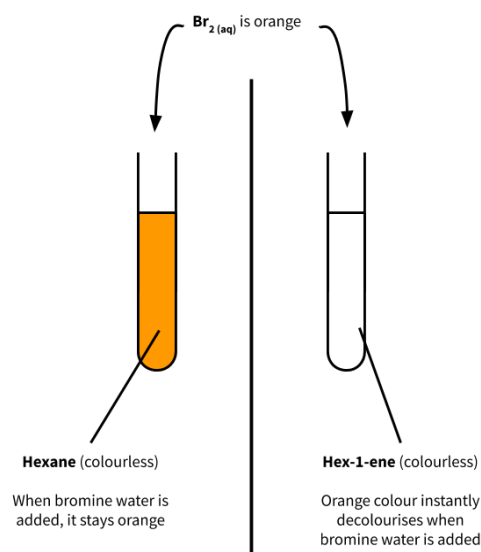
Addition reactions

Alkenes contain a $C=C$. This is made up of a regular single bond ($C-C$), plus a second type of bond called a pi/ π bond. The π bond is weaker than single bonds, and it breaks fairly readily. This means that other atoms can add on to the carbons.



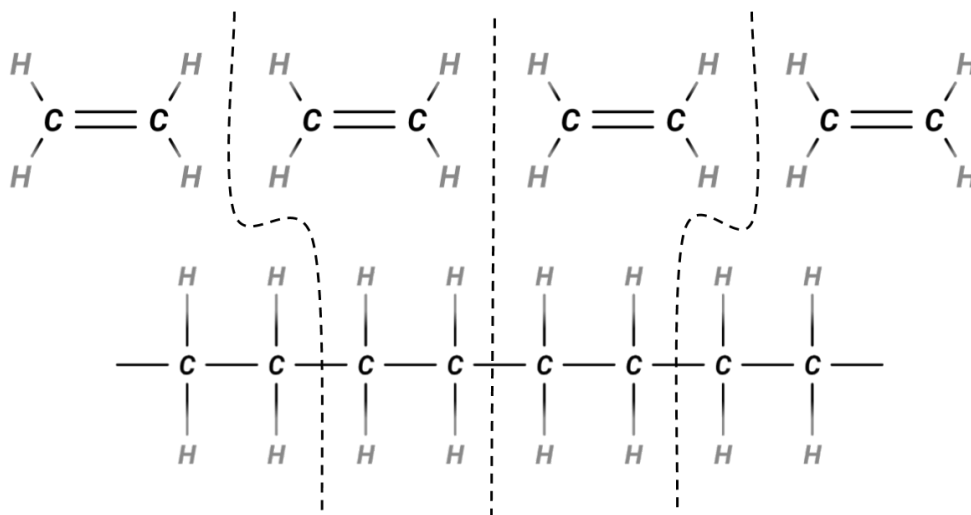
Propene + bromine \rightarrow 1,2-dibromopropane (colourless)

This reaction happens instantly. No UV light is needed. Bromine water is used as a test to distinguish between alkanes and alkenes.



Polymerisation

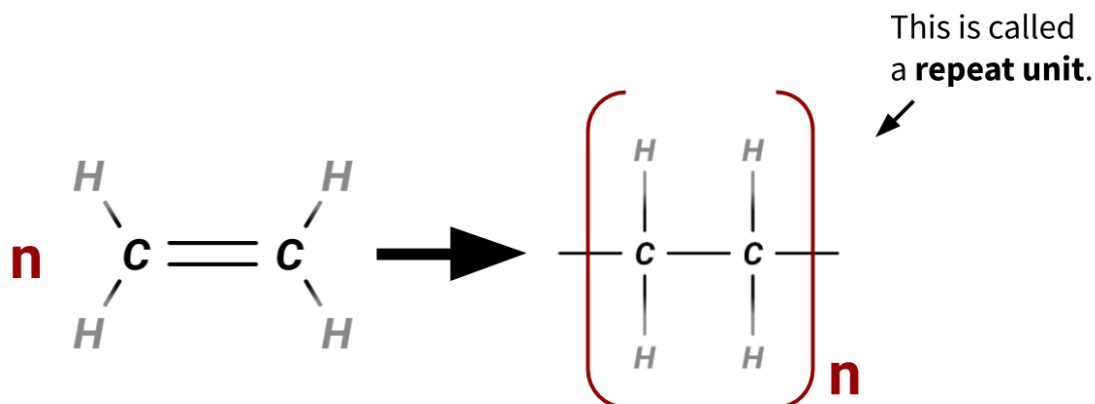
Alkenes are most widely used to make **polymers**. An alkene molecule is a **monomer**. When it reacts with many other monomers it makes a polymer. Most polymers are used as plastics. When any alkene polymerises, *only* the $C=C$ reacts. $C=C \rightarrow -C-C-$ Consider ethene:



Atom economy of 100%

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To write this as a balanced symbol equation, we use 'n'



The problem with polymers

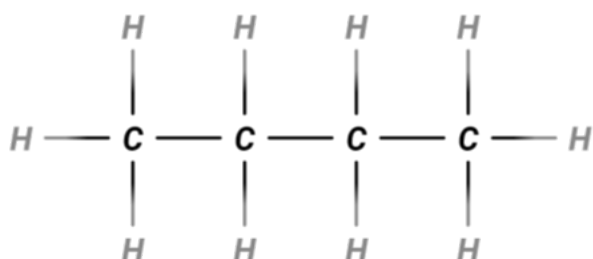
The big problem with polymers is that they are non-biodegradable. (Think plastic bags) They are all chemically unreactive. This makes plastic rubbish harmful to wildlife and ecosystems. (Turtles, swans, fish)

So, to dispose of them, we could burn the plastics, using them as fuel in power stations. However, burning plastics produces toxic fumes, including CO_2 , CO and HCl.

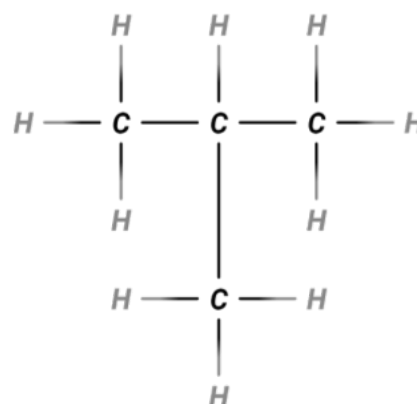
So, we could alternatively recycle them. But this is a very expensive process, especially due to the need for sorting and cleaning.

Or we could use chemistry to make life better, by developing biodegradable plastics, developing bacteria to "eat" plastics or by putting filters (such as NaOH) on chimneys to absorb toxic fumes when burning.

Isomers



Butane is C_4H_{10}



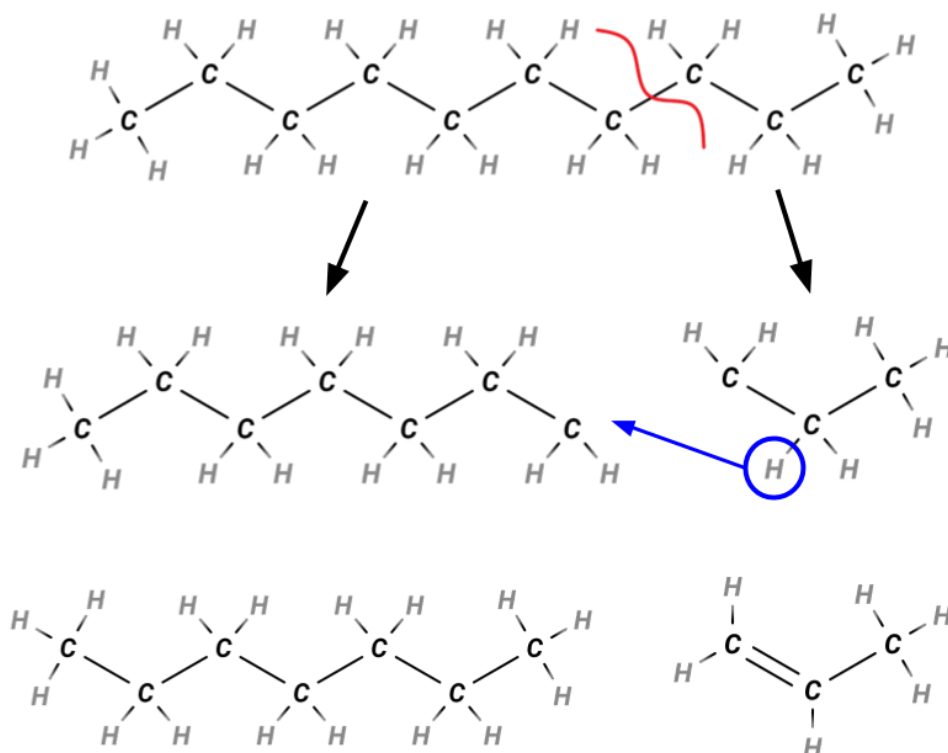
So what is this...?

This is an **isomer** of butane. It has the same molecular formula but its atoms are arranged in a different order. At GCSE, remember that the isomer of butane is called *2-methylpropane*.

Cracking

Cracking is a chemical process in which a long chain alkane is broken down; it is a type of thermal decomposition. It is a way of producing valuable alkenes. Cracking is a totally random process. The alkane molecule can be cracked in any position and any numbers of times.

The long-chain alkane molecule (which is usually liquid at RTP) is heated to give a gas and then passed over a catalyst of silicon dioxide (also known as *silica*) and aluminium dioxide (also known as *alumina*) at about 600-700°C. Alternatively, cracking can be carried out at higher temperatures without a catalyst, known as thermal cracking.



We can split the decane (C₁₀H₂₂) molecule in any place.

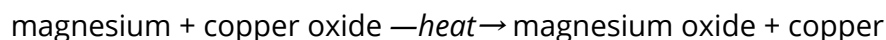
However, once we have cracked it, we have two unhappy fragments. So, the H moves and C=C forms.

Now we have cracked decane to form heptane and propene.

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Reactivity and Redox

A reminder of common entrance:



This is described as a **displacement reaction**. The more reactive magnesium removes the less reactive copper, displacing the copper and leaving it as an element.

At GCSE, we describe reactions like this as **redox reactions**. The magnesium is becoming oxidised because it is gaining oxygen to become magnesium oxide. The copper oxide is being reduced because it is losing oxygen to become copper. We also say that magnesium is a **reducing agent** because it removes the oxygen from copper oxide. The copper oxide can be described as an **oxidising agent** because it provides oxygen for magnesium.

Reactivity Series

Metal	Reaction with water	Reaction with acids	Extraction method
K	React with water to form a metal hydroxide (alkaline) and hydrogen gas. e.g. $\text{Ca} + 2\text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2 + \text{H}_2$	Would react with acids but <i>not safely</i> (in a lab) so we do not do it	Extracted from their ores by electrolysis of molten compounds.
Na			
Li			
Ca			
Mg	Either reaction (above or below)	React with acids to form a metal salt and hydrogen gas. e.g. $\text{Zn} + 2\text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2$	
Al	React with steam to form a metal oxide and hydrogen gas. e.g. $\text{Mg} + \text{H}_2\text{O}_{(\text{g})} \rightarrow \text{MgO} + \text{H}_2$		
(C)			(for comparison)
Zn			
Fe			
Pb	Do not react with water or steam		Extracted from their ores by chemical reduction using <i>carbon</i> .
(H)		(included for comparison)	
Cu		Do <i>not</i> react with acids	Extracted by reduction with <i>carbon</i> <u>or</u> <i>hydrogen</i> .
Ag			Often found " native " and do not need to be extracted
Au			

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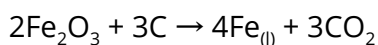
Metal Extraction

Metals have been used since the earliest civilizations. They are useful and prized. Some very unreactive metals, such as *silver*, *gold* and *platinum* can be found “**native**”. This means that they are uncombined with other elements and do not need to be extracted. All other metals are present as compounds — usually oxides and sulfides.

A **mineral** is a naturally occurring **metal compound**, for example, *halite* is NaCl, *zinc blende* is ZnS, *aragonite* is CaCO₃. An **ore** is a rock that contains enough of a metal compound or mineral to make it economically worth extracting. Iron ore (*haematite*) and aluminium ore (*bauxite*) are two very important ones.

Metals that are below carbon in the reactivity series can be extracted using *carbon* in a **chemical reduction** process. Their ores are heated with carbon in a furnace and carbon acts as a reducing agent. Carbon is cheap and is produced from charcoal. Charcoal comes from trees, so it is considered renewable.

The chemical reduction of iron ore is a **redox reaction**:

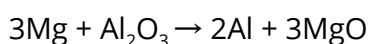


Fe₂O₃ is reduced to form **Fe**. (Fe is not reduced!)

C is oxidised to form CO₂

Carbon acts as a reducing agent as it releases the Fe from the Fe₂O₃. The furnace is hot enough to melt the metals. They are collected as molten liquids.

For metals above carbon in the reactivity series, this kind of cheap chemical reduction is not an option. It is *chemically* possible to produce aluminium by heating its ore with magnesium, but since magnesium is more expensive than aluminium this process is not an economically viable option:

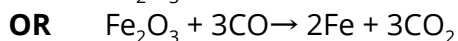
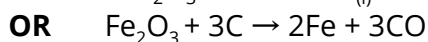
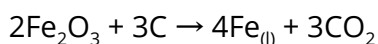


For all metals above carbon, a process called **electrolysis** is used. The metal ore is melted and a large electric current is passed through it. The **ions** in the compound are attracted to different electrodes and molten aluminium is collected. *For more information on electrolysis, see the section about it later in this booklet.*

Iron

Extraction

Iron is extracted from its **ore** in a large container called a **blast furnace**. The most common iron ore is *haematite*, which contains *iron (III) oxide* (also called *iron sesquioxide*), Fe₂O₃. A process of **chemical reduction** is used to extract the iron from its ore. In the reactivity series above, we can see that *carbon* is more reactive than iron. In hotter parts of the furnace, *carbon monoxide* will also reduce the iron oxide.



C is oxidised to form CO₂

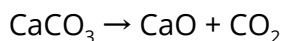
C is oxidised to form CO.

CO is oxidised to form CO₂

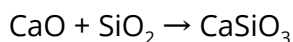
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In all cases, **Fe₂O₃** is reduced to form **Fe**. (the Fe is not reduced!) The iron melts and flows to the bottom of the furnace, where it can be tapped off.

To remove impurities, limestone (calcium carbonate) is added into the blast furnace. In the furnace, it thermally decomposes.



The calcium oxide reacts with acidic impurities such as silica (also known as quartz and found in sand) in the haematite to produce slag (calcium silicate).



The slag accumulates and can be tapped off from the furnace using a separate tap.

Preventing Rust

Iron rusts in the presence of both *oxygen* **and** *water*. There are many ways to protect the iron from rusting. To protect it, we must stop oxygen and water reaching the surface of the metal.

Sacrificial Protection

Iron can be worked into **stainless steel** to prevent it rusting. Stainless steel is just like regular steel in terms of hardness and its applications, except it contains *chromium* and *nickel*. The **chromium** is what makes it 'stainless'. It is an example of a **sacrificial metal**. Chromium reacts with oxygen in the air more readily than iron does, and forms a colourless layer of *chromium oxide* on the external sides of the stainless steel, preventing the iron underneath from losing its electrons and becoming oxidised. Even when scratched away, the chromium oxide layer will reform before the iron can rust. The **nickel** is added to increase the **hardness** of the steel and its ability to absorb considerable energy before cracking.

Galvanising

Galvanising is a method where the steel is coated in a thin layer of *zinc*. The zinc protects the iron in two ways: it stops oxygen and water reaching the metal underneath and it acts as a sacrificial metal.

Oiling and greasing

Oiling or greasing are the most common methods used to protect the iron from rusting, for example in bike chains. Greasing is often used in nuts and bolts as it also lubricates. Oil and water do not mix and so the oil forms a protective layer on the iron that prevents access to oxygen and water.

Painting

Paint protects the iron from rusting as a **barrier protection** method. One disadvantage of this is that paint chips and must be repainted every so often, and so is not as reliable as other methods, however, painting is a cheap solution in the short term.

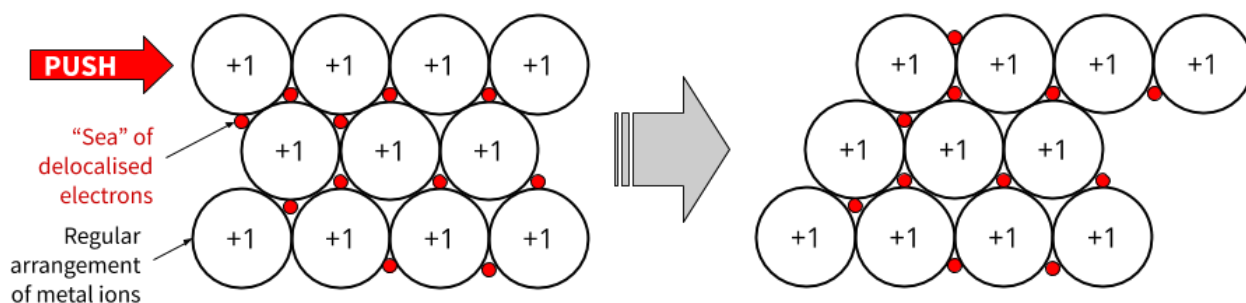
Coating with metal or plastic

Iron and steel objects may also be covered with a layer of metal or plastic. Food cans, for example, are plated with a thin layer of tin.

Alloys

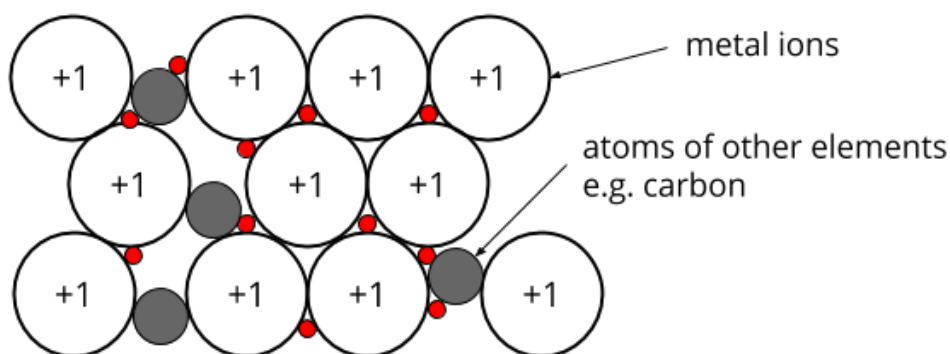
Alloys are **mixtures** of a metal with one or more other elements. They do **not** have a fixed ratio of atoms. There are hundreds of different alloys. Many are made for a certain purpose.

A malleable material can be hammered or bent into shape — it is **not brittle**. All pure metals are **malleable** and **soft** and will dent when hit hard. This is because of their regular arrangement and structure. The ions are able to slide over each other, whilst maintaining their **metallic bonding** (electrostatic attractions with the delocalised electrons). When given a shove, rows of atoms move but they still hold together.



For more information about metallic structure, see the pages about it earlier in this booklet.

When an alloy is made, other atoms of a **different size** become part of the structure.



This makes it far more difficult for the layers to slide over each other and they get stuck. Different sizes mean the layers cannot slide. The alloy is **harder** and **less malleable** than the pure metal.

Uses of metals and alloys

Metal	Uses	Properties
iron	<ul style="list-style-type: none"> Iron railings Manhole covers Iron gates 	<i>Iron</i> is hard . It is not as useful as steel (an iron alloy) and it rusts so things made from iron need to be painted or otherwise protected.
aluminium	<ul style="list-style-type: none"> Overhead cables window/door frames Tube trains Aeroplanes 	<i>Aluminium</i> is a very low-density metal. It forms an unreactive layer of oxide on its surface, which prevents it from being corroded.

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copper	<ul style="list-style-type: none">• Water pipes• Electric wiring• saucepans	<i>Copper</i> is very unreactive so it does not react with water or food. It is a very good conductor of electricity and heat.
mild steel (an iron alloy with 0.25% carbon)	<ul style="list-style-type: none">• Nails• Car bodies, ships• Girders and bridges	<i>Mild steel</i> has a very high tensile strength. It can be hammered into a number of shapes (it is malleable). CAUTION: mild steel rusts!
high carbon steel (0.6% to 1.2% carbon content)	<ul style="list-style-type: none">• Knives• Cutting tools	<i>High carbon steel</i> is harder and more resistant to wear than mild steel but is more brittle and less malleable.
stainless steel (alloy of iron, chromium and nickel)	<ul style="list-style-type: none">• Cutlery• Sinks• Draining boards	<i>Stainless steel</i> does not rust since the chromium is similar to aluminium and forms a corrosion-resistant layer on the surface.

Chemical Analysis

Chemical analysis is a series of test tube and other simple reactions that identify ions, gases and water.

Testing for water

Physical tests for pure water

- The boiling point of pure water is exactly 100°C
- The freezing/melting point of pure water is exactly 0°C

Chemical tests for the presence of water

- The addition of anhydrous copper (II) sulfate turns it from white to blue
- The addition of anhydrous cobalt (II) chloride turns it from blue to pink

Testing for gases

Gas	Test and observation
Hydrogen	Put a lit splint at the top of a test tube with hydrogen. Gives a squeaky pop.
Oxygen	Put a glowing splint into a test tube of the gas. The splint should ignite, because the concentration of oxygen is higher in the test tube than in the air.
Chlorine	Test with <i>damp</i> blue litmus paper. The paper turns red and then bleaches white. Also makes damp starch-iodide paper turn blue black.
Carbon dioxide	Bubble the gas through limewater [$\text{Ca}(\text{OH})_{2(aq)}$]. The limewater turns cloudy/milky due to calcium carbonate precipitate.
Ammonia	Extinguishes a lit splint. Test with <i>damp</i> red litmus paper. The litmus paper turns blue. A drop of hydrochloric acid into ammonia produces white NH_4Cl smoke.

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Identifying anions

Anion	Test and observation	Ionic equation
CO_3^{2-}	Add $\text{HCl}_{(\text{aq})}$ and test the gas given off with limewater. The limewater goes milky if CO_3^{2-} is present.	$\text{CO}_3^{2-}(\text{aq}) + 2\text{H}^+(\text{aq}) \rightarrow \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$
SO_4^{2-}	Add $\text{HCl}_{(\text{aq})}$ (to remove any CO_3^{2-} which would give an unclear result), followed by $\text{BaCl}_{2(\text{aq})}$. $\text{BaSO}_{4(\text{s})}$, a white ppt. is formed.	$\text{Ba}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{BaSO}_{4(\text{s})}$
Cl^-	To test for any halide ion, add $\text{HNO}_{3(\text{aq})}$ (to remove any CO_3^{2-}), then add $\text{AgNO}_{3(\text{aq})}$ <ul style="list-style-type: none"> Cl^- white precipitate AgCl Br^- cream precipitate AgBr I^- yellow precipitate AgI 	$\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}_{(\text{s})}$
Br^-		$\text{Ag}^+(\text{aq}) + \text{Br}^-(\text{aq}) \rightarrow \text{AgBr}_{(\text{s})}$
I^-		$\text{Ag}^+(\text{aq}) + \text{I}^-(\text{aq}) \rightarrow \text{AgI}_{(\text{s})}$

Flame Tests

Some metal cations give characteristic colours when put into an invisible Bunsen flame. This is because electrons, which are excited by the heat of the flame, move to higher energy levels. When they fall back down to their 'ground state' they emit the extra energy as electromagnetic radiation. If this happens within the visible region of the spectrum, we see a colour.

To do a flame test, an inert wire made of nichrome (or alternatively platinum) is cleaned using concentrated hydrochloric acid. The clean wire is then moistened with clean hydrochloric acid and dipped into the solid. *Alternatively*, wooden splints are soaked in ionic solutions overnight so that they become saturated. The splints are then held in the flame and the colour observed.

Cation	Flame Colour	Cation	Flame Colour
Li^+	red	Ca^{2+}	orange-red
Na^+	(persistent) yellow	Ba^{2+}	apple green
K^+	lilac	Cu^{2+}	blue-green

Another cation test

Sodium hydroxide solution can also be used to test for certain cations. This test works even if the cation concentration is very low. To do the sodium hydroxide test:

1. Dissolve your unknown compound in distilled water to make a solution
2. Add drops of $\text{NaOH}_{(\text{aq})}$
3. Look for a precipitate (a solid being formed)
4. If testing for NH_4^+ ions, warm the solution gently and test for ammonia gas.

Cation	Result with $\text{NaOH}_{(\text{aq})}$	Equation
$\text{Cu}^{2+}(\text{aq})$	Gelatinous blue precipitate	$\text{Cu}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq}) \rightarrow \text{Cu}(\text{OH})_{2(\text{s})}$
$\text{Fe}^{2+}(\text{aq})$	Gelatinous green precipitate	$\text{Fe}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq}) \rightarrow \text{Fe}(\text{OH})_{2(\text{s})}$

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$\text{Fe}^{3+}_{(\text{aq})}$	Gelatinous brown precipitate	$\text{Fe}^{3+}_{(\text{aq})} + 3\text{OH}^{-}_{(\text{aq})} \rightarrow \text{Fe}(\text{OH})_{3(\text{s})}$
$\text{NH}_4^{+}_{(\text{aq})}$	Ammonia gas produced when warmed	$\text{NH}_4^{+}_{(\text{aq})} + \text{OH}^{-}_{(\text{aq})} \rightarrow \text{NH}_{3(\text{g})} + \text{H}_2\text{O}_{(\text{l})}$

Quantitative Chemistry

Relative formula mass

The Periodic Table gives the relative **atomic** masses of the elements. The relative **formula** mass is the sum of the masses of the atoms in a compound's formula.

For example, the relative formula mass of potassium permanganate, KMnO_4 is

K	= 39
+ Mn	= 55
+ 4 × (O)	= 4 × 16 = 64
Total	= 158

Empirical and Molecular Formulae

Every chemical compound has an **empirical formula**. This is the simplest whole number ratio of atoms in the compound. Molecules also have a **molecular formula**. This is the *actual* number of atoms in the compound.

Name	Molecular formula	Empirical formula
butane	C_4H_{10}	C_2H_5
ethene	C_2H_4	CH_2
methane	CH_4	CH_4
hexane	C_6H_{14}	C_3H_7
ethanol	$\text{C}_2\text{H}_5\text{OH}$	$\text{C}_2\text{H}_5\text{OH}$
glucose	$\text{C}_6\text{H}_{12}\text{O}_6$	CH_2O
pentene	C_5H_{10}	CH_2

Experimental analysis can give the percentage composition of a compound in terms of its elements. We can use moles to determine the empirical formula from this. Example below.

An oxide of sulfur was found to be 50% sulfur and 50% oxygen. Determine its empirical formula.

Elements	S	O
either % or mass	50	50
RAM	32	16
Moles (= % ÷ RAM)	1.5625	3.125
Divide by smallest number to obtain the ratio	1	2

$\Rightarrow 1 : 2 \text{ ratio} \rightarrow \text{SO}_2$

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Moles

A mole is a unit of substance.

A mole of a substance is the amount of that substance that contains as many particles as there are atoms in exactly 12.00g of ^{12}C (carbon-12).

This (very large number) is known as Avogadro's constant and $\approx 6.02 \times 10^{23}$

To use the mole concept we use the **mole equation**

$$n = \frac{m}{M_r} \quad \text{where} \quad \begin{array}{l} n : \text{number of moles} \\ m : \text{mass of the substance (grams)} \\ M_r : \text{molar mass (g mol}^{-1}\text{)} \end{array}$$

Note that this equation is not given in the exam and you must learn it.

Moles and Reacting Masses

The mole concept enables us to calculate the exact masses of reactants and products in any chemical reaction. Demonstration by example.

10g of titanium (IV) carbonate was reacted with an excess of hydrochloric acid. What mass of carbon dioxide was produced?

1. Give a balanced symbol equation



2. Work out the moles of something (usually there is only one product/reactant for which you can work out its number of moles)

$$n(\text{Ti}(\text{CO}_3)_2) = \frac{10\text{g}}{48+60 \times 2} = \frac{10}{168} = 0.0595 \text{ moles}$$

3. Look at the ratio between the substance you have calculated the moles for and the substance the question is asking about.

Ratio of	$\text{Ti}(\text{CO}_3)_2$:	CO_2	
	1	:	2	(from the balanced symbol equation)
	0.0595	:	0.119	

4. Answer the question.

$$0.119 \text{ moles of } \text{CO}_2 \Rightarrow m = n \times M_r \Rightarrow 0.119 \times 44 = 5.24\text{g}$$

Percentage by mass calculations

Percentage by mass is a way in which compounds are compared. We use the formula

$$\% \text{ by mass} = \frac{\text{mass of element in a compound}}{\text{molar mass of a compound}} \times 100$$

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Percentage Yield

Virtually no reactions give a 100% yield. There are a number of reasons:

- product may be left in the experimental glassware
- starting material may be impure
- side reactions may occur, giving other products
- reaction may be reversible

To calculate yield we use the simple formula

$$\% \text{ yield} = \frac{\text{actual mass of product obtained}}{\text{expected mass from calculation}} \times 100$$

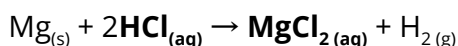
Moles and Gases

Using moles, we can calculate the *masses* of reactants and products in a given chemical reaction. Many chemical reactions involve gases, and gases are very difficult to weigh. Instead of weighing gases, we use their *volume*. Remember the following:

*1 mole of **any** gas has a volume of 24dm³ at room temperature and pressure*

Moles and Solutions

Many chemical reactions involve solutions



Solutions may be concentrated or dilute. Concentration is measured in mol dm⁻³ which is the number of moles of solute dissolved in 1dm³ (=1000cm³) of solution. To calculate concentration we use

$$n = cV \quad \text{where} \quad \begin{array}{l} n : \text{number of moles} \\ c : \text{concentration (mol dm}^{-3}\text{)} \\ V : \text{volume (dm}^3\text{)} \end{array}$$

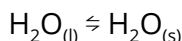
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Chemical Equilibria [NOT IN THE MOCK]

Some chemical reactions are reversible. They are shown with the reversible arrow as such ⇌

Any reversible reaction, if within a **closed system** (i.e. nothing enters or leaves and the temperature and pressure remain unchanged) is able to reach equilibrium.

Consider a mixture of water and ice at exactly 0°C:



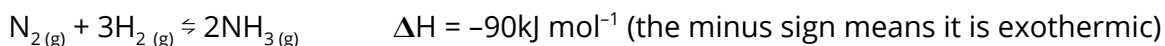
The water freezes to become ice. The ice melts to become water. At 0°C these processes occur at exactly the same rate, and so, the amount of ice and the amount of water are unchanged. (although you should never use the word “amount” in chemistry, the topic of *moles* will be covered in 6th form, and so for now we remain vague.) This system is said to be at equilibrium.

Manipulating equilibrium

All chemical equilibria are subject to a chemical law called “Le Chatelier’s Principle”. This is not on the specification, however, it makes understanding this topic far easier. In very simple terms, “Le Chatelier’s Principle” states that if you do something to an equilibrium it will do the opposite.

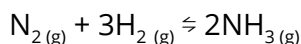
In the water/ice equilibrium, $\text{H}_2\text{O}_{(l)} \rightarrow \text{H}_2\text{O}_{(s)}$ is an exothermic process because bonds are being made. $\text{H}_2\text{O}_{(s)} \rightarrow \text{H}_2\text{O}_{(l)}$ is an endothermic process because bonds are being broken. If the temperature is increased for the $\text{H}_2\text{O}_{(l)} \rightleftharpoons \text{H}_2\text{O}_{(s)}$ equilibrium, the position of the equilibrium moves to the left in the endothermic direction, in order to oppose the increase in temperature and the ice melts.

Ammonia production



This reaction describes the **Haber Process**. This is the industrial production of ammonia. It is very important because the ammonia is then converted into agricultural fertilisers. In order to get a decent yield of ammonia, the reaction conditions are modified to force the equilibrium to move to the right. This is done in 5 ways:

1. **Keep adding N_2** — the equilibrium moves right to remove N_2
2. **Keep adding H_2** — the equilibrium moves right to remove H_2
3. **Keep removing NH_3** — the equilibrium moves right to make more NH_3
4. **Keep the reactor cold** — the equilibrium moves right in the *exothermic* direction to oppose the temperature change.



4 moles on this side ⇌ 2 moles of gas on other side

5. **Increase pressure** — the equilibrium moves right where there are fewer moles of gas.

What happens in reality?

N_2 and H_2 are constantly pumped into the reactor. The reactor is kept at 200atm pressure and 450 °C. The ammonia is condensed and removed from the reactor as it is formed.

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As you may have noticed, cost is also a factor at play here. We could use 2000atm pressure instead of 200atm, giving a better yield. However, as a compromise, we use 200atm, because the equipment for 2000atm costs more and the increase in yield is not worth it. The same applies for temperature, and we must compromise between rate and yield. We could use 20°C and get a 98% yield of ammonia but it would (quite literally) take centuries. Instead we use 450°C and get a 50% yield, but it takes mere minutes.

Electrochemistry [NOT IN THE MOCK]

Metals conduct electricity because they have delocalised electrons which are free to move and can carry a charge. Insulators such as wood and plastic and most non-metals do not conduct electricity because they have no mobile charged particles.

Ionic compounds do not conduct when solid but do when molten or in solution. In doing so, the compounds are broken down and this is called **electrolysis**. The ionic compound that undergoes electrolysis is called an **electrolyte**. In electrolysis, the electrolyte is split into its ions and the ions are able to move.

We can perform electrolysis of **melts** e.g. $\text{NaCl}_{(l)}$
Or electrolysis of **solutions** e.g. $\text{NaCl}_{(aq)}$

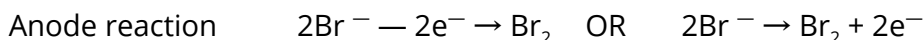
Electrolysis of molten compounds

As an example we take PbBr_2 . The Pb^{2+} and Br^- ions are attracted to the electrodes.

Pb^{2+} goes to the cathode because Pb^{2+} is a cation. (and **cations** are **pussy**tive) When the Pb^{2+} gets to the cathode it is reduced to Pb. This is a reduction reaction because Pb^{2+} is gaining electrons.

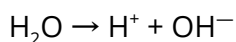


Br^- ions will move to the anode. Br^- is an anion. When Br^- ions reach the anode, they lose electrons to form bromine molecules. This is an oxidation reaction because Br^- is losing electrons.

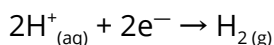


Electrolysis of aqueous solutions

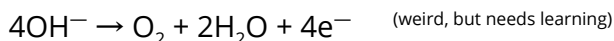
Many ionic compounds are soluble. Their aqueous solutions can undergo electrolysis *but* water molecules can form ions too.



This means we have two cations going to the cathode and two anions heading for the anode. If H^+ from water reacts at the cathode, hydrogen gas is formed.



If OH^- from water reacts at the anode, oxygen gas is formed.



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Rules exist that allow us to predict what happens at the electrodes in these cases

At the cathode, whichever element is *lower* in the reactivity series will be released. The more reactive element will stay in solution.

At the anode, if a halide is present (F^- , Cl^- , Br^- or I^-) it will always react. Otherwise OH^- reacts to give O_2 .