

Part I

Unit 1: Chemical fundamentals – structure, properties, and reactions

Section 1

Properties and structure of atoms

1.1 Atomic structure

1.1.1 What is matter?

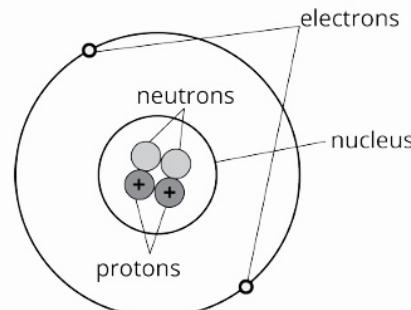
Matter is anything that has **mass and takes up space**. So, everything around you is matter! This book, yourself, and the air that you are breathing are all examples of matter. An **atom** is made up of a *very small* amount of matter!

Everything is made up of atoms. Scientists had to try for many years to work out what the inside of an atom looks like, because atoms cannot be seen by the human eye. As you may already know, atoms are made up of **subatomic particles**, including protons, neutrons, and electrons. I've summarised some information about these particles in the table below.

	Proton	Neutron	Electron
Symbol	p	n	e
Mass	$1.673 \times 10^{-27} \text{ kg}$	$1.675 \times 10^{-27} \text{ kg}$	$9.109 \times 10^{-31} \text{ kg}$
Charge	+1	0	-1

As you can see in the table above, protons are positively charged, neutrons have no charge, and electrons are negatively charged. Protons and neutrons have similar masses and are much heavier than electrons.

Protons and neutrons are found inside the small, positively charged **nucleus**. Therefore, protons and neutrons are called **nucleons**. Meanwhile, electrons are found outside of the nucleus in what we call the **electron cloud**. A simplified model of an atom is pictured here on the right. This diagram is not to scale, because in real life the nucleus is much smaller in comparison to the size of the atom.



Fun fact: the nucleus is around 10,000 times smaller than the size of an atom! Despite being very small, the nucleus is **really dense**, as it contains approximately 99.97% of the atom's mass.

You may remember from previous years that **opposite charges attract**. Therefore, the positively charged protons are attracted to the negatively charged electrons. This is called **electrostatic attraction**, which is what binds the electrons to the nucleus.

KEY POINT :

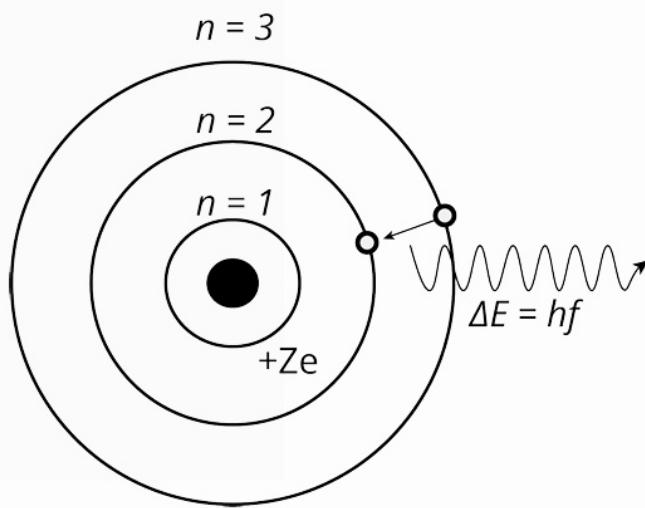
An atom contains the same number of protons and electrons. Since protons and electrons have equal but opposite charges, atoms are **neutrally charged** because the charges cancel each other out.

EXTENSION :

Since we know that like charges repel each other, and we also know that neutrons have no charge and protons are positively charged, you may be wondering why protons don't repel each other in the nucleus. This is because of the **strong nuclear force**. This is an attractive force which binds neutrons and protons together, and is integral in keeping the nucleus together, because without this force, nuclei wouldn't exist.

1.1.2 Electron shells

In the **Bohr model** of the atom, the electrons are grouped into **energy levels** (also known as **electron shells**). This is shown by the diagram below.



As you can see in the diagram, electrons which are in the same shell are the same distance away from the nucleus. Electrons in shells closer to the nucleus have lower energy, whereas electrons in the outer shells have higher energy. You can see in the diagram above that the lowest energy shell is labelled $n = 1$, whereas the highest energy shell is labelled $n = 3$.

Electrons can move up into a higher energy level by absorbing energy, and electrons can move down into a lower energy level by releasing energy. When electrons move into a higher energy shell (which is called the **excited state**), we can describe the electrons as being ‘excited.’ Whereas, the lower energy shell that the electron is originally found in is called the **ground state**.

In the diagram above, we can observe that energy is being released as an electron moves from the $n = 3$ electron shell to the $n = 2$ electron shell.

Different electron arrangements can release different energies of light. We can observe this because the light is different colours. The released light produces an **emission spectrum**. Don’t worry, I will talk more about this later (see page 26).

1.1.3 Electron configurations using the Bohr Model

An **electron configuration** describes the location of electrons in shells for a given atom.

Each shell has a different maximum number of electrons that it can hold. You can calculate this using the following formula:

$$\text{Maximum number of electrons in the } n^{\text{th}} \text{ shell} = 2n^2$$

Some examples of the maximum number of electrons in the 1st, 2nd, 3rd, and 4th shells are given below.

Shell (n)	Maximum number of electrons ($2n^2$)
1	$2 \times 1^2 = 2$
2	$2 \times 2^2 = 8$
3	$2 \times 3^2 = 18$
4	$2 \times 4^2 = 32$

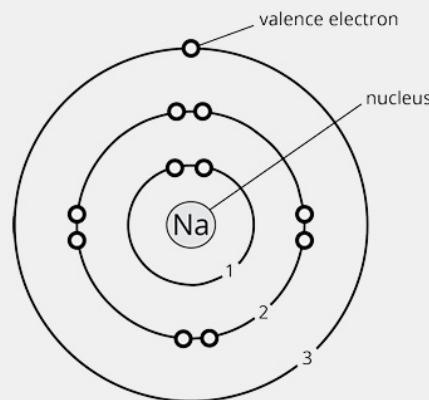
The arrangement of electrons in shells around a nucleus can be shown by a **Bohr diagram**. You may have learnt this in previous years, but it's good to revise it. The below diagram shows electrons orbiting the nucleus like the planets orbit the Sun.

SAMPLE :

Let's look at a Bohr diagram of a sodium atom.

Sodium (which is represented by the symbol Na) has 11 protons, 12 neutrons, and 11 electrons. I will talk more about how to work this out later using the periodic table. As you can see in the diagram, there are two electrons in the 1st shell, eight electrons in the 2nd shell, and one electron in the 3rd shell.

This one electron in the outer shell is known as a **valence electron** (discussed in the next section!). This configuration occurs because the 1st shell can only hold two electrons and the 2nd shell can only hold eight electrons. Electron shells are filled from the inner shells to the outer shells. Therefore, we can write the electron configuration of sodium as 2,8,1.



1.1.4 Introduction to the periodic table

Before you start learning about the periodic table, there are a few terms that you need to know. I've listed these below, but you may already know about some of them:

- **Element:** contains only one type of atom (which have the same number of protons).
 - E.g. sodium (Na) and oxygen gas (O_2 , which contains two oxygen atoms bonded together).
- **Compound:** contains two or more elements bonded together.
 - E.g. water (H_2O , which contains two hydrogen atoms bonded to an oxygen atom).
- **Molecule:** contains two or more atoms bonded together.
 - E.g. water (H_2O) and oxygen gas (O_2).

KEY POINT :

Make sure that you don't get **molecules** confused with **compounds**! This is important because molecules aren't necessarily compounds. For example, oxygen gas is an *element* (as it only contains oxygen atoms) and it is a *molecule* (as it is made up of two atoms), but it is not a *compound*.

The periodic table displays the 118 elements in order of increasing **atomic number**, which is the number of protons in the atom. The elements are displayed using their chemical symbols on the periodic table (e.g. oxygen is represented as O).

The columns on the periodic table are called **groups**. Elements in the same group have a similar outer shell electron configuration (and therefore similar chemical and physical properties).

The atom's outer shell can be called the **valence shell**, and the electrons in this shell are called **valence electrons**. These electrons are involved in chemical reactions because they require the least amount of energy to be removed. Later I will talk about bonding and how atoms can lose, gain, and share electrons (starting on page 17). Atoms tend to do this so they can achieve **eight valence electrons** – this is known as the **octet rule**.

As you can see in the periodic table on the following page, there are 18 groups (columns). The elements are grouped into columns based on their number of valence electrons. I've listed the groups in the table on the following page.

Group	Description
Group 1	Elements have one valence electron and are known as the alkali metals .
Group 2	Elements have two valence electrons and are known as the alkaline earth metals .
Group 3–12	Elements have varying numbers of valence electrons and are known as the transition metals .
Group 13	Elements have three valence electrons.
Group 14	Elements have four valence electrons.
Group 15	Elements have five valence electrons.
Group 16	Elements have six valence electrons.
Group 17	Elements have seven valence electrons and are known as the halogens .
Group 18	Elements have eight valence electrons (a full outer shell) with the exception of helium, which only has one electron shell so has two valence electrons. These elements are known as the noble gases .
Two bottom rows	The two rows at the bottom which are separated from the main part of the table are called the lanthanides (elements with atomic numbers between 57 and 71) and actinides (elements with atomic numbers between 89 and 103).

KEY POINT :

An easy way to remember the table above is that groups 1, 2, and 13–18 have the same number of valence electrons as their final digit.

The rows on the periodic table are known as **periods** and are labelled 1–7. Elements in each period have a valence shell with the same number as the period. For example, elements in period 5 will have valence electrons in the 5th shell.

Periodic table of the elements

The periodic table displays the following data:

- Periods:** Rows 1 through 7.
- Groups:** Columns 1 through 18.
- Elements:** Hydrogen (H), Helium (He), Lanthanides (La-Lu), Actinides (Ac-Lr), and various transition metals.
- Properties:** Atomic number, symbol, relative atomic mass, and some isotopic data (e.g., ¹³⁸La, ¹⁴⁰Ce).

Groups are numbered according to IUPAC convention 1–18.

*Values in brackets are for the isotope with the longest half-life.

Source: https://www.qcaa.qld.edu.au/downloads/senior-qce/sciences/snr_chemistry_19_formula_data_book.pdf

KEY POINT :

In the external exam you will have access to a periodic table in the formula and data book, so you don't need to memorise the elements' atomic numbers, mass numbers, or symbols.

1.1.5 Nuclear symbol notation

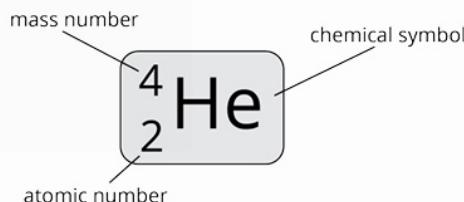
The periodic table displays each element's **atomic number** and **mass number**. As mentioned previously, the atomic number is the number of protons in an atom. Atomic numbers have the symbol Z, and all atoms that are the same element have the same atomic number. Mass numbers are represented by the symbol A and are equal to the number of protons plus the number of neutrons.

KEY POINT :

To find the number of neutrons in an element, use the mass number in the formula below:

$$\text{Number of neutrons} = \text{mass number (A)} - \text{atomic number (Z)}$$

Nuclear symbol notation shows an atom's atomic number and mass number. An example of helium (which has two protons and two neutrons) is given below:



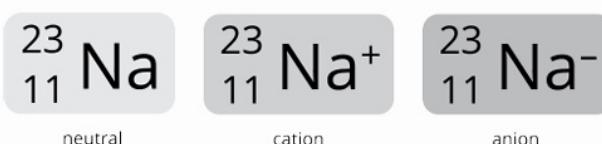
QCAA also wants us to consider this notation for ions and isotopes. **Ions** are charged atoms. They can be positively charged (these are called **cations**) or negatively charged (these are called **anions**).

KEY POINT :

You can remember this by thinking that **cations** are **pawsitive!**

If an atom loses an electron, then it becomes a cation (as it has fewer electrons than are negatively charged). Whereas, if an atom gains an electron, then it becomes an anion (as it has excess electrons). For example, a sodium (Na) atom has 11 electrons, whereas a sodium cation (Na^+) has 10 electrons and a sodium anion (Na^-) has 12 electrons.

Since nuclear symbol notation only includes information about the number of protons and neutrons, the only thing that differentiates a neutral element from its ionised form is the superscript (the + or - symbols above that show the charge). For example, the nuclear notation for a neutral sodium atom, a sodium cation and a sodium anion are given below.



Isotopes are atoms which are the same element (as they have the same atomic number) but they have different numbers of neutrons (as they have different mass numbers). Carbon has fifteen known isotopes. For example, carbon-12 has six protons and six neutrons, whereas carbon-13 has six protons and seven neutrons. Make sure you remember that isotopes are named by writing '[element]-[mass number]'

It's also important to remember that isotopes of an element have the **same chemical properties** but **different physical properties**. For example, isotopes have different masses (which is a physical property) due to the different numbers of neutrons. The chemical properties are the same because these are determined by the electron configurations (which are the same for all isotopes of an element). Since isotopes have different mass numbers, they have different nuclear symbol notations. For example, let's look at carbon-12 and carbon-13.



Note that isotopes can also be represented as **A Symbol** (e.g. carbon-12 can be written as ^{12}C).

1.1.6 Orbitals and electron configuration

Now that we've looked at the nucleus, let's spend some more time thinking about the arrangement of electrons. Scientists found that electron shells contain **subshells** and **orbitals** and this model of the atom is called the **quantum mechanical model**.

KEY POINT :

Subshells are separate energy levels within a shell. There are four types of subshells: **s**, **p**, **d**, and **f**. Make sure that you know the difference between subshells and orbitals! Orbitals are regions within subshells, which hold two electrons.

Each subshell can only contain a certain number of electrons. This is because different subshells have various numbers of orbitals, and this is described in the table below:

Subshell	Number of orbitals	Max. number of electrons
s	1	2
p	3	6
d	5	10
f	7	14

Some shells don't contain all the types of subshells (s, p, d and f), so I've summarised this information for the first four shells in the table below:

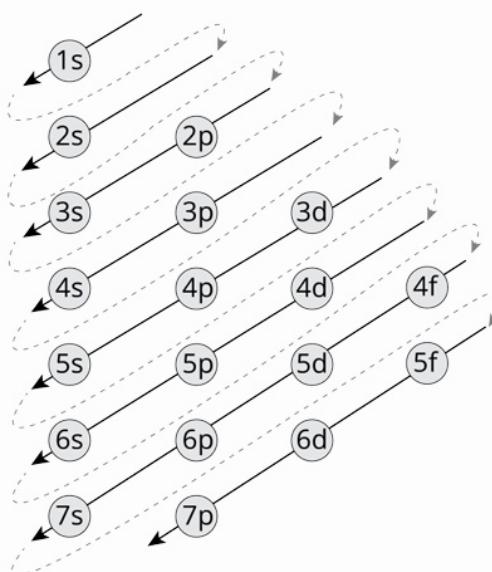
Shell	Subshells present
1	s
2	s, p
3	s, p, d
4	s, p, d, f

Earlier on page 3, we covered how to write electron configurations using the Bohr model. Now we need to write electron configurations using the subshell model. When we do this, there are three main rules to consider: the **Aufbau principle**, the **Pauli exclusion principle**, and **Hund's rule**.

The **Aufbau principle** states that shells and subshells are filled from the lowest energy to highest energy levels. This means that we fill the inner shells before the outer shells. The order of filling the subshells is displayed in the diagram on the right.

Note that the shell number is listed first and then the subshell (e.g. 1s means the 1st shell, subshell s). When we write the electron configuration including subshells, we use a similar method.

From the diagram above, we can tell that the 2p subshell is filled before the 3s subshell, because 2p subshell is of lower energy than the 3s subshell.



SAMPLE :

Let's go through how to write an electron configuration including subshells using the example of chlorine:

1. The first thing that you do is you check how many electrons chlorine has. You can do this by using the periodic table and finding the atomic number. Chlorine has 17 protons and therefore has 17 electrons.
2. Next you must work out how many electron shells chlorine has. The first shell can hold two electrons (in the s subshell), the second shell can hold eight electrons (in the s and p subshells), so the third shell must contain seven electrons (in the s and p subshells). You can work out the order that the subshells are filled by looking at the previous diagram.
3. Now you can write out this information: $1s^2 2s^2 2p^6 3s^2 3p^5$. The coefficients signify the shells (e.g. 1 means the first shell). The letters describe the subshells. The number of electrons in each subshell are denoted by the superscripts (e.g. the superscript 2 in $1s^2$ signifies that there are two electrons in the s subshell in the first shell).

KEY POINT :

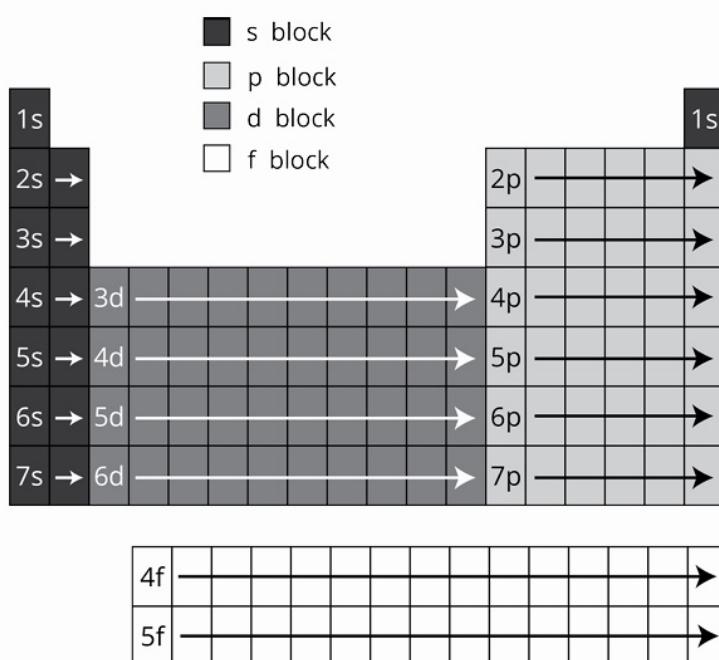
The QCAA syllabus says that you need to be able to write the electron configurations for all of the elements up to krypton ($Z = 36$) on the periodic table. Krypton contains the p subshell in the fourth electron shell. So, make sure you memorise the diagram on the previous page up to 4p!

It is very important to recognise that there are **exceptions** to the Aufbau principle. The electron configurations for chromium and copper do not follow this principle:

	Actual electron configuration	Predicted electron configuration using the Aufbau principle
Chromium (Cr)	$[Ar]3d^5 4s^1$	$[Ar]3d^4 4s^2$
Copper (Cu)	$[Ar]3d^{10} 4s^1$	$[Ar]3d^9 4s^2$

Both elements are exceptions because the 3d subshell is filled before the 4s subshell, which doesn't follow the Aufbau principle (this is shown in the diagram above). The reason for this is that half-filled and fully filled subshells are more stable.

A quicker way of writing electron configurations is using the periodic table, as it shows the highest energy subshell for each element:



QCAA also specifies that you need to know about **condensed electron configuration**. Again, I'll explain this using the example of chlorine.

SAMPLE :

The first thing that you need to do is look at the periodic table and find the last noble gas before chlorine (which is neon).

Then by looking at the periodic table we can tell that neon is $1s^2 2s^2 2p^6$ and chlorine is $1s^2 2s^2 2p^6 3s^2 3p^5$.

Finally, we combine the information and remove the subshells in common to write the condensed electron configuration for chlorine: [Ne] $3s^2 3p^5$.

1.1.7 Orbital diagrams

Now that we've looked at the Aufbau principle, we need to discuss the other two principles (Pauli exclusion principle and Hund's rule) and orbital diagrams.

Orbital diagrams show specific details of electrons in each orbital (in particular, it displays their spin).

The Pauli exclusion principle states that electrons in the same orbital have opposite spin. In an orbital diagram, boxes are used to represent orbitals and arrows symbolise electrons. The arrows for two electrons in an orbital are in opposite directions to signify that they have opposite spin (as one electron spins-up and the other electron spins-down).

	1s	2s	2p
Boron: $1s^2 2s^2 2p^1$	↓↑	↓↑	1
Carbon: $1s^2 2s^2 2p^2$	↓↑	↓↑	1 1
Nitrogen: $1s^2 2s^2 2p^3$	↓↑	↓↑	1 1 1
Oxygen: $1s^2 2s^2 2p^4$	↓↑	↓↑	↓↑ 1 1

You can see in the diagram above that for nitrogen there is one electron in each orbital. This is due to the **Hund's rule**, which states that one electron must occupy each orbital in a subshell before any orbital is doubly occupied.

1.2 Periodic table and trends

Now that we've discussed the structure of atoms, let's talk about how elements appear on the periodic table.

1.2.1 Effective nuclear charge

Before we start looking at trends on the periodic table, it's important that you understand what effective nuclear charge is.

KEY POINT :

Effective nuclear charge = number of protons – number of inner shell electrons

Across each period (row) from **left to right**, as the number of protons increases but the number of electron shells remain constant, **effective nuclear charge increases**.

SAMPLE :

Let's look at the effective nuclear charges for sodium (Na) and magnesium (Mg):

Sodium is in group 1 and period 3. It has 11 protons and 11 electrons. Therefore, sodium has one valence electron, and 10 electrons in the two inner shells:

$$\begin{aligned}\text{Effective nuclear charge} &= 11 - 10 \\ &= 1\end{aligned}$$

Magnesium is in group 2 and period 3. It has 12 protons and 12 electrons. Therefore, magnesium has two valence electrons, and 10 electrons in the two inner shells:

$$\begin{aligned}\text{Effective nuclear charge} &= 12 - 10 \\ &= 2\end{aligned}$$

Therefore, sodium (1) has a smaller effective nuclear charge than magnesium (2). This demonstrates that as you go along a period effective nuclear charge increases.

Down each group, as both the number of electron shells and protons increase, effective nuclear charge remains constant.

SAMPLE :

Let's contrast the effective nuclear charges for sodium (Na) and potassium (K):

We previously calculated the effective nuclear charge for sodium (1), which is in group 1 and period 3. So, sodium has 11 protons, 1 valence electron, and 10 electrons in the two inner shells.

Potassium is below sodium on the periodic table, as it is in group 1 and period 4. Therefore, potassium has 19 protons and 19 electrons. Potassium has 1 valence electron, and 18 electrons in the three inner shells:

$$\begin{aligned}\text{Effective nuclear charge} &= 19 - 18 \\ &= 1\end{aligned}$$

Therefore, potassium and sodium have the same effective nuclear charge (1). This supports the trend that down each group effective nuclear charge remains the same.

Now that you understand effective nuclear charge, let's look at the trends for the attraction between the positively charged nucleus and the negatively charged valence electrons.

Across each period, the valence electrons become more attracted to the nucleus, as the number of protons in the nucleus increases. This is because the nucleus' positive charge is greater, which attracts the negatively charged electrons.

For example, let's consider sodium and magnesium which are next to each other in period 3. Sodium has an atomic number of 11 and magnesium has an atomic number of 12. Since magnesium has one more proton in its nucleus than sodium, magnesium's valence electrons feel a stronger attraction to the nucleus than sodium's valence electrons.

Whereas, down a group, there is reduced electrostatic attraction between the valence electrons and the nucleus. This is because the inner electron shells 'shield' the valence electrons by reducing the outer electrons' attraction to the nucleus. This is called the **shielding effect**. Also, as there are more electron shells, there is increased repulsion from the inner-shell electrons.

For example, let's think about sodium and potassium, which are vertically next to each other, as sodium is in period 3 and potassium is in period 4. Therefore, sodium has two inner electron shells, whereas potassium has three inner electron shells. So, the valence electrons in potassium are 'shielded' more from the nucleus' positive charge than the valence electrons in sodium. Therefore, the valence electrons in sodium have increased electrostatic attraction to the nucleus than the valence electrons in potassium.

KEY POINT :

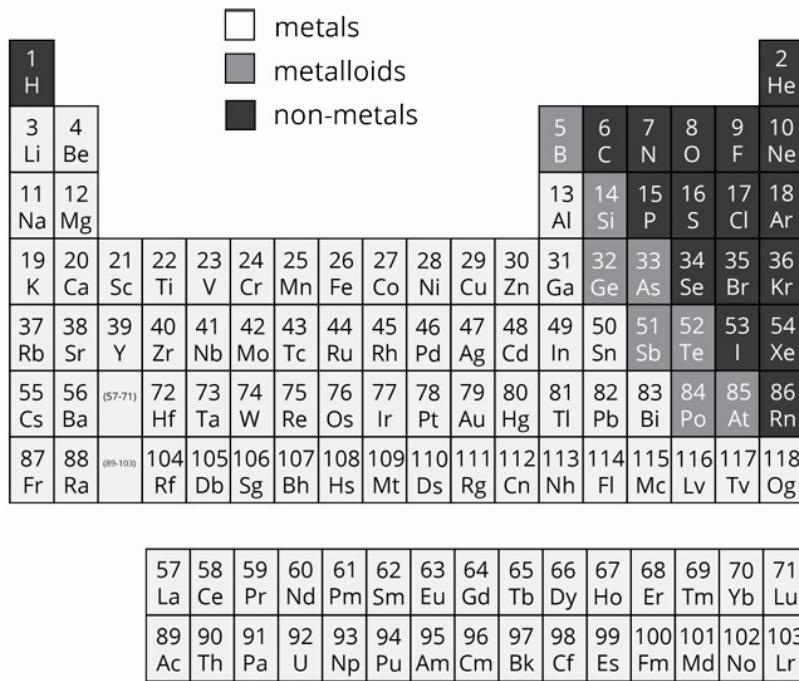
In the next chapters I will discuss other trends on the periodic table. When explaining these trends, make sure that you link the reason back to the attraction of the valence electrons to the nucleus.

I've summarised the trends from this chapter in the table below:

Attraction between nucleus and valence electrons		Effective nuclear charge
Across each period (left to right)	Increases	Increases
Down each group	Decreases	Remains the same

1.2.2 Metals, non-metals and metalloids

Now that you understand effective nuclear charge, let's discuss some trends on the periodic table. We'll first start by looking at the different types of elements. The elements on the periodic table are either classified as metals, non-metals or metalloids. These are marked on the periodic table below:



Let's start by looking at **metals**. As you can see in the diagram, the metals are located to the left of the table (except for hydrogen which is a non-metal).

KEY POINT :

Metals tend to lose their valence electrons in order to obtain eight valence electrons (remember the octet rule) by forming cations (which are positively charged).

- **Alkali metals:** lose their one valence electron to form cations which have a charge of +1.
- **Alkaline earth metals:** lose their two valence electrons to form +2 charged cations.
- **Transition metals:** these elements act differently to other elements because they form variously charged ions. For example, cobalt can form both Co^{2+} and Co^{3+} . Notice that the charge of +2 is written as a superscript of $^{2+}$. (You don't need to fully understand the transition metals, as this is more advanced chemistry!)
- **Group 13:** lose their three valence electrons to form +3 charged cations.

To measure the **reactivity of metals**, we look at how easily they lose electrons (to form cations). Across each period, the reactivity of metals decreases. This makes sense because metals are found on the left-hand side of the table, whereas the non-metals are found on the right-hand side. As we go across the rows, electrons are *more* attracted to the nucleus. Therefore, it is harder for these atoms to lose electrons. As you go down each group, the reactivity of metals increases. This is because the valence electrons feel less attracted to the nucleus due to the **shielding effect** and therefore these valence electrons are easier to remove.

Now that we've considered metals, let's have a look at **non-metals**.

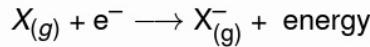
KEY POINT :

Non-metals react by gaining electrons in order to form a stable octet arrangement (eight valence electrons).

- **Group 15:** have five valence electrons, so they gain three electrons. This allows them to have a full outer shell of eight electrons. This forms an anion with a charge of -3.
- **Group 16:** have six valence electrons, so they gain two electrons (to have a full outer shell). This forms an anion with a charge of -2.
- **Halogens:** have seven valence electrons, so they gain one electron (to have a full outer shell). This forms an anion with a charge of -1.
- **Noble gases:** already have full outer shells (stable electron arrangements), so they don't form ions. Therefore, noble gases are mostly **inert** (which means that they don't react).

The **reactivity of non-metals** is measured by their tendency to gain electrons to form anions. Across each period, the reactivity of non-metals increases, as electrons are more attracted to the nucleus, so they gain electrons more easily. Down each group, the reactivity of non-metals decreases, as electrons are less attracted to the nucleus due to the shielding effect.

The reactivity of non-metals can be related to **electron affinity**. Electron affinity measures the energy released when an atom gains an electron:



If a non-metal is more reactive then its electron affinity will be higher (as it releases more energy). So, across each period, electron affinity increases, and down each group, electron affinity decreases.

EXTENSION :

Fluorine is an exception to the trend for electron affinity. This is because as you go down the halogens, fluorine has a lower electron affinity than chlorine, even though it is more reactive than chlorine. This occurs because the reaction of fluorine involves other transformations which release energy. You probably don't need to know this but it's important to understand that there is an exception to the rule!

KEY POINT :

Metalloids (which are also called semi-metals) display properties of both metals and non-metals. The metalloids form a 'stairway' shape on the periodic table (see the image on the previous page). The metalloids include boron, silicon, geranium, arsenic, antimony, tellurium, polonium, and astatine.

You may have noticed that group 14 contains metals, metalloids, and non-metals. So, carbon (which is a non-metal) forms anions with a charge of -4. Silicon and geranium (which are metalloids) can form anions and cations with charges of -4 and +4. Tin and lead (which are metals) form cations with a charge of +4.

I've summarised the trends from this chapter in the table below.

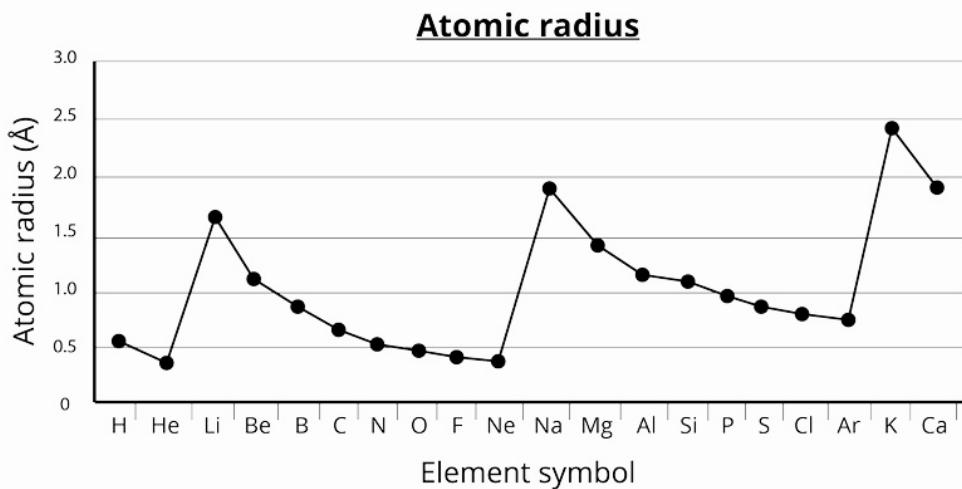
	Reactivity of metals	Reactivity of non-metals	Electron affinity
Across each period (left to right)	Decreases	Increases	Increases
Down each group	Increases	Decreases	Decreases

1.2.3 Atomic radius

Now let's discuss the trends for atomic radius. **Atomic radius** is the distance from the nucleus to the valence electrons. Across each period, atomic radius decreases. This is because the valence electrons are more attracted to the nucleus, so they are bound closer to the nucleus.

Down each group, atomic radius increases, as there are more electron shells. So, due to the shielding effect, electrons are less attracted to the nucleus, and therefore they are further away from the nucleus.

You can see these trends in the graph below. For example, down group 1, which consists of hydrogen (H), lithium (Li), sodium (Na), and potassium (K), atomic radius increases. The graph also shows, for example, that across period 2, atomic radius decreases from lithium (Li) to neon (Ne).



I've summarised the trends for atomic radius in the table below:

	Atomic radius
Across each period (left to right)	Decreases
Down each group	Increases

1.2.4 Ionic radii

Now that we've covered atomic radius, let's have a look at ionic radius. This describes the radius of an atom's ion. I'll explain the trends using a few examples.

Let's consider sodium (Na) and magnesium (Mg).

SAMPLE :

Sodium and magnesium are both in period 3 and are in groups 1 and 2 respectively. Therefore:

- Sodium has the electron configuration of: $1s^2 2s^2 2p^6 3s^1$.
- Magnesium has the electron configuration of: $1s^2 2s^2 2p^6 3s^2$.

These metals lose electrons and become cations to achieve an outer shell full of eight electrons. Sodium loses one electron to become Na^+ ($1s^2 2s^2 2p^6$). Whereas, magnesium loses two electrons to also become Mg^{2+} ($1s^2 2s^2 2p^6$). Na^+ and Mg^{2+} have the same electron configuration, which is the same as neon's.

Therefore, Na^+ and Mg^{2+} are **isoelectronic**, which is when two ions or atoms have the same electron configuration.

Since Na^+ and Mg^{2+} have the same electron configuration, Mg^{2+} has a smaller radius because it contains more protons, so the electrons are more attracted to the nucleus.

Now let's consider, sulphur (S) and chlorine (Cl).

SAMPLE :

Sulphur and chlorine are both in period 3 and groups 16 and 17 respectively.

- Sulphur has an electron configuration of: $1s^2 2s^2 2p^6 3s^2 3p^4$.
- Chlorine has an electron configuration of: $1s^2 2s^2 2p^6 3s^2 3p^5$.

The non-metals gain electrons to form anions which have a full outer shell, so sulphur becomes S^{2-} and chlorine becomes Cl^- .

S^{2-} and Cl^- are isoelectronic as they have the same electron arrangement as argon (Ar) of $1s^2 2s^2 2p^6 3s^2 3p^6$. Therefore, since Cl^- has more protons than S^{2-} , we know that Cl^- has a smaller radius.

So, across each period from left to right, ionic radius decreases for the elements that form cations and then for the elements that form anions.

Now we need to compare the ionic radius for cations and anions in the same period. For example, let's think about Mg^{2+} and S^{2-} . Of these two, S^{2-} has more electron shells, so it is larger than Mg^{2+} . Therefore, cations are smaller than anions in the same period.

Down each group, ionic radius increases. Like atomic radius, this is caused by the increased number of electron shells.

I've summarised the trends for ionic radius in the table below:

	Ionic radius
Across each period (left to right)	Decreases for elements that form the same type of ion (i.e. decreases for cations across each period and then decreases for anions)
Down each group	Increases

1.2.5 Ionisation energy

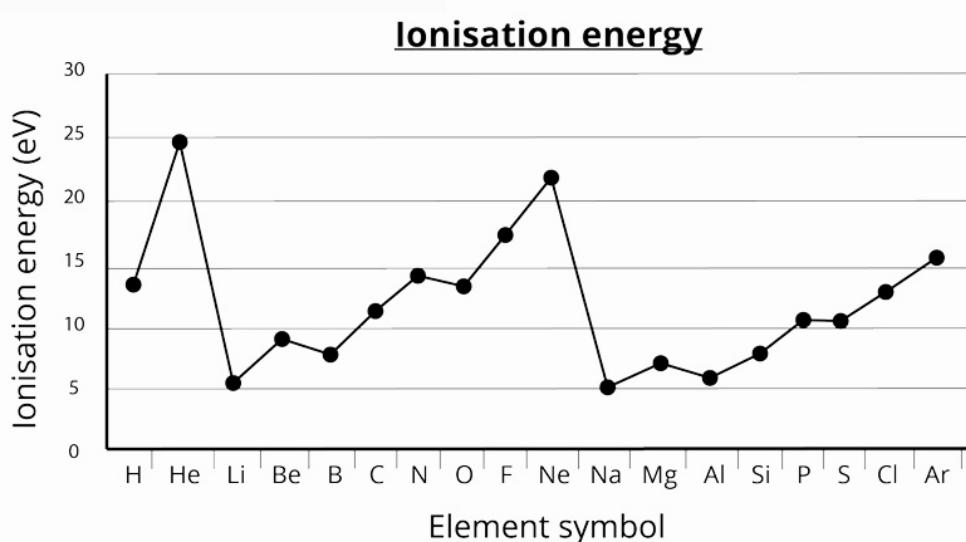
KEY POINT :

First ionisation energy is the amount of energy needed to remove one mole of electrons from one mole of gaseous atoms.

Across each period from left to right, first ionisation energy increases. This is because the electrons are bound more strongly to the atom, so more energy is required to overcome the attraction between the electron and the nucleus in order to remove the electron.

Down each group, the first ionisation energy decreases. This is because there are more electron shells, so the valence electrons are more ‘shielded’ from the nucleus. Therefore, less energy is required to remove an electron.

These trends are displayed in the graph below. You can see that down group 18 (consisting of helium, neon and argon) ionisation energy decreases. Whereas, across period 2, ionisation energy increases from lithium to neon.



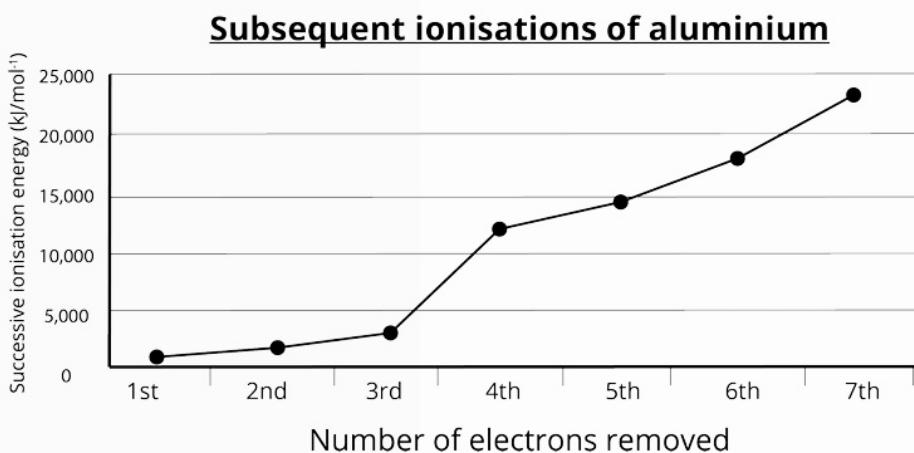
However, there are exceptions to the trends! For example, boron (B) has a lower first ionisation energy than beryllium (Be). This is due to the nature of the orbitals. Boron has an electron configuration of $1s^2 2s^2 2p^1$, whereas beryllium has an electron configuration of $1s^2 2s^2$. Using the trends on the periodic table, beryllium should have a smaller first ionisation energy than boron. However, this is not the case, because when beryllium is ionised, an electron is removed from a full s-subshell. It is harder to remove an electron from a fully filled subshell, because they are stable, so therefore it requires more energy. Whereas, when boron is ionised, an electron is removed from a p subshell that isn't full. Therefore, it is easier to remove an electron from boron, so it has a lower first ionisation energy.

Likewise, another exception to the trend is that nitrogen (N) has a larger first ionisation energy than oxygen (O). This is because when nitrogen is ionised an electron is removed from a half-filled p-subshell ($1s^2 2s^2 2p^3$), but this is not the case for oxygen ($1s^2 2s^2 2p^4$). It is harder to remove electrons from half-filled subshells because they are more stable. Therefore, oxygen has a smaller first ionisation energy than nitrogen, even though this does not follow the trend.

I've summarised the trends for first ionisation energy in the table below:

First ionisation energy	
Across each period (left to right)	Increases
Down each group	Decreases

Next, let's consider successive ionisation energies. For example, **second ionisation energy** is the amount of energy needed to remove one mole of electrons from one mole of the gaseous anion (of charge -1). The graph below displays the trend for the successive ionisation energies for aluminium.



You can see that the successive ionisation energy continues to increase for aluminium. As an aluminium atom continues to be ionised, more electrons are removed, so there is increased attraction between the electrons and the nucleus. Therefore, more energy is required to counteract the increased electrostatic attraction in order to remove an electron.

There is a large jump between the 3rd and 4th ionisation energies for aluminium. This is because once three electrons are removed from an aluminium atom it is stable, as it has a full outer shell of eight valence electrons, so lots of energy is required to remove a fourth aluminium electron.

1.2.6 Electronegativity

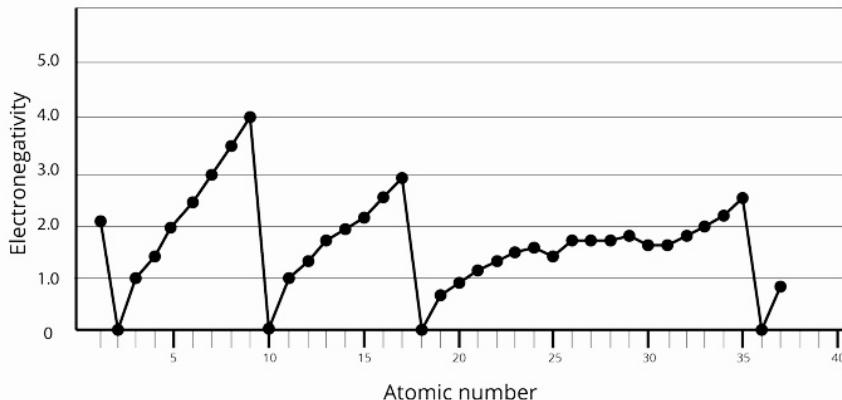
KEY POINT :

Electronegativity describes an atom's ability to attract and gain electrons.

Across each period electronegativity increases. This is because electrons are more attracted to the nucleus, as it contains more protons. Whereas, down each group, electronegativity decreases. This is caused by the increased number of electron shells, which shield electrons from the attraction to the nucleus.

These trends are shown in the graph below. As you can see, electronegativity increases across period 2 from lithium ($Z = 3$) to fluorine ($Z = 9$). Down group 17, you can see that electronegativity decreases from fluorine ($Z = 9$) to chlorine ($Z = 17$) to bromine ($Z = 35$). It is important to understand that the noble gases generally have no electronegativity, because they have a stable electron arrangement and don't gain electrons.

Electronegativity vs. atomic number



I've summarised the trends for electronegativity in the table below:

Electronegativity	
Across each period (left to right)	Increases
Down each group	Decreases

Note that the formula and data booklet gives data relating to atomic radii, ionic radii, first ionisation energies, and electronegativities, so you may need to look at this information for data related questions.

1.2.7 Acids and bases

We now need to look at how the acidic nature of the oxides of elements change on the periodic table. Oxides are compounds that are made up of oxygen and another element. For example, magnesium oxide is made up of magnesium and oxygen atoms.

The oxides of the elements on the left-hand side of the table are basic. This is implied by the name of group 1 and group 2, as they are called alkali metals and alkaline earth metals. An alkaline is a base dissolved in water, so the name states that the oxides of these elements are basic. Whereas, the oxides of the elements on the right-hand side of the periodic table are acidic.

Therefore, as you move across each period, the oxides of the elements go from basic to amphoteric to acidic. Amphoteric means that a substance can react as an acid and as a base – don't worry, I will talk about acids and bases in more detail later on page 61.

1.3 Introduction to bonding

When different elements bond together, they form a compound. There are three types of chemical bonds: **ionic, covalent, and metallic bonds**. The elements' electron configurations determine which bonds they can form, as bonds are caused by electrostatic attraction.

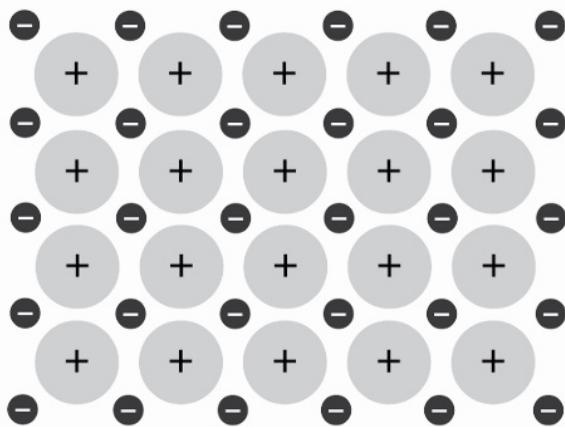
1.3.1 Valency

Valency is the number of bonds which an atom can form. This is determined by the number of valence electrons. This is because as stated by the octet rule, atoms want to obtain a stable arrangement of valence electrons (eight electrons in the outer shell).

- Group 1: form cations with a charge of **+1** (as they lose one electron), so they can form **one bond**.
- Group 2: form cations with a charge of **+2** (as they lose two electrons), so they can form **two bonds**.
- Groups 3–12 (transition metals): form **ions of various charges**. Therefore, they can form **different numbers of bonds**.
- Group 13: form cations with a charge of **+3** (as they lose three electrons). Therefore, they can form **three bonds**.
- Group 14: form **+4** and/or **-4** ions depending on whether the element is a metal, non-metal or metalloid. However, all elements in group 14 can form **four bonds**.
- Group 15: form anions with a charge of **-3** (as they gain three electrons), so they can form **three bonds**.
- Group 16: form anions with a charge of **-2** (as they gain two electrons), so they can form **two bonds**.
- Group 17: form anions with a charge of **-1** (as they gain one electron), so they can form **one bond**.
- Group 18: **don't form ions nor bonds** because they are stable.

1.3.2 Metallic bonding

Metallic bonds form between metal atoms. As you would already know, metal atoms' valence electrons aren't very attracted to the nucleus. This means that the outer shell electrons are able to detach from their atoms and move freely about the metal. This creates a 'sea' of delocalised electrons, which surrounds a lattice of positive metal cations. Note that the metal atoms' inner shell electrons are not capable of detaching from the atom because they are more strongly attracted to the nucleus. This diagram shows the structure of metallic bonding.

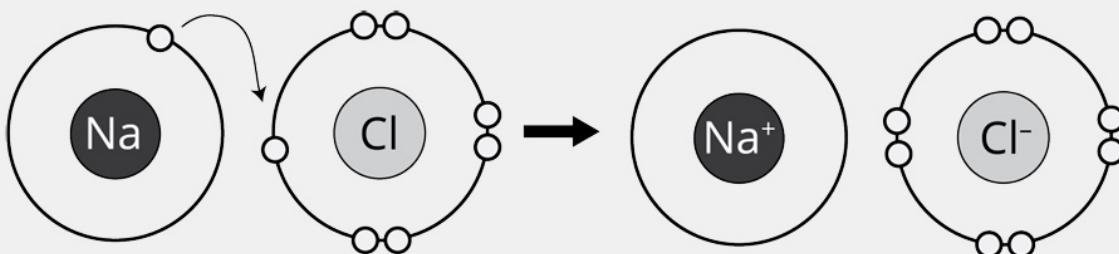


1.3.3 Ionic bonding

Ionic bonds form between a non-metal and a metal. You already know that metals tend to lose electrons to obtain eight valence electrons, whereas non-metals tend to gain electrons. In an ionic bond, electrons are transferred between atoms, as the non-metal atoms gain electrons (form anions) and the metal atoms lose electrons (form cations). When this occurs, the metal cations are electrostatically attracted to the non-metal anions.

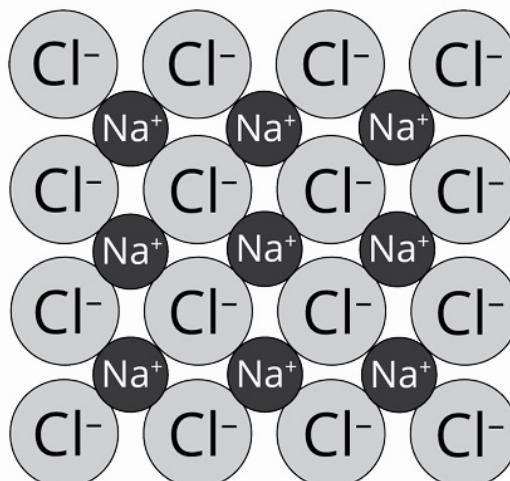
SAMPLE :

Let's take a look at the example of sodium chloride (NaCl), which has ionic bonding. Sodium chloride is made up of sodium and chlorine atoms which form charged ions. As shown on the periodic table, sodium (Na) is an alkali metal so it forms Na^+ (a cation with a charge of +1) by losing one electron. Chlorine (Cl) is a halogen so it forms Cl^- (an anion with a charge of -1) by gaining one electron. Therefore, sodium chloride is formed as electrons are transferred from sodium atoms to chlorine atoms. This forms Na^+ and Cl^- which are electrostatically attracted to each other. This is shown by the diagram below (which displays the atoms' outer shells).



When cations and anions combine, they normally form a three-dimensional lattice. For example, in NaCl , the Na^+ ions and the Cl^- ions combine to form lattice structure (as shown in the diagram on the right).

These lattices are quite strong because the electrostatic attraction between ions of opposite charge outweighs the repulsion between the ions of like charges.



Now let's have a look at how you can write the chemical formulas for ionic compounds. You may have covered this in year 10 but it's good to review this process.

SAMPLE :

Let's find the chemical formula for zinc (II) nitride:

1. First, we need to write the chemical symbols for the ions. Zinc is a transition metal, so it can form various ions. But these zinc atoms have a valency of two, as shown by the name. So, zinc forms Zn^{2+} cations.
2. The non-metal in this compound is nitrogen (N). This forms N^{3-} anions.
3. Now we need to determine how many ions are needed of each to balance the charges. This is because the compound is neutral. To do this, we need to calculate the lowest common multiple of 2 and 3, which is 6. So, three Zn^{2+} cations balance the charge of two N^{3-} anions.
4. The number of ions is represented by the subscript. So, the compound is Zn_3N_2 .

Now that you know how to write the chemical formulas, let's look at how you can name ionic compounds.

SAMPLE :

Let's discuss how you name Zn_3N_2 :

1. First, look at which types of atoms make up the compound. In Zn_3N_2 , we can see that the compound consists of zinc (Zn) and nitrogen (N) atoms.
2. When naming an ionic compound, first you state the name of the metal (which forms cations). Since zinc is a transition metal, make sure you state its valency in brackets. So, for this compound, zinc (II) is the metal.
3. Next you state the name of the non-metal (which forms anions). For this compound, nitrogen is the non-metal. In ionic compounds, '-ide' is added to the end of the non-metal's name. So 'nitrogen' becomes 'nitride'.
4. Then you simply combine the names for the metal and the non-metal. So, this compound is called zinc (II) nitride.

It's important that we also consider when ions contain atoms of two or more elements. These are called **polyatomic ions**. So far, we've only considered **monatomic ions** (which are made up of only one type of element). The names and formulas for common polyatomic ions are given at the back of the QCAA formula book, so you don't have to memorise them. For example, K_2CrO_4 is called potassium chromate, as the CrO_4^{2-} ion is called chromate. You may have noticed that the suffix '-ate' is associated with ions containing oxygen.

1.3.4 Covalent bonding

Covalent bonds form between non-metal atoms. When this occurs, the electrons are shared between the atoms. There are two types of covalent bonding: **non-polar** and **polar**. In non-polar bonds, electrons are shared equally between the atoms. Whereas, in polar bonds, electrons are shared unequally between the atoms, as they are more attracted to one atom.

A polar bond occurs when there is a large difference in electronegativity between the atoms. For example, a polar bond is formed between hydrogen and fluorine because fluorine has a much larger electronegativity value (4.0) than hydrogen (2.2). A non-polar bond is formed between carbon and hydrogen because the electronegativity of carbon (2.6) is very close to the electronegativity of hydrogen (2.2).

Note that in comparison to metallic and ionic compounds, covalent compounds can be made up of molecules or a lattice. For example, water (H_2O) is a molecule, whereas diamond (which is a network of carbon atoms covalently bonded together) is a lattice. These lattices are called covalent networks.

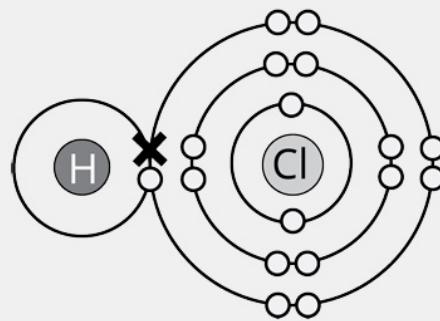
Differences in electronegativity between the atoms in a compound determine the type of bonding. This is shown in the table below:

Type of bond	Difference in electronegativity
Non-polar covalent	<0.5
Polar covalent	0.5-1.7
Ionic	>1.7

In covalent bonds, electrons are shared between atoms so they can obtain a stable arrangement of valence electrons. Let's take a look at an example.

SAMPLE :

Let's discuss hydrogen chloride (HCl). As the name suggests, it is made up of hydrogen (H) and chlorine (Cl) atoms (which are both non-metals). We can work out using the periodic table that hydrogen has one valence electron and chlorine has seven valence electrons. Therefore, to obtain a stable arrangement of electrons, a pair of electrons are shared between the hydrogen and chlorine atoms. This results in the atoms having full outer shells, as chlorine has eight valence electrons and hydrogen has two valence electrons.



Covalent compounds are named in a very similar way to ionic compounds. The non-metal furthest to the left of the periodic table is stated first. The last element in the name also has the 'ide' prefix. However, when naming covalent compounds prefixes are included. These indicate the number of each atom:

Number of atoms	Prefix
1	Mono-
2	Di-
3	Tri-
4	Tetra-
5	Penta-
6	Hexa-
7	Hepta-
8	Octa-
9	Nona-
10	Deca-

SAMPLE :

Let's take a look at some examples of how to name covalent compounds:

- CO = carbon monoxide (note that the prefix 'mono' isn't used when naming the first element)
- N₂O = dinitrogen oxide
- SF₆ = sulphur hexafluoride
- CF₄ = carbon tetrafluoride

We now need to look at how to write the chemical formulas for covalent compounds:

SAMPLE :

Let's write the chemical formula for nitrogen trifluoride:

By reading the prefixes, we know that there is one nitrogen atom and three fluorine atoms in this compound. So, the chemical formula is NF₃.

It is also important to understand that some elements exist as **diatomic molecules** (which contain two atoms). For example, hydrogen exists as two hydrogen atoms bonded together (H₂).

This also occurs for:

- Nitrogen (N₂)
- Fluorine (F₂)
- Oxygen (O₂)
- Iodine (I₂)
- Chlorine (Cl₂)
- Bromine (Br₂)

All of the halogens exist as diatomic molecules. The other type of molecules are polyatomic molecules, which are made up of more than two atoms. For example, water (H₂O) is a polyatomic molecule, as it contains three atoms.

Another example of a covalent compound is **ammonia (NH₃)**. It is important to recognise the name of this compound because it is commonly used in questions.

1.3.5 Lewis dot diagrams

Lewis dot diagrams (which are also known as electron dot diagrams) are used to represent covalent bonding. These diagrams help you distinguish between bonding and non-bonding electrons. Pairs of non-bonding electrons are known as **lone pairs**.

There are different types of covalent bonds. A **single covalent bond** occurs when a pair of electrons is shared, whereas a **double covalent bond** occurs when two pairs of electrons are shared. Following the pattern, a **triple covalent bond** occurs when three pairs of electrons are shared.

In Lewis dot diagrams, symbols are used for elements, lines are used to show bonds (which are made up of pairs of bonding electrons), and dots show non-bonding electrons. Lewis dot diagrams only show valence electrons. Here are the steps to drawing a Lewis dot diagram:

1. Count the total number of valence electrons.
2. Assemble bonding using only single bonds. Make sure that the central atom can form the most bonds or is the least electronegative.
3. Place non-bonding electrons on each outer atom.
4. Assign the remaining valence electrons to inner atoms or complete octets.

Note that carbon, nitrogen, oxygen, and fluorine always obey the octet rule.

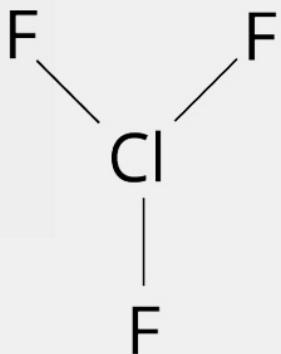
Now I'm going to show you how to draw the Lewis dot diagram for chlorine trifluoride (ClF_3):

SAMPLE :

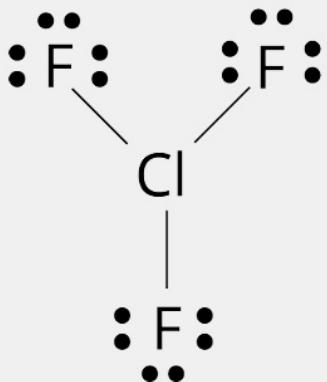
1. Total number of valence electrons:

$$\begin{aligned}
 &= \text{chlorine valence electrons} + 3 \times \text{fluorine valence electrons} \\
 &= 7 + (3 \times 7) \\
 &= 28
 \end{aligned}$$

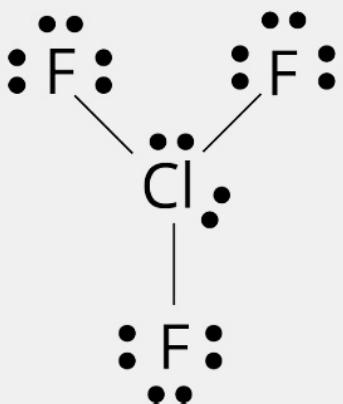
2. Chlorine is the central atom as it is less electronegative than fluorine.



3. The fluorine atoms have eight electrons surrounding them, so they are following the octet rule.



4. There are four electrons remaining. Chlorine has only six electrons surrounding it, so the extra electrons will be added to the central atom.



So, from our final Lewis dot diagram, we can see that there are three pairs of bonding electrons and eleven lone pairs (i.e. non-bonding electrons).

1.4 Analytical techniques

In this section, we'll look at some techniques chemists use to determine what a substance is made of.

1.4.1 Mass spectrometry

Mass spectrometry can be used to determine which isotopes are present in a substance. It can also be used to measure the abundance of isotopes in the sample. We've already discussed isotopes on page 6, but as a refresher, remember that the isotopes of an element have different numbers of neutrons.

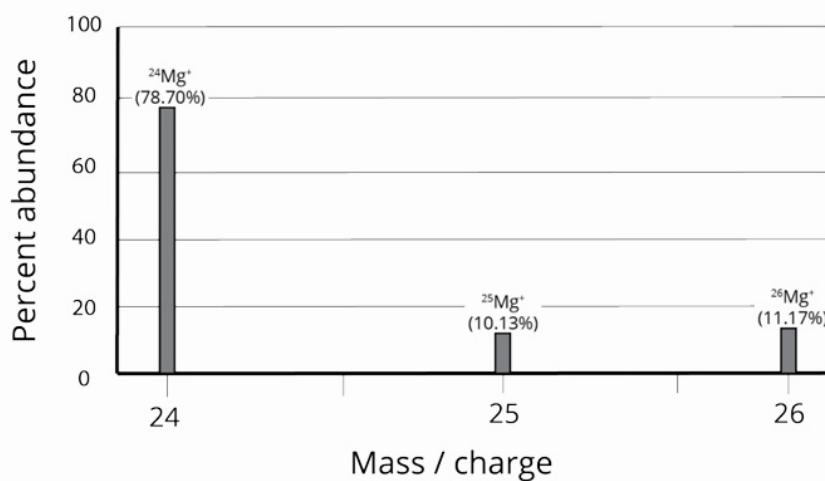
First, before we discuss the process of mass spectrometry, you need to understand relative isotopic mass. **Relative isotopic mass** is an isotope's mass relative to the mass of a carbon-12 atom. So, the relative isotopic mass of a carbon-12 atom (which has six protons and six neutrons) is 12. Therefore, the relative isotopic mass is equal to the sum of the number of protons and neutrons. For example, the relative isotopic mass of a hydrogen atom (which has one proton and no neutrons) is equal to 1.

There are five main steps required to complete mass spectrometry:

1. The sample is vaporised and then the gas is ionised to form positively charged ions.
2. These ions are accelerated by an electric field.
3. The ions are then separated by a magnetic field based on their mass-to-charge ratio.
4. Ions of a specific mass and charge reach a detector, measuring the current produced by the ions.
5. This output is used to create a **mass spectrum** (a graph showing percentage abundance vs mass-to-charge ratio).

For example, see the mass spectrum of magnesium below:

Magnesium mass spectrum



Using the mass spectrum, we can see that there are three isotopes (as this is indicated by the number of peaks). Since the charge of the ions was +1, the relative mass of each isotope is given by the mass-to-charge ratio. The percentage abundance indicates how much of each isotope was present in the sample. So, we can see that the percentage abundance of magnesium-24 is 78.70%, magnesium-25 is 10.13%, and magnesium-26 is 11.17%.

Using this information, we can calculate the **relative atomic mass** (A_r) of magnesium. This is the weighted average of the relative isotopic masses.

For an element with three isotopes, we can use the formula below:

$$A_r = \frac{\text{relative isotopic mass} \times \% \text{ isotopic abundance}}{100} + \frac{\text{relative isotopic mass} \times \% \text{ isotopic abundance}}{100} + \frac{\text{relative isotopic mass} \times \% \text{ isotopic abundance}}{100}$$

SAMPLE :

We can use this formula to calculate the relative atomic mass for magnesium:

$$A_r = \frac{24 \times 78.70\%}{100} + \frac{25 \times 10.13\%}{100} + \frac{26 \times 11.17\%}{100}$$

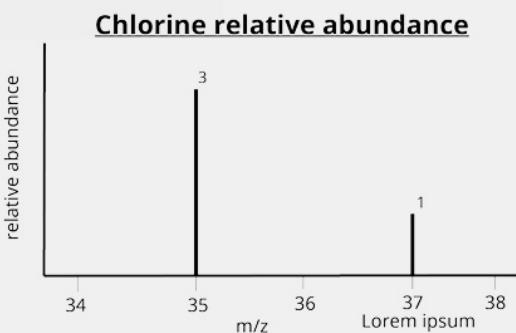
$$A_r = 24.32 \text{ (rounded to two decimal places)}$$

The relative atomic mass values for all elements are shown on the periodic table.

In an exam, instead of percentage abundance, the *y*-axis of the mass spectrum may be **relative abundance**. This can be used to calculate percentage abundance.

SAMPLE :

Let's look at the mass spectrum for chlorine. Using the mass spectrum, we can see that the relative abundance is 3 for chlorine-35 and 1 for chlorine-37. In an exam, these numbers may not be given to you, so you will need to use a ruler to measure the lengths of the lines.



Now we can calculate percentage abundance for each isotope:

$$\text{Total peak height} = 3 + 1 = 4$$

$$\% \text{ abundance} = \frac{\text{peak height}}{\text{total peak height}} \times 100\%$$

Chlorine-35:

$$\% \text{ abundance} = \frac{3}{4} \times 100\% = 75\%$$

Chlorine-37:

$$\% \text{ abundance} = \frac{1}{4} \times 100\% = 25\%$$

Now, we can use the percentage abundances to calculate the relative isotopic mass for chlorine.

A harder exam equation may give you the relative atomic mass of an element and ask you to calculate the percentage abundances of each isotope. Below is an example.

SAMPLE :

Boron's relative atomic mass is 10.811. The isotopic masses of boron's two isotopes are 10.013 and 11.009. Calculate the percentage abundances for each isotope:

Let the % abundance of 10.013 isotope = x .

Let the % abundance of 11.009 isotope = $100 - x$ (this is because the sum of the percentage abundances will equal 100%).

Substitute these values in the formula for relative atomic mass:

$$\begin{aligned} 10.811 &= \frac{10.013 \times x}{100} + \frac{11.009 \times (100 - x)}{100} \\ 10.811 &= \frac{10.013x + 1100.9 - 11.009x}{100} \\ 1081.1 &= -0.996x + 1100.9 \\ -19.8 &= -0.996x \\ x &= \frac{-19.8}{-0.996} \\ x &= 19.88\% \end{aligned}$$

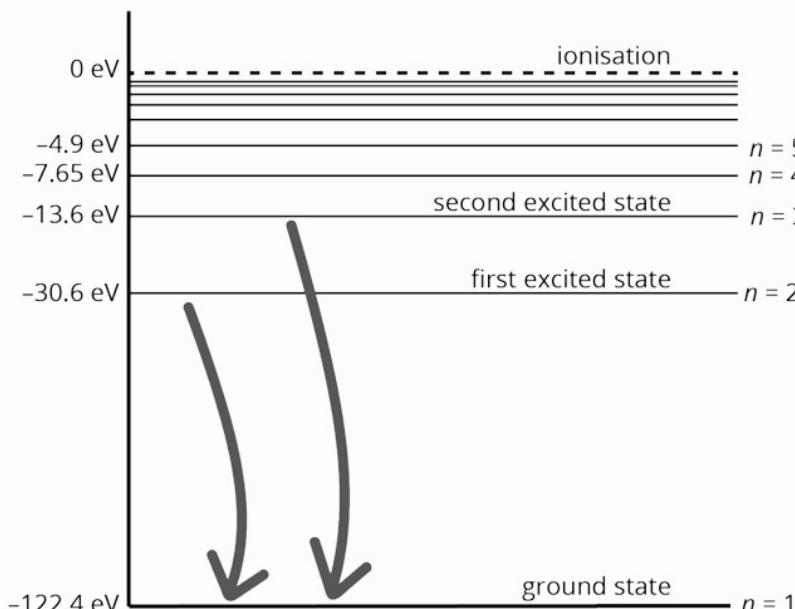
Therefore, the percentage abundance of the 10.013 isotope is 19.88%.

$$100 - x = 80.12\%$$

So, the percentage abundance of the 11.009 isotope is 80.12%.

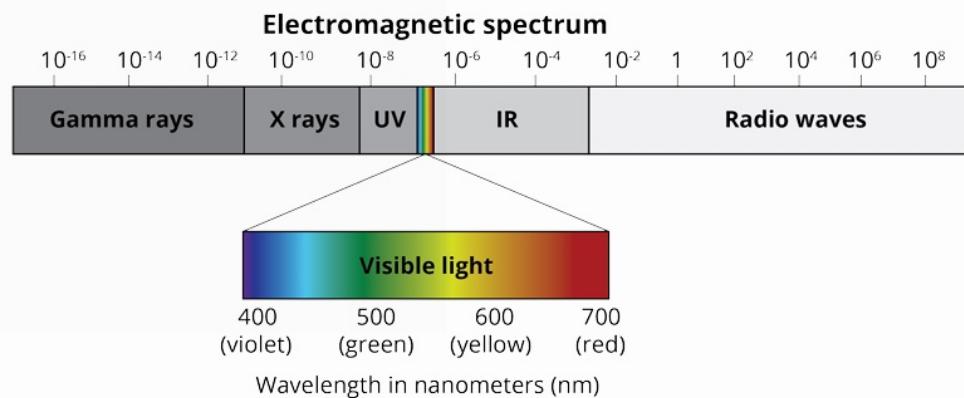
1.4.2 Flame tests

In a flame test, a substance is passed through a flame to determine which ions are present. You already know that when an electron absorbs energy it becomes 'excited' and goes into a higher energy level. This occurs when an electron absorbs thermal energy from the flame. Energy is released when this electron returns to ground state. This energy is released in the form of light, which is a type of electromagnetic radiation. A photon is emitted (which is a particle of light) that has a specific amount of energy. The amount of energy released is equal to the difference in energy levels between the excited and ground levels. This is shown in the energy level diagram below.



For example, in the previous diagram, you can see that more energy would be released when an electron returns from the second excited state to the ground state than from the first excited state. This is because there is a larger energy difference between the third ($n = 3$) and first ($n = 1$) energy levels than the second ($n = 2$) and first ($n = 1$) energy levels.

You may already know about the electromagnetic spectrum. So, you would know that only visible light (which has wavelengths between 400 nm and 700 nm) can be seen by the human eye. Light of different energies have different wavelengths, and this corresponds to seeing different colours.

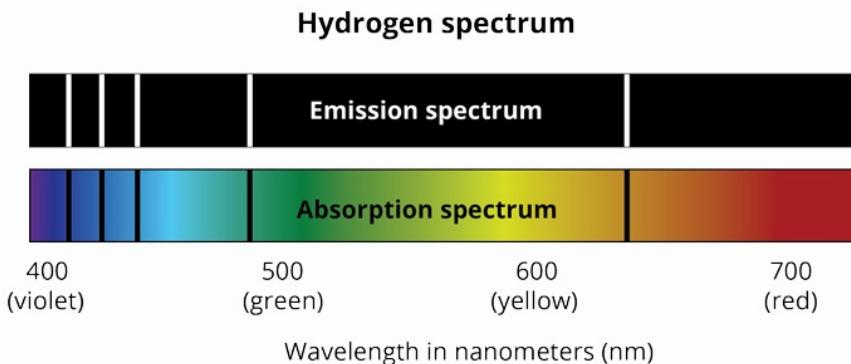


Therefore, due to their different electron configurations, different ions will produce different colours, as electrons are excited to different energy states.

See the table below for some of the colours produced by certain metals:

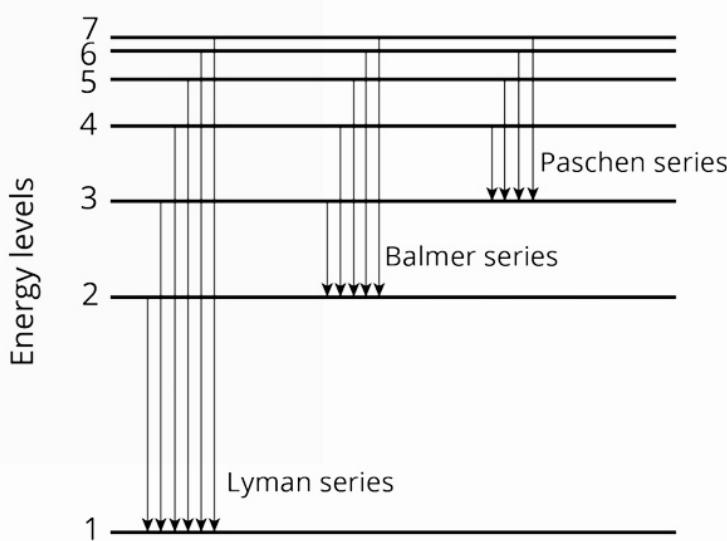
Metal	Flame colour
Lithium	Red
Strontium	Scarlet
Barium	Yellow-green
Sodium	Yellow
Copper	Green
Potassium	Lilac
Calcium	Yellow-red

We can also look at absorption spectra. These show all the different wavelengths of light which are absorbed by an element when an electron goes into different excited states. Similarly, emission spectra show all the different wavelengths of light which are emitted by an element when an electron releases energy and returns to a lower energy level.



As you can see in the diagram above, the lines on the emission spectrum show the emitted wavelengths of light. The black lines on the absorption spectrum show the absorbed wavelengths of light.

The Lyman series is the series of lines observed which correspond to an electron in a hydrogen atom returning to the first energy level ($n = 1$). The Balmer series is for an electron in a hydrogen atom returning to the second energy level ($n = 2$), and the Paschen series is for an electron returning to the third energy level ($n = 3$).



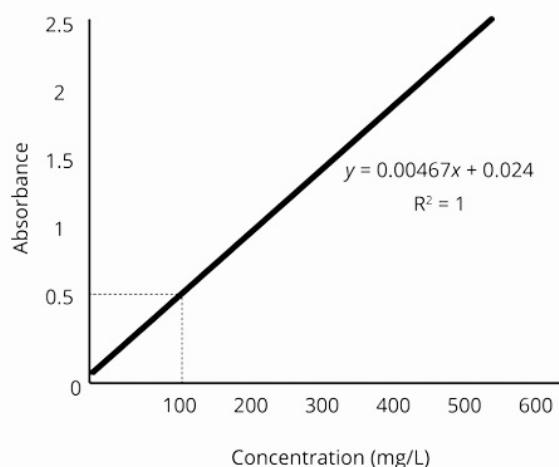
1.4.3 Atomic absorption spectroscopy (AAS)

Atomic absorption spectroscopy (AAS) is used to determine what metals are present in a sample, and what their concentrations are. This is determined by measuring the amount of light absorbed by the sample to excite the electrons.

The following steps are required to conduct AAS:

1. Different wavelengths of light are emitted by a hollow source cathode lamp. This light is absorbed by the metal sample.
2. The sample is sprayed into the flame, so it is vaporised.
3. A monochromator is used to select a wavelength for analysis.
4. The amount of light that reaches a detector is recorded, and this information is used to calculate the amount of light that was absorbed by the sample.
5. The absorbance of light is measured, and this can be used to find the concentration of the metal. To do this a calibration curve is needed.

For example, a calibration curve for copper (II) is shown below:



A calibration curve is created by conducting AAS on many solutions of metals and measuring their concentration and absorbance. For example, let's say that we conduct AAS and find that the absorbance of a copper (II) sample is 0.5. Using the calibration curve above, we can conclude that the concentration is 100 mg/L.

Section 2

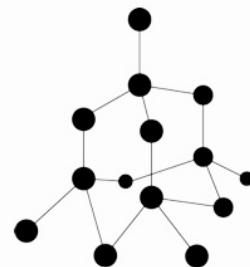
Properties and structure of materials

2.1 Compounds and mixtures

2.1.1 Pure substances vs mixtures

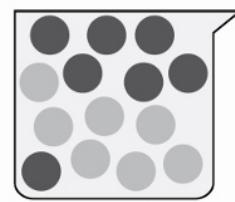
Elements and compounds are **pure substances**. Pure substances, which contain only one element or compound, have definite and constant composition. These substances have distinct measurable properties (e.g. density, melting and boiling points).

For example, diamond (which is a lattice of carbon atoms) is a pure substance which is displayed in the diagram.

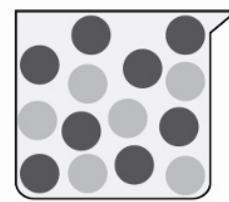


When pure substances physically combine (without chemical bonding), this produces a **mixture**. For example, sand and water is a mixture. In contrast to pure substances, mixtures don't have distinct characteristics, as these depend on the relative amounts of the substances in the mixture. For example, mixtures don't have distinct density nor melting and boiling points.

There are two types of mixtures: **homogenous and heterogenous mixtures**. Homogenous mixtures have a uniform composition (e.g. air) whereas heterogenous mixtures have a non-uniform composition. They are separated into phases, so the different substances in the mixtures are easily visually distinguished (e.g. a salad).

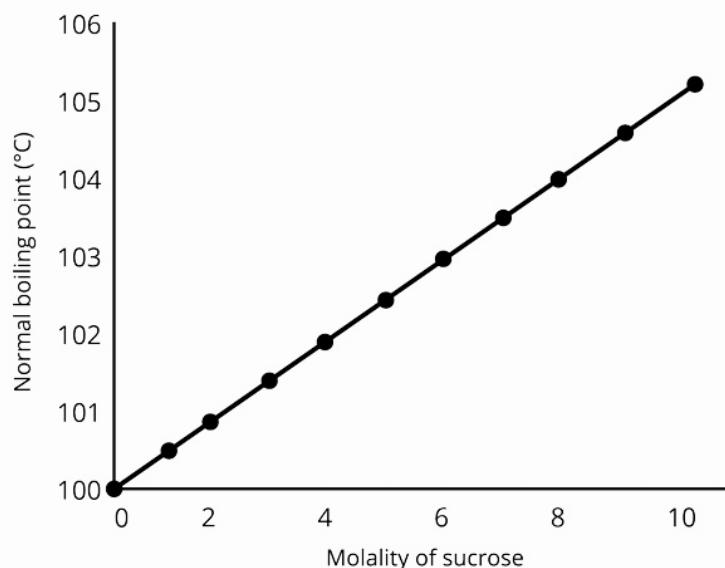


Heterogenous mixture



Homogenous mixture

Let's look at the graph below. This graph is based on a mixture of water and sucrose. The graph shows the relationship between the boiling point and the molarity of sucrose. Don't worry we will learn about molarity later in this book, but it basically means the amount of sucrose in the mixture. We can see from the graph that the boiling point linearly increases as the amount of sucrose in the mixture increases. So, this shows that the properties of mixtures are dependent on the amounts of each substance in it.



2.1.2 Nanomaterials

You may have heard of **nanoscience**, which is the study of materials on a very small scale (this can be called the **nanoscale**). 1 **nanometre** (which has the symbol nm) is equal to one billionth of a metre ($1 \text{ nm} = 1 \times 10^{-9} \text{ m}$).

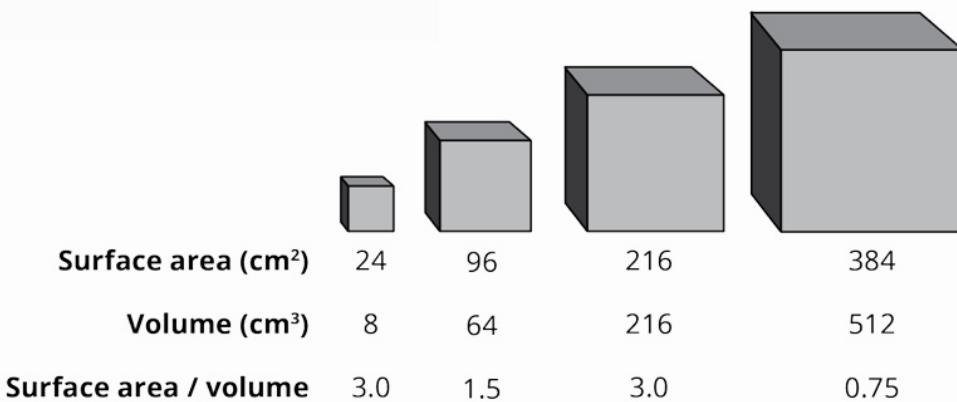
For example, let's convert 2 mm into nanometres:

1. First, we will convert 2 mm into m, which is 0.002 m.
2. Then, we will convert 0.002 m into nanometres by multiplying it by 10^9 . So, 0.002 m is 2×10^6 nm.

Nanoparticles are used for a range of purposes and have different properties based on the size of these particles. One of the most important properties of nanoparticles is their large surface area compared to their volume. This can be described as a **large surface area to volume ratio**.

KEY POINT :

The smaller the particles are, the greater the surface area to volume ratio.



Therefore, having lots of small particles rather than a few large particles will increase the surface area to volume ratio.

A large surface area to volume ratio is important for three reasons:

- It allows nanoparticles to **absorb** many molecules. So, nanoparticles such as charcoal can remove unwanted gases. This is why charcoal was used to create masks in World War I to protect soldiers from dangerous gases.
- Nanoparticles can **transport** other molecules which are absorbed into them. Therefore, they are used in medicine to deliver drugs through the body.
- It allows nanoparticles to act as **catalysts** (as they speed up reactions), due to their large surface area. Reactant molecules are absorbed into the nanoparticles so they can combine to form the product.

Carbon nanomaterials are made up of carbon to carbon bonds. In particular, carbon nanotubes, which are made up of nanoparticles, are good electrical and thermal conductors, strong, and good absorbers. This allows carbon nanotubes to be used for a range of purposes, including drug delivery.

2.2 Bonding and properties

We've already discussed the three types of bonds: covalent, ionic and metallic. Now, we need to consider the properties of these compounds.

The table below includes the different properties of covalent, ionic, and metallic compounds.

	Melting and boiling point	Tensile strength (resistance to breaking when a force is applied)	Thermal conductivity	Electrical conductivity
Ionic compounds	High	Low (brittle)	High only in liquid or aqueous state	High only in liquid or aqueous state
Covalent compounds	Low	Low	No	No
Metals	High (higher for transition metals)	High	High	High
Explanation	<p>A large amount of energy is required to overcome the strong electrostatic attraction between cations and anions in ionic compounds.</p> <p>The attraction between covalent molecules/compounds is relatively weak, so it is easier to overcome this attraction.</p> <p>A large amount of energy is required to overcome the strong electrostatic attraction between the metal cations and the delocalised electrons in metals.</p>	<p>Ionic compounds are brittle because when a force is applied this causes like charges to be shifted next to each other. Like charges repel, so the compound breaks easily.</p>	<p>When ionic compounds are liquids or in aqueous solutions, the ions are free to move. This is because when ionic compounds dissolve in water, the ionic bonds break, and the ions can move freely. Therefore, these ions can move throughout and transfer electrical and thermal energy.</p> <p style="text-align: center;">Electrical conductivity of ionic compounds</p>	<p>In covalent compounds, there are no free charges. When covalent compounds dissolve in water, they do not divide into ions and instead split into molecules. In metals, free electrons move throughout and when they collide with cations thermal energy is transferred. These electrons are also able to move from negative to positive terminals in a circuit to transfer electrical energy.</p>

Metals have other special properties:

- **Malleable** (can be shaped by beating or rolling) and **ductile** (can be drawn into a wire): this is because when a force causes the metal cations to move, the cation layers are still held together by the delocalised electrons moving between them.
 - **Lustrous and reflective**: because free electrons move throughout the metal and reflect light.

Now let's consider some other types of substances and their properties.

Alloys are created when other metals or carbon are added to a melted metal, forming a mixture. As we discussed before, mixtures have different properties depending on the amounts of each substance in the mixture.

Allotropes are made up of the same element but have different structures. This is because the atoms are bonded together differently. Therefore, allotropes have different chemical and physical properties. For example, graphite and diamond are allotropes of carbon, and yet they look completely different!



Graphite, a carbon allotrope

Source: Robert Lavinsky, iRocks.com – CC-BY-SA-3.0

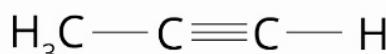


Diamond, also a carbon allotrope

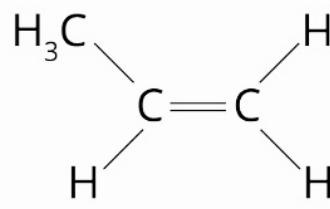
Source: Catherine Lewis, estatediamondjewelry.com – CC-BY-SA-4.0

It is also important to recognise that having multiple bonds (i.e. double or triple bonds) affects the properties of a substance. For example, a triple carbon to carbon bond has a shorter bond length than double and single carbon to carbon bonds. So, propyne will have different properties to propene and propane, even though they contain similar atoms. For example, propyne will have the highest boiling point because a triple bond requires more energy to break than a double or single bond.

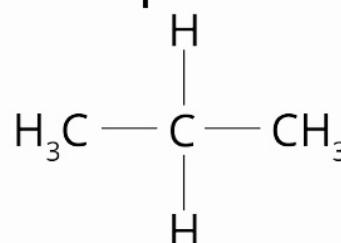
Propyne



Propene



Propane



These compounds are called hydrocarbons, as they only contain hydrogen and oxygen. You will learn more about these substances in Year 12.

Section 3

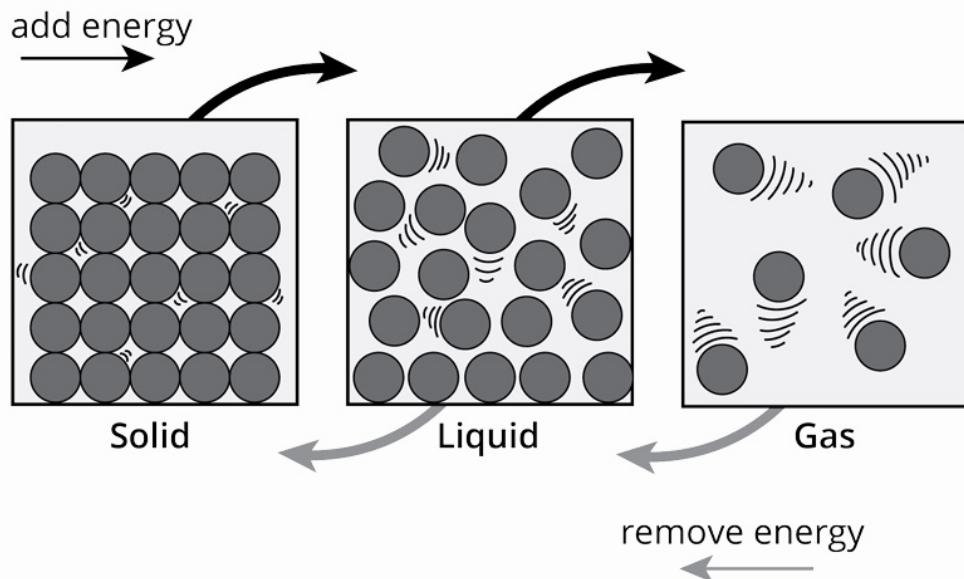
Chemical reactions – reactants, products, and energy change

3.1 Chemical reactions

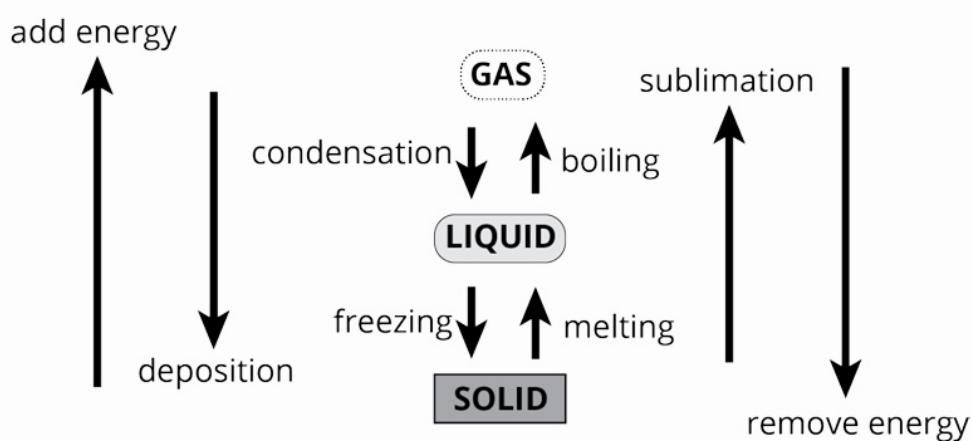
3.1.1 Chemical reaction vs. physical change

A **chemical reaction** is when one or more reactants change into one or more products, as the reactants' atomic structures are rearranged. Whereas, a **physical change** occurs when matter is changed into a different form and its chemical composition remains the same. So, frying an egg is a good example of a chemical reaction, as chemical bonds are broken and formed to produce new particles (and we can't un-fry an egg!). In comparison, when ice melts, this is a physical change because the water molecules can move more freely, as energy is added in the form of heat, but this water can be re-frozen.

When a substance changes state, this is called a **phase change**.



The names of these phase changes are described in the diagram below:

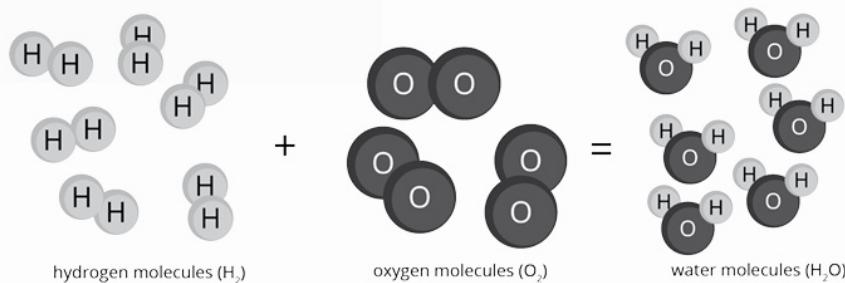


3.1.2 Chemical equations

Now that we've talked about physical changes, let's look at chemical reactions. Chemical equations are a visual representation of chemical reactions. They include reactants on the left-hand side, and products on the right-hand side. The reactants and the products are separated by an arrow (\longrightarrow) to show the chemical change. Coefficients are used to show the correct proportions of reactants and products in the reaction. When you write chemical equations, you need to include symbols to indicate what state the reactants and products are in:

- **s** = solid
 - **l** = liquid
 - **g** = gas
 - **aq** = aqueous solution

For example, $2\text{H}_2\text{(g)} + \text{O}_2\text{(g)} \rightarrow 2\text{H}_2\text{O(g)}$ shows that two molecules of hydrogen gas will react with one molecule of oxygen to form two molecules of water vapour. This reaction is shown visually below:



KEY POINT :

The reaction in the equation above is balanced because the number of hydrogen atoms (4) and oxygen atoms (2) in the reactants is equal to the number of hydrogen atoms (4) and oxygen atoms (2) in the products. This follows the **law of conservation of mass**, which states that the total mass of the reactants is equal to the total mass of the products.

There are six main types of chemical reactions:

- **Single displacement:** $A + BX \rightarrow B + AX$
 - In this reaction, one element is substituted for another element in a compound.
 - For example, $Zn + CuCl_2 \rightarrow ZnCl_2 + Cu$
 - **Double displacement:** $AY + BX \rightarrow BY + AX$
 - In this reaction, the cations and anions of two ionic compounds switch places.
 - For example, $AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$
 - **Acid-base:** acid + base \rightarrow salt + water
 - An acid-base reaction is a special type of double displacement reaction. A salt is an ionic compound.
 - For example, hydrochloric acid (HCl) and sodium hydroxide (base) react: $HCl + NaOH \rightarrow NaCl + H_2O$
 - **Combustion:** hydrocarbon + oxygen \rightarrow carbon dioxide + water + heat
 - This reaction occurs when a fuel reacts with oxygen and gives off heat. Combustion is just a scientific name for burning.
 - For example, methane (CH_4) reacts with oxygen: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$
 - **Combination:** $A + B \rightarrow AB$
 - This occurs when two or more reactants combine to form one product.
 - For example, $CaO + H_2O \rightarrow Ca(OH)_2$
 - **Decomposition:** $AB \rightarrow A + B$
 - This occurs when one reactant breaks down to produce multiple products.
 - For example, $NaCl \rightarrow Na + Cl$

Now let's look at how to balance chemical equations.

SAMPLE :

I'm going to use the example of hydrogen gas and oxygen gas forming water vapour.

Write out the chemical symbols and states for the reactants and products:



Write out the number of atoms of each element in the reactants and products:

	Reactants	Products
Number of hydrogen atoms	2	2
Number of oxygen atoms	2	1

You can now see that the chemical equation above isn't following the law of conservation of mass. This is because the number of oxygen atoms in the reactants and products aren't equal.

Now we can add coefficients to the chemical equation to balance it. We can see that there needs to be two molecules of water, so there are two oxygen atoms in the products. This will create four hydrogen atoms in the products: $\text{H}_2\text{(g)} + \text{O}_2\text{(g)} \rightarrow 2\text{H}_2\text{O(g)}$.

	Reactants	Products
Number of hydrogen atoms	2	4
Number of oxygen atoms	2	2

Therefore, we will need two molecules of hydrogen gas to balance the four hydrogen atoms in the products: $2\text{H}_2\text{(g)} + \text{O}_2\text{(g)} \rightarrow 2\text{H}_2\text{O(g)}$.

	Reactants	Products
Number of hydrogen atoms	4	4
Number of oxygen atoms	2	2

Balancing chemical equations can be tricky, so make sure you do lots of practice questions!

Both chemical reactions and phase changes involve changes in energy. For example, when a liquid is boiled, this process requires energy in the form of heat. Another example is the chemical reaction which occurs in a heat pack, as this produces energy in the form of heat. Energy emission and absorption in chemical reactions and phase changes can also be observable by the emission or absorption of light.

3.2 Exothermic and endothermic reactions

3.2.1 What is heat?

Now let's look at how reactions release or absorb energy in the form of heat. But first you need to understand what energy is. **Energy** is the capacity to do work measured in joules (J). Energy comes in many forms, including kinetic, chemical, and electrical energy.

The **law of conservation of energy** states that during any chemical or physical process, energy is neither created nor destroyed. So, energy is exchanged between the system and the surroundings. In this case, the system is the chemical reaction and the surroundings is everything else.

Heat is energy that flows between objects due to a temperature difference. If heat is added to an object, its temperature increases. As you would already know, heat spontaneously flows from hotter to cooler objects until they reach the same temperature. For example, if you have a cup of coffee, heat is transferred from the hot coffee to the surroundings, and therefore the coffee cools down until it is the same temperature as the surroundings.

Temperature is a measure of the average kinetic energy of the particles which make up an object. If you measure your temperature when you are sick, you will get a temperature reading in degrees Celsius ($^{\circ}\text{C}$). However, in science, temperature is usually recorded in kelvin (K). This is because 0 K (which equals $-273\ ^{\circ}\text{C}$) is the temperature at which particles have no energy. This temperature is called absolute zero. The formula for converting between temperature in kelvin and degrees Celsius is shown below:

$$\text{Temperature in K} = \text{temperature in } ^{\circ}\text{C} + 273$$

As you would know, as temperature increases, an ice cube would melt. The ice cube is a solid, so it has a definite shape as the particles are packed closely together. These particles vibrate and cannot move freely. However, as temperature increases, these particles gain kinetic energy, so they can move more freely. Therefore, the ice cube becomes liquid water, which doesn't have a definite shape. This is described by **kinetic theory**, which states that as temperature increases, the movement of particles increases.

3.2.2 Endothermic and exothermic reactions

Chemical energy is stored in the bonds of atoms and molecules in a substance. In a chemical reaction, the reactants and products have different amounts of chemical energy stored in their bonds. Therefore, this difference causes energy to be absorbed or released as heat.

In an **exothermic reaction**, the reactants have more chemical energy than the products, so energy is released. If an exothermic reaction occurs in a container, then it will heat up the container. Combustion (burning) reactions are exothermic. For example, if methane is burned, then it will produce water, carbon dioxide gas and heat.

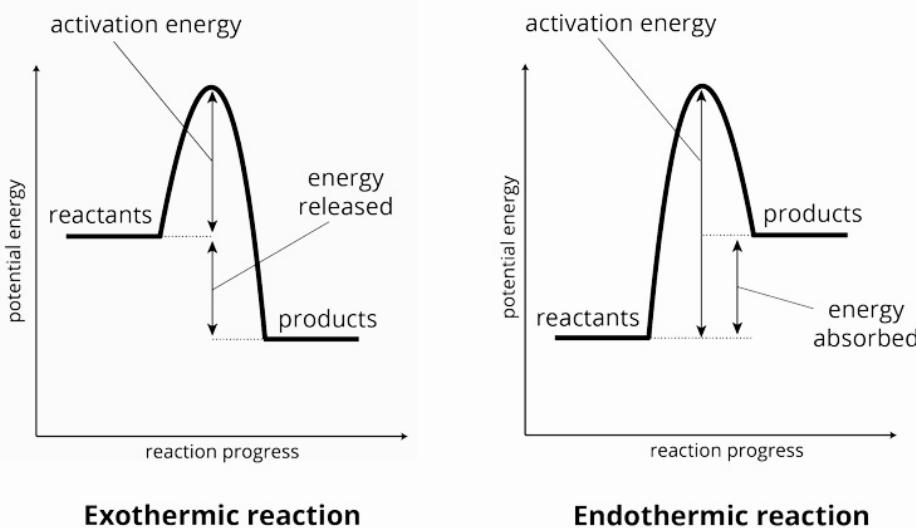
In an **endothermic reaction**, the reactants have less chemical energy than the products, so energy is added to the system for the reaction to occur. For example, the decomposition reaction of CaCO_3 is endothermic, as energy needs to be added for the reaction to occur.

KEY POINT :

You can remember this by thinking that in an **exothermic** reaction energy **exists**, and in an **endothermic** reaction energy **enters**.

The products in an exothermic reaction are more stable than the reactants. So, the bonds formed are stronger than the bonds broken. Whereas, in an endothermic reaction, the reactants are more stable than the products. So, the broken bonds are stronger than the bonds formed. You can see this in the graphs on the following page.

The **activation energy** you can see in these graphs is the amount of energy required to break the bonds of the reactants so that the reaction can occur. This is an energy barrier which occurs in both endothermic and exothermic reactions. However, as shown by the graphs, the activation energy for an endothermic reaction is larger.

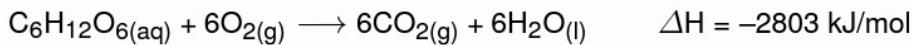


The chemical energy of a substance can be called its **enthalpy** or **heat content (H)**. The amount of energy released or absorbed in a reaction is called **enthalpy change (ΔH)**:

$$\Delta H = H_{\text{products}} - H_{\text{reactants}}$$

Therefore, since the chemical energy (H) is greater for the reactants in an exothermic reaction, ΔH is negative. So, in an endothermic reaction ΔH is positive.

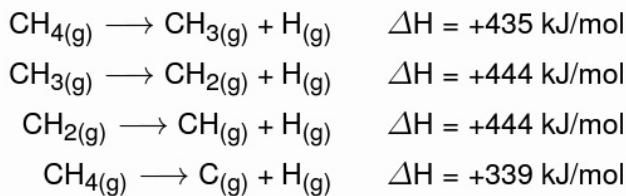
Thermochemical equations include the ΔH for the reaction. For example, the equation for photosynthesis is shown below, which is an exothermic reaction:



3.2.3 Bond enthalpies

As you already know, chemical reactions involve breaking the bonds in the reactants and then rearranging the atoms to create different products by forming bonds. Breaking chemical bonds is an endothermic process because energy is required to break the electrostatic attraction which forms the bond. Whereas, forming bonds is an exothermic process, as it releases heat. This is because bonds are formed by electrostatic attraction. This is commonly misunderstood by students so make sure you get this the right way around!

Bond enthalpy is the energy required to break the bonds in one mole of gaseous molecules under standard conditions (1 atm pressure, 25 °C). For example, breaking down methane (CH_4) requires four processes:



These equations show a C–H bond being broken; however, the ΔH values are not the same. So, the average bond enthalpy for C–H bond is calculated by averaging the ΔH values above:

$$\text{average bond enthalpy} = \frac{435 + 444 + 444 + 339}{4} = 415.5 \text{ kJ/mol}$$

The average bond enthalpies are found in the formula and data booklet, so you might need to use these in exam questions. The average bond enthalpies can be used to calculate the enthalpy change in a reaction:

$$\Delta H = \sum(\text{bonds broken}) - \sum(\text{bonds formed})$$

SAMPLE :

Let's calculate the enthalpy change for $\text{H}_{2(\text{g})} + \text{F}_{2(\text{g})} \rightarrow 2\text{HF}_{(\text{g})}$

- First using the formula sheet, we can find that the bond enthalpies for breaking H–H and F–F bonds are +436 kJ/mol and +159 kJ/mol respectively.
- Next, using the formula sheet, we find that bond enthalpy for forming two H–F bond is –1134 kJ/mol (-567×2).
- Now, we can use the formula for enthalpy change:

$$\begin{aligned}\Delta H &= (436 + 159) - (1134) \\ &= -539 \text{ kJ/mol}\end{aligned}$$

3.2.4 Hess' law

Hess' law states that if you add two thermochemical equations then the enthalpy change of the final reaction will be equal to the sum of the two previous enthalpy changes.

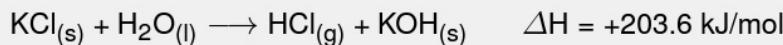
SAMPLE :

Find the enthalpy change for the following reaction: $2\text{KCl}_{(\text{s})} + \text{H}_2\text{SO}_{4(\text{l})} \rightarrow 2\text{HCl}_{(\text{g})} + \text{K}_2\text{SO}_{4(\text{s})}$.

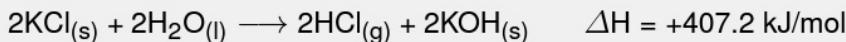
Using these reactions:

- $\text{HCl}_{(\text{g})} + \text{KOH}_{(\text{s})} \rightarrow \text{KCl}_{(\text{s})} + \text{H}_2\text{O}_{(\text{l})} \quad \Delta H = -203.6 \text{ kJ/mol}$
- $2\text{KOH}_{(\text{s})} + \text{H}_2\text{SO}_{4(\text{l})} \rightarrow 2\text{H}_2\text{O}_{(\text{l})} + \text{K}_2\text{SO}_{4(\text{s})} \quad \Delta H = -342.4 \text{ kJ/mol}$

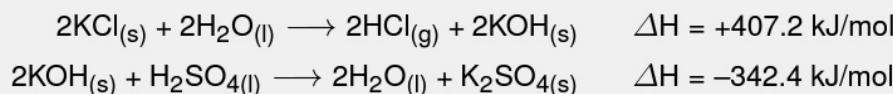
- First, we need to rearrange the equations, so we have the reactants on the right-hand side. Therefore, we need to flip the first equation. This changes the sign of the enthalpy change:



- The reactants contain two molecules of KCl. So, we need to multiply the equation above by two. The enthalpy change is also multiplied by two:



- Now we have two equations:



- If we add these equations together, we will get the total reaction:



3.2.5 Calorimetry

Calorimeters are used to measure the energy change in a reaction. The reactants are put in water in the calorimeter and the temperature change of the water is calculated, as energy is absorbed or released from/into the water. A calorimeter is a polystyrene cup with a cover, so no heat leaves the cup.

By measuring the temperature change of the water, the heat energy transferred to the water can be calculated using this formula:

$$Q = mc\Delta T$$

where Q is the heat transferred in (J), m is the mass of the water in (g), c is the specific heat capacity of the water ($4.18 \text{ J/g}^{\circ}\text{C}$), and ΔT is the change in temperature (K or $^{\circ}\text{C}$).

$$\Delta T = T_{\text{final}} - T_{\text{initial}}$$

The **specific heat capacity** is the amount of energy required to increase the temperature of a given substance by 1°C . Specific heat capacity is different for different substances. Water has the highest specific heat capacity, and this value is given in the formula and data booklet.

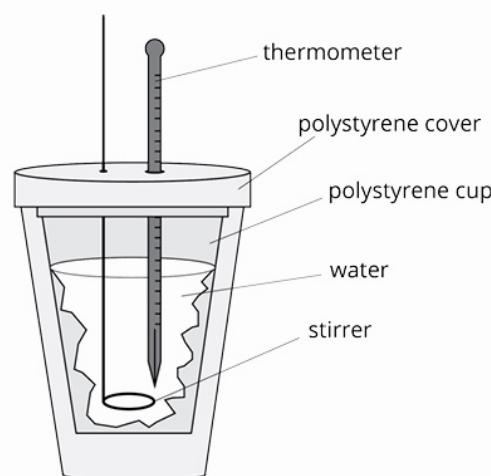
3.3 Fuels

A **fuel** is a material that is made to react with other substances to release energy, which can be used for work. The combustion of fuels provides us with energy for daily use as it is an exothermic reaction. The amount of energy released depends on the amount and type of fuel being burnt, and whether complete or incomplete combustion occurs.

Complete combustion occurs when there is a plentiful supply of oxygen and carbon dioxide is produced. Whereas, **incomplete combustion** occurs when the oxygen supply is limited, so carbon monoxide is produced.

You would have already heard of **fossil fuels**, including coal, natural gas, and petroleum. These fuels are derived from remains of organic matter produced by photosynthesis over many years. They are used faster than they can be replaced, so they are non-renewable resources. Meanwhile, **biofuels** are derived from animal or plant materials, including vegetable waste, plant oils, and sugar cane. The three main biofuels are biogas, biodiesel, and bioethanol.

I've summarised the key differences between petrodiesel (a fossil fuel) and biodiesel (a biofuel) below:



	Biodiesel	Petrodiesel
Renewability	Renewable	Non-renewable
Emissions	More carbon neutral because even though CO ₂ is produced, CO ₂ is used to produce the triglycerides used to make the fuel	CO ₂ and other pollutants
Energy content	Marginally lower	Marginally higher
Environmental impact	Spilt biodiesel will biodegrade. Biodiesel can be made from waste products (e.g. cooking oil) which has beneficial environmental effects.	Oil fields are prone to spills which damage the environment. Oil refineries produce harmful emissions.

3.4 Mole concept and law of conservation of mass

3.4.1 What is a mole?

Moles can seem like an intimidating concept in Chemistry, but a mole is just a number! For example, we call 12 eggs, 'a dozen' eggs. In the same way, we can call 6.02×10^{23} eggs 'one mole' of eggs (I know, that's a lot of eggs!). So, one mole is equal to 6.02×10^{23} . This number is referred to as **Avogadro's number** (this constant is given in the formula book). We commonly use moles as a measure of the number of particles in chemistry.

SAMPLE :

Calculate the number of molecules in 2.5 moles of water (H_2O).

We know that 1 mole = 6.02×10^{23} .

Therefore, 2.5 moles = $2.5 \times 6.02 \times 10^{23} = 1.51 \times 10^{24}$

So, in 2.5 moles of water there are 1.51×10^{24} molecules. Note that in 2.5 moles of water, there are 2.5 moles of oxygen atoms and 5 moles of hydrogen atoms.

Below is a diagram that shows how to convert between number of particles and moles.



Molar mass is the mass (in grams) of one mole of a given element or compound. For example, since one mole of carbon-12 atoms has a mass of 12 g, the molar mass of carbon-12 is 12.01 g/mol. This is because naturally occurring carbon is mostly made up of the carbon-12 isotope. Note that the molar mass of an element is equal to the relative atomic mass found on the periodic table. To calculate the number of moles, a substance's mass and molar mass can be used:

$$n = \frac{m}{M}$$

where n is the number of moles, m is the mass in (g), and M is the molar mass in (g/mol).

SAMPLE :

Let's calculate the number of molecules in 4 g of sucrose ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$):

We will use the formula: $n = \frac{m}{M}$:

$$\begin{aligned} m &= 4 \text{ g} \\ M &= 12 \times 12.01 + 22 \times 1.01 + 11 \times 16.00 \\ &= 342.34 \text{ g/mol} \end{aligned}$$

Now, we will substitute these values into the formula:

$$\begin{aligned} n &= \frac{4}{342.34} \\ &= 0.01168 \text{ mol} \end{aligned}$$

$$\begin{aligned} \text{number of molecules} &= 0.01168 \times 6.02 \times 10^{23} \\ &= 7.034 \times 10^{21} \end{aligned}$$

Therefore, in 4 g of sucrose, there is 7.034×10^{21} sucrose molecules.

3.4.2 Empirical formula

An **empirical formula** expresses the *simplest* whole number ratio of elements in a compound. This is different from the **molecular formula** which shows the *actual* number of atoms of each element in a compound. For example, ethane has a molecular formula of C₂H₆, as it has two carbon atoms and six hydrogen atoms. However, ethane's empirical formula is CH₃ as the ratio of carbon to hydrogen is simplified from 2:6 to 1:3.

It is important to recognise that for some compounds the empirical and molecular formulas are the same (e.g. water, H₂O).

First, before you learn how to write empirical formulas, you need to understand how to calculate percentage composition.

KEY POINT :

This can be calculated dividing the molar mass of the element by the molar mass of the compound:

$$\% \text{ by mass of an element in a compound} = \frac{\text{mass of the element in 1 mole of the compound}}{\text{molar mass of the compound}} \times 100$$

SAMPLE :

Let's calculate the percentage by mass of oxygen in carbon dioxide (CO₂):

1. First, we need to calculate the molar mass of carbon dioxide:

$$\begin{aligned} M(\text{CO}_2) &= 12.01 + 2 \times 16.00 \\ &= 44.01 \text{ g/mol} \end{aligned}$$

2. Next, we need to find the mass of oxygen in 1 mole of CO₂:

$$\begin{aligned} \text{mass of O in 1 mole of CO}_2 &= 2 \times M(\text{O}) \\ &= 2 \times 16.00 \\ &= 32.00 \text{ g/mol} \end{aligned}$$

3. Now we can substitute the values above into the formula:

$$\begin{aligned} \% \text{ by mass of oxygen in CO}_2 &= \frac{32.00}{44.01} \times 100 \\ &= 72.71\% \end{aligned}$$

Therefore, the percentage by mass of oxygen in CO₂ is 72.71%.

Now, you can calculate the molecular formula of a compound given percentage by mass values.

SAMPLE :

Calculate the molecular formula of a compound that is made up of 85.60% carbon and 14.40% hydrogen, which has a molar mass of 154 g/mol.

- First, assuming we have 100 g of this compound, calculate the masses for the elements present:

$$m(C) = 85.60 \text{ g}$$

$$m(H) = 14.40 \text{ g}$$

- Next, we need to calculate the number of moles of each element:

$$\begin{aligned} n &= \frac{m}{M} \\ n(C) &= \frac{85.60}{12.01} \\ &= 7.1274 \text{ mol} \\ n(H) &= \frac{14.40}{1.01} \\ &= 14.2574 \text{ mol} \end{aligned}$$

- Now, we need to calculate the simplest ratio of the moles of each atom:

$$\begin{aligned} \text{C : H} \\ 7.1274 : 14.2574 \\ \frac{7.1274}{7.1274} : \frac{14.2574}{7.1274} \\ 1 : 2 \end{aligned}$$

- Write the empirical formula: CH_2

- Find the molar mass of CH_2 :

$$\begin{aligned} M(\text{CH}_2) &= 12.01 + 2 \times 1.01 \\ &= 14.03 \text{ g/mol} \end{aligned}$$

- Now, we need to calculate how many CH_2 units make up this compound:

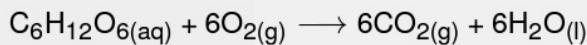
$$\begin{aligned} \text{number of } \text{CH}_2 \text{ units} &= \frac{154}{14.03} \\ &= 11 \\ \text{molecular formula} &= 11 \times \text{CH}_2 \\ &= \text{C}_{11}\text{H}_{22} \end{aligned}$$

3.4.3 Stoichiometry

Stoichiometry is the study of ratios of moles of substances. The coefficients in chemical equations tell us the ratios of moles for each reactant and product in a chemical reaction. Let's take a look at an example.

SAMPLE :

Calculate how many moles of glucose ($C_6H_{12}O_6$) and oxygen gas are required to make 15 moles of carbon dioxide and 15 moles of water:



In the reaction above, one mole of glucose reacts with six moles of oxygen gas to form six moles of carbon dioxide and six moles of water. We can write this information as a mole ratio:

$$C_6H_{12}O_6 : O_2 : CO_2 : H_2O = 1 : 6 : 6 : 6$$

So, we know that there are 15 moles of carbon dioxide produced. We will use the ratios above to calculate the amount of reactants needed:

$$\begin{aligned} C_6H_{12}O_6 : CO_2 &= 1 : 6 \\ n(C_6H_{12}O_6) &= 6 \times n(CO_2) \\ &= 6 \times 15 \\ &= 90 \text{ mol} \end{aligned}$$

Now, we will repeat the process again to work out the amount of oxygen gas that is needed:

$$\begin{aligned} O_2 : CO_2 &= 6 : 6 = 1 : 1 \\ n(O_2) &= n(CO_2) \\ &= 15 \text{ mol} \end{aligned}$$

Therefore, 2.5 moles of glucose and 15 moles of oxygen gas are required.

Questions are a bit more complex if they give you the mass rather than the number of moles of the reactant/product. Let's take a look at another example.

SAMPLE :

If 48 g of magnesium burns in the presence of oxygen gas to produce magnesium oxide, what is the mass of oxygen gas required?

1. The first thing you need to do is to write the *balanced* chemical equation. It is extremely important to do this! Otherwise, you won't know the mole ratios.



Next, we need to calculate the number of moles of magnesium:

$$\begin{aligned} n &= \frac{m}{M} \\ n(\text{Mg}) &= \frac{48}{24.31} \\ &= 1.9745 \text{ mol} \end{aligned}$$

2. Use the mole ratio to calculate the number of moles of oxygen gas:

$$\begin{aligned} \text{Mg : O}_2 &= 2 : 1 \\ n(\text{O}_2) &= n(\text{Mg}) \div 2 \\ &= 1.9745 \div 2 \\ &= 0.98725 \text{ mol} \end{aligned}$$

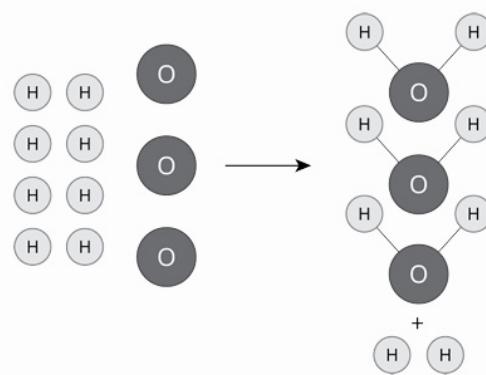
3. Now let's calculate the mass of oxygen gas:

$$\begin{aligned} m &= n \times M \\ &= 0.98725 \times 16.00 \\ &= 15.796 \text{ g} \end{aligned}$$

Therefore, 15.796 g of oxygen gas is required.

Now we need to consider reactions where the reactants aren't fully used up. When this occurs, the reactant that is not fully consumed is the **excess reactant**. The reactant that is fully used up is called the **limiting reactant** and it stops the reaction from continuing further. You can see this visually in the diagram on the right.

In this reaction, hydrogen and oxygen atoms react to form water. You can see that oxygen is the limiting reactant and hydrogen is the excess reactant. This is because there is left-over hydrogen that does not react. Now let's do a calculation for a reaction where there are limiting and excess reactants.

**SAMPLE :**

50.0 g of aluminium nitrate reacts with 40.0 g of dissolved sodium hydroxide. Find the mass of the aluminium hydroxide formed:

- First, we need to write a balanced chemical equation. This reaction is a double-displacement reaction:



- Next, we need to determine which reactant is the limiting reactant:

$$\begin{aligned} n(\text{NaOH}) &= \frac{m}{M} \\ &= \frac{40}{40} \\ &= 1 \text{ mol} \end{aligned}$$

$$\begin{aligned} n(\text{Al}(\text{NO}_3)_3) &= \frac{50}{213.01} \\ &= 0.2347 \text{ mol} \end{aligned}$$



If 0.2347 moles of $\text{Al}(\text{NO}_3)_3$ reacts, the amount of moles of NaOH needed is:

$$\begin{aligned} n(\text{NaOH}) &= 3 \times n(\text{Al}(\text{NO}_3)_3) \\ &= 0.7041 \text{ mol} \end{aligned}$$

Since there is 1 mole of NaOH, we know that NaOH is in excess.

- To calculate the mass of Al(OH)_3 we will use the mole ratio with the limiting reactant, $\text{Al}(\text{NO}_3)_3$:

$$\begin{aligned} \text{Al(OH)}_3 : \text{Al}(\text{NO}_3)_3 &= 1 : 1 \\ n(\text{Al(OH)}_3) &= n(\text{Al}(\text{NO}_3)_3) \\ &= 0.2347 \text{ mol} \\ m(\text{Al(OH)}_3) &= n \times M \\ &= 0.2347 \times 78.01 \\ &= 18.31 \text{ g} \end{aligned}$$

3.4.4 Experiment yield

In an experiment, you may do calculations using stoichiometry before conducting your experiment to predict how much product should be formed in a reaction. This value is called the **theoretical yield**, and it assumes that all of the reactants are converted into products. However, when you do the experiment, the amount of product that you create may be slightly different to your predicted value. The amount of product that is created in the experiment is called the **experimental yield**.

There are many reasons why the experimental yield may be less than the theoretical yield in an experiment. It may be due to experimental error or maybe the reaction rate is very slow, so the reaction didn't have enough time to go to completion. Another possible reason is that the reaction does not go to completion, as it reaches a state of equilibrium. You will learn more about equilibria in Unit 3.

When you analyse your results from the experiment, you may calculate the **percentage yield**. This measures the efficiency of the production process in the experiment. The higher the percentage yield, the higher degree of conversion from reactants to products:

$$\text{percentage yield} = \frac{\text{experimental yield}}{\text{theoretical yield}} \times 100$$

The equation above is given in the formula and data book.

Part II

Unit 2: Molecular interactions and reactions

Section 1

Intermolecular forces and gases

1.1 Intermolecular forces

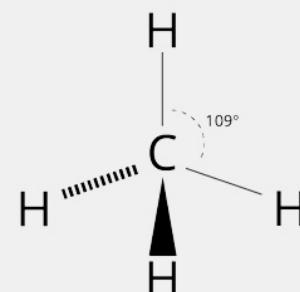
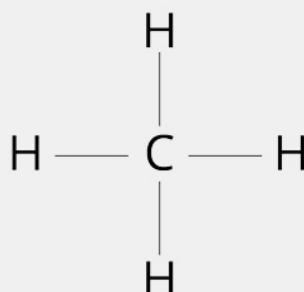
1.1.1 VSEPR theory

You already know that electrons repel each other. We can use this fact to determine the shape of molecules due to the repulsion of electrons. This is stated by the **valence shell electron pair repulsion (VSEPR) theory**, which says that electron pairs repel each other so they are as far away as possible.

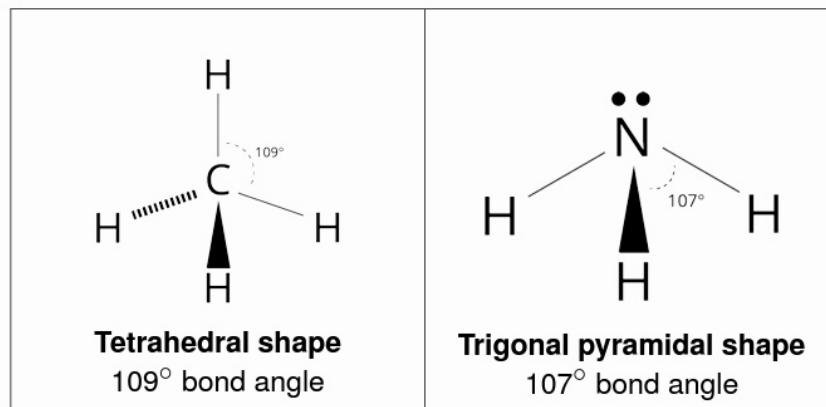
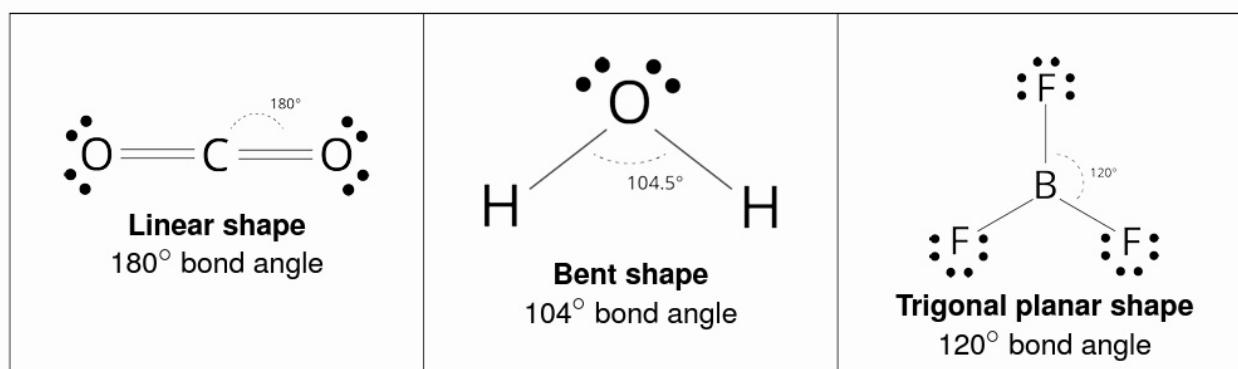
SAMPLE :

Let's have a look at methane (CH_4). To determine its shape we need to think about the electron pairs and how they repel each other. To visualise this, I've drawn the Lewis dot diagram below on the left.

The bonding electron pairs repel each other, and this gives the molecule **tetrahedral** geometry and bond angles of 109° . Note that a thick line represents a bond coming out of the page and a dashed line shows a bond going into the page.



There are many different shapes which molecules can be, and I've listed the main geometries below.

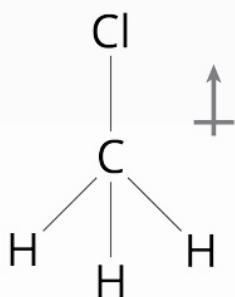


1.1.2 Polarity of molecules

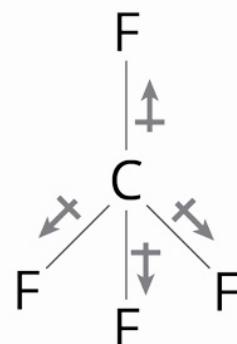
We have already discussed that a polar bond occurs when there is a large difference in electronegativity between the atoms involved. As the difference in electronegativity increases, the bond becomes more polar.

However, compounds containing polar bonds aren't necessarily polar. All asymmetrical molecules are polar because they have a net dipole. A **dipole** is a separation of positive and negative charges, which is represented by an arrow pointing from the positive charge to the negative charge.

For example, the molecule below is polar. Note that carbon to hydrogen bonds are considered non-polar, due to the small difference in electronegativity. In the carbon to chlorine bond, the carbon is positively charged, and the chlorine is negatively charged, because chlorine is more electronegative than carbon.



Meanwhile, symmetrical molecules are non-polar as the dipoles may cancel each other out. Carbon tetrafluoride (CF_4) is an example of this, shown below.

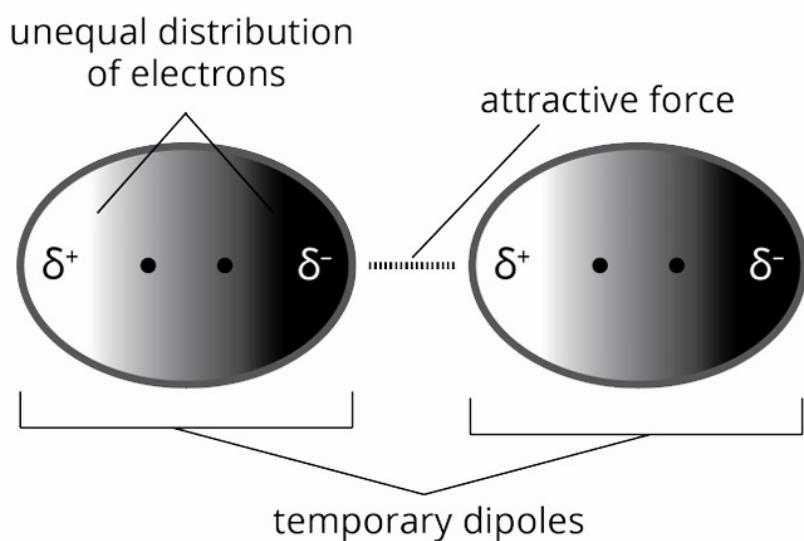


1.1.3 Intermolecular forces

Intermolecular forces exist between molecules and hold compounds together. Make sure that you don't get these confused with **intramolecular forces** which exist inside molecules and hold the atoms together. Intramolecular forces include covalent, metallic, and ionic bonding.

There are three types of intermolecular forces: **dispersion forces**, **dipole-dipole forces**, and **hydrogen bonding**. The strength and type of intermolecular forces are determined by electron configurations.

Dispersion forces are present in every molecule. These forces are due to instantaneous dipoles, which are caused by the random movement of electrons throughout molecules. A visual representation of this is shown below:

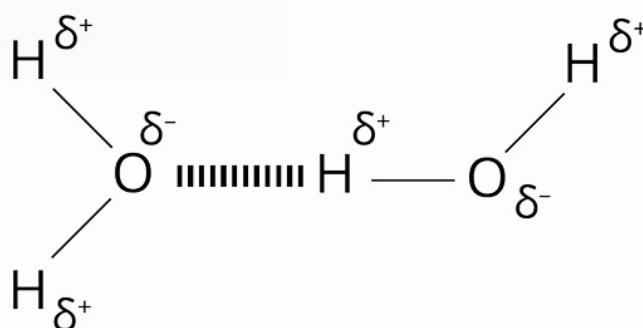


Larger molecules with greater molar masses have stronger dispersion forces because there are more electrons present. Also, these electrons are on average further away from the nucleus, so they are more easily distorted. This increases the chance of instantaneous dipoles occurring.

Dipole-dipole forces occur between polar molecules. This results from the attraction between the positive and negative ends of polar molecules. The more polar the molecules are, the stronger these interactions. This is shown by the diagram below:



Hydrogen bonding is a strong form of dipole-dipole forces. This only occurs between hydrogen and either oxygen, nitrogen, or fluorine. This is because these atoms are highly electronegative, so the molecules are very polar. Water is a good example of a compound containing hydrogen bonding. This is shown in the diagram below:



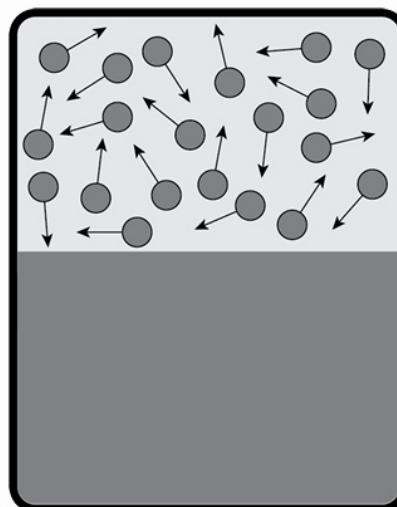
1.1.4 Properties and intermolecular forces

The order of intermolecular forces from weakest to strongest is the following:

$$\text{dispersion forces} < \text{dipole-dipole forces} < \text{hydrogen bonding}$$

However, when comparing molecules of very different molar masses, dispersion forces may outweigh other intermolecular forces. This is because dispersion forces are very strong in larger molecules.

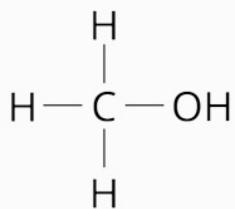
Vapour pressure is the pressure that gaseous molecules exert on the walls of a closed container when rates of condensation and evaporation are equal. Evaporation occurs when molecules in the liquid escape from the surface and become a gas. The stronger the intermolecular forces in the liquid, the more strongly the molecules are held together, so it is harder for them to escape. Therefore, less evaporation occurs, so vapour pressure is lower.



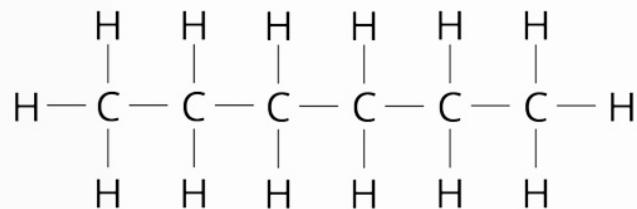
Now, we need to look at how to determine **solubility**. Generally, if the solute and solvent have similar intermolecular forces, then the solute will dissolve.

For example, methanol dissolves in water because both molecules are polar and can form hydrogen bonding. By contrast, hexane (C_6H_{14}) cannot dissolve in water because hexane only contains dispersion forces and water has hydrogen bonding.

methanol



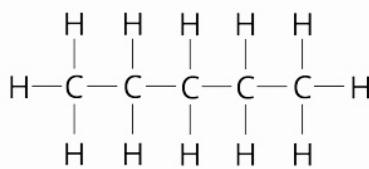
hexane



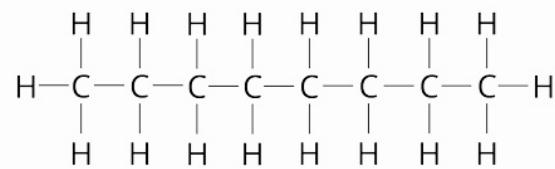
This is because for a solute to dissolve, the intermolecular forces in the solute and the solvent need to be broken first. If the energy released, when the intermolecular forces form between the solute and the solvent, is not enough to compensate for this energy required, then the solute won't dissolve in the solvent. Therefore the energy released, when dispersion forces form between hexane and water, won't compensate for the energy required to break the strong hydrogen bonds between the water molecules.

Non-polar solutes dissolve in non-polar solvents because they only contain dispersion forces. For example, pentane (C_5H_{12}) dissolves in octane (C_8H_{18}).

pentane



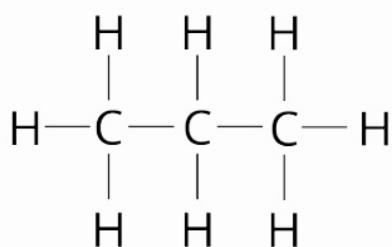
octane



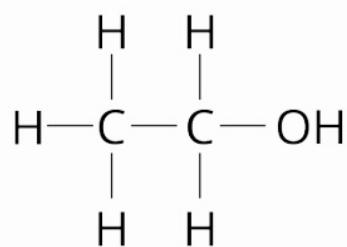
We also need to consider the impact of intermolecular forces on **melting and boiling point**. Substances containing stronger intermolecular forces generally have higher melting and boiling points. This is because more energy is required to overcome the attraction between the molecules.

For example, propane (C_3H_8), which only contains dispersion forces, has a boiling point of -42°C , whereas ethanol (C_2H_6O), which contains hydrogen bonding, has a boiling point of 78°C , and water (which contains stronger hydrogen bonds) has a boiling point of 100°C .

propane



ethanol

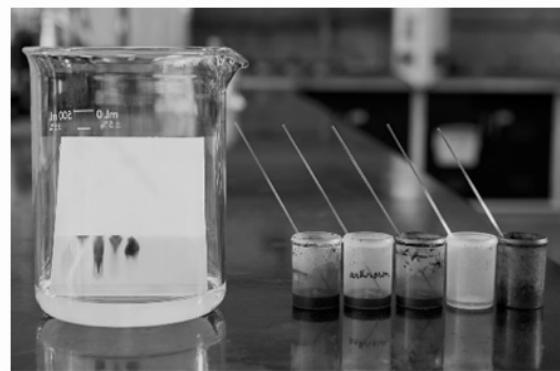


1.2 Chromatography techniques

Chromatography is a technique used by chemists to separate components in a mixture. There are many types of chromatography, but we will be focusing on paper, thin layer, gas, and high-performance liquid chromatography.

1.2.1 Paper and thin layer chromatography

Paper and thin layer chromatography are the two simplest forms of chromatography. You may have even done an experiment on this in previous years at school. Have you ever had a piece of white paper and drawn dots using different coloured pens on it, and then put the bottom of this paper in water? If you have you would have seen that the different coloured inks separate as they run up the paper. You would get something like what is shown in the image on the right. This is paper chromatography! The original position where the ink spots were drawn on the paper is called the **origin**.

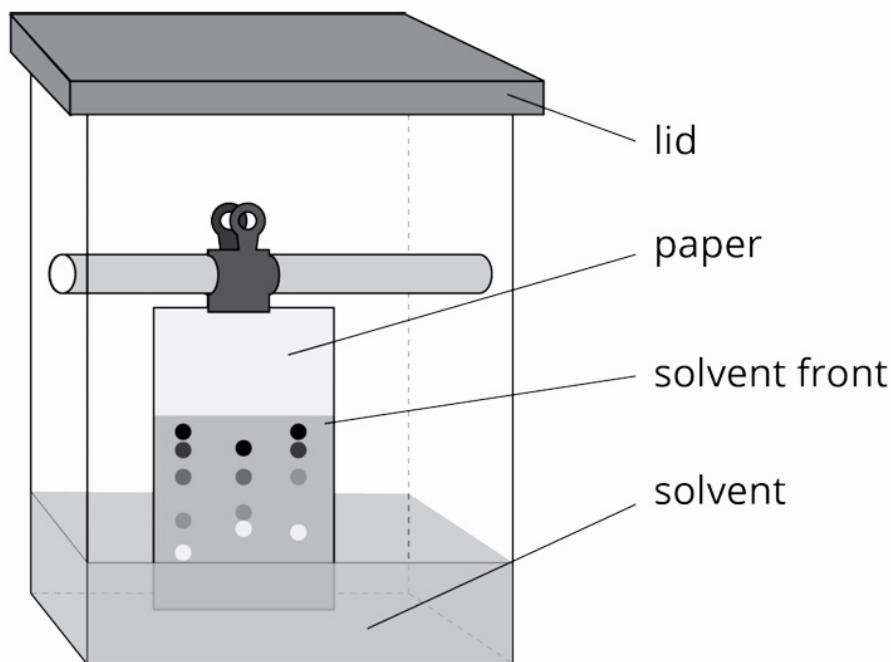


<https://www.britannica.com/science/paper-chromatography>. Image © Ggw1962/Dreamstime.com

In all types of chromatography, there is a stationary phase and mobile phase. The stationary phase is either a solid or liquid, and the mobile phase is either a liquid or gas.

In this case, the **stationary phase** is the high-quality absorbent paper and the **mobile phase** is the water travelling up the paper. The ink travels up the page with the solvent (i.e. the water) and continuously undergoes **desorption** into the mobile phase and **absorption** onto the stationary phase. The distance that the ink components travel up the page is determined by how readily it absorbs onto the stationary phase and dissolves into the mobile phase. This is determined by intermolecular forces because we already know that a solute will dissolve in a solvent if they have similar intermolecular forces. Therefore, since water is a polar solvent, polar molecules will dissolve in the water and travel further up the page than non-polar molecules.

Note that in paper chromatography the mobile phase doesn't have to be water.



Before the water reaches the end of the paper, it is removed from the beaker and the distance which the solvent travelled is measured (this is called the **solvent front**). The distances which each component of the ink travelled are also recorded. This allows us to calculate a **retardation factor** (R_f) for the sample:

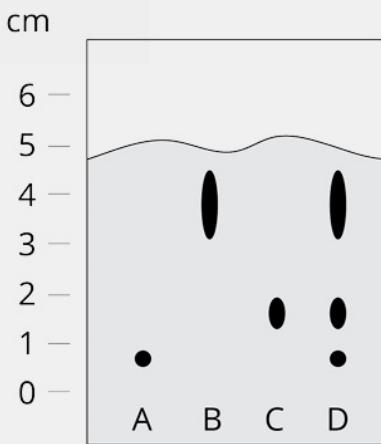
$$R_f = \frac{\text{distance component travelled from the origin}}{\text{distance solvent front travelled from the origin}}$$

KEY POINT :

Note that retardation factors are always less than 1. The higher the retardation factor, the more strongly the component dissolved into the mobile phase, and the further it travelled.

SAMPLE :

Let's calculate the retardation factor for sample B shown below:



$$\begin{aligned} R_f &= \frac{\text{distance component travelled from the origin}}{\text{distance solvent front travelled from the origin}} \\ &= \frac{4 \text{ cm}}{4.7 \text{ cm}} \\ &= 0.85 \text{ cm (rounded to two decimal places)} \end{aligned}$$

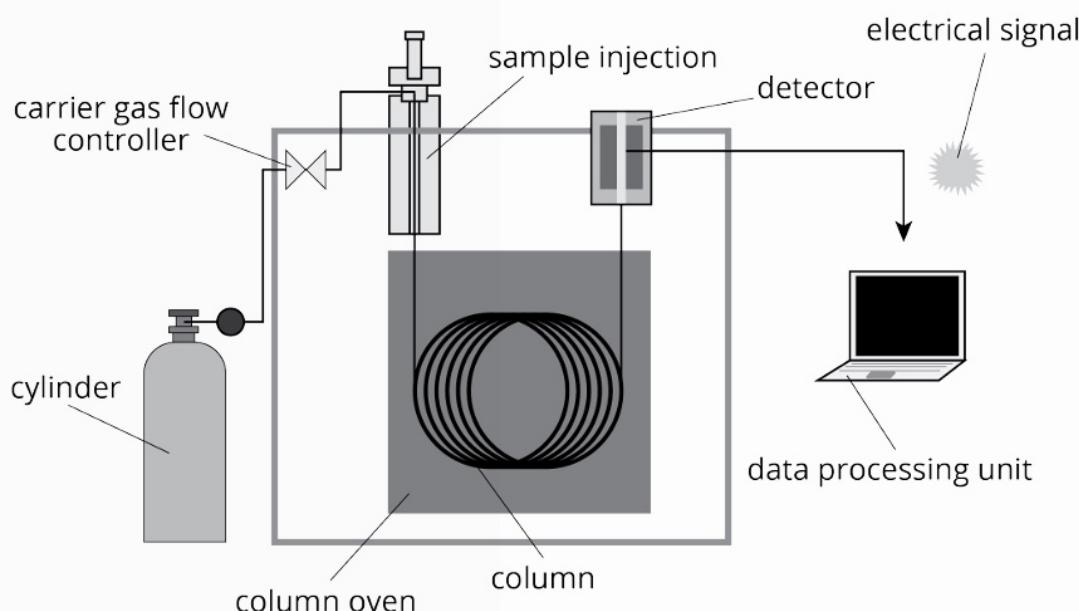
Therefore, sample B was strongly attracted to the mobile phase and moved far up the paper.

Thin layer chromatography is the same process as paper chromatography; however, instead of paper, the stationary phase is a thin layer of silica gel spread over a glass plate.

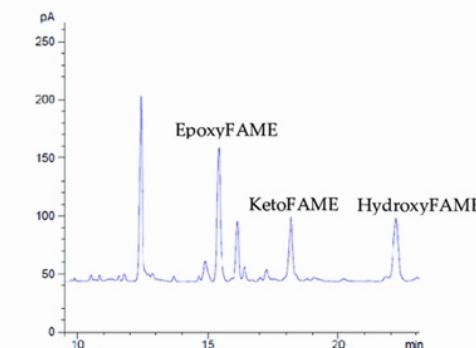
1.2.2 Gas chromatography

In gas chromatography, the sample being analysed is first heated to vaporise it. Gas chromatography has many applications, including drug testing in sport. There are two types of gas chromatography: **gas-liquid chromatography (GLC)** and **gas-solid chromatography (GSC)**.

In both types of gas chromatography, the mobile phase is the carrier gas, which is generally nitrogen gas. The gaseous sample is swept through a series of loops by the carrier gas. In GLC, these columns are filled with a porous solid coated with liquid ester or hydrocarbon. Therefore, GLC has a liquid stationary phase. In comparison, GSC has a solid stationary phase, as the columns are packed with an absorbent solid such as silica gel.



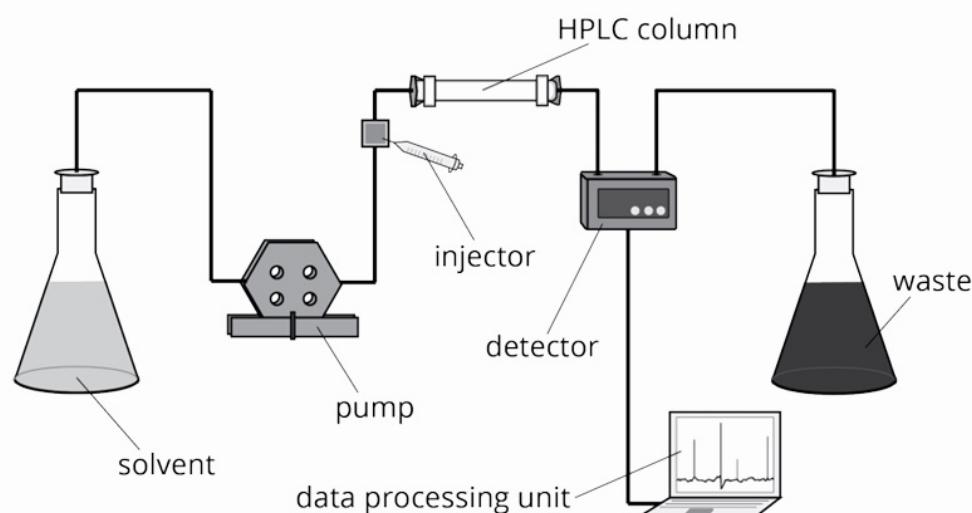
Instead of calculating retardation factors (like in paper and thin layer chromatography), in gas chromatography we measure **retention times**. This is the amount of time it takes for the gas to travel through the column. The longer it takes, the more the gas absorbs onto the stationary phase. This recorded data is used to create a **chromatograph** which shows the different components of the sample. For example, the chromatograph on the right shows that HydroxyFAME had the longest retention time. Each peak on the chromatograph represents a different component of the mixture.



Source: Khor, et al. 2019, www.researchgate.net/figure/a-Sample-gas-chromatograph-GC-chromatogram-for-the-determination-of-epoxy-keto-and_figs1_336447190, CC-BY-SA-4.0

1.2.3 High-performance liquid chromatography

High-performance liquid chromatography (HPLC) is extremely important as it is used to detect the presence of oil spills in water and pesticides in food. The mobile phase is a liquid which is pumped through a column containing small particles (which make up the stationary phase). The sample is injected into the column and the time it takes to flow through it (retention time) is measured. The separated components that leave the column with the solvent are called the **eluent**. To detect the different components, a beam of UV light is passed through the eluent stream. This is because most organic compounds absorb UV light, so the detector measures the reduced signal. This information is used to create a chromatograph.



1.3 Gases

1.3.1 Kinetic theory of gases

We've already discussed some of the properties of solids, liquids, and gases, but I've summarised these in the table below.

	Solids	Liquids	Gases
Volume and shape	Fixed volume and shape	Fixed volume, takes the shape of the container	Takes volume and shape of the container
Density	High	High	Low
Ability to mix	Don't mix	Require stirring	Mix rapidly
Compressibility	Almost incompressible	Almost incompressible	Compressible

These properties are caused by the gas particles' ability to move independently of each other, so they are more spread out. In comparison, particles in solids are only able to vibrate, and liquid particles are only free to move relative to each other. Gas particles are very small so most of the volume occupied by a gas is empty space, as the intermolecular forces between gas particles are weak and are easily broken.

You've already learnt about kinetic theory, which states that all particles are moving and as temperature increases kinetic energy of particles increase. So, gas particles are in constant random motion, as they freely move and collide with each other. These collisions are **elastic** as kinetic energy is transferred between the particles.

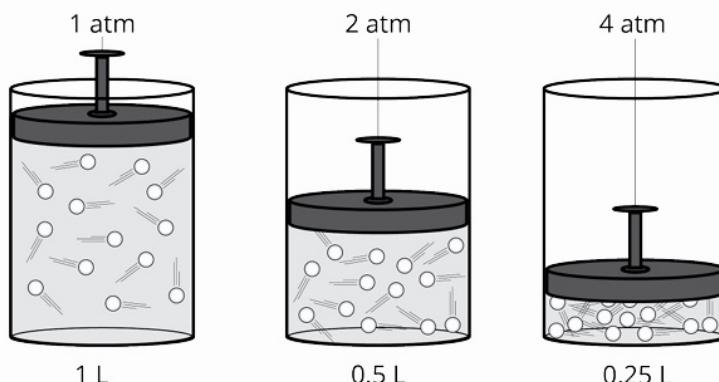
1.3.2 Volume, pressure, and temperature

Let's define the terms volume and pressure:

1. **Volume:** the amount of space that a substance occupies.
 - Commonly expressed in litres (L).
 - $1 \text{ L} = 1 \text{ dm}^3 = 1000 \text{ cm}^3 = 0.001 \text{ m}^3 = 1000 \text{ mL}$
2. **Pressure:** the force applied on a surface compared to the surface's area.

$$\text{Pressure} = \frac{\text{force}}{\text{area}}$$
 - Therefore, the units of pressure are N/m².
 - $1 \text{ N/m}^2 = 1 \text{ Pa}$ (pascal)

There is a relationship between the pressure, temperature, and volume of a gas. As volume decreases, pressure increases if the amount of gas and temperature remain constant. So, volume is inversely proportional to pressure ($V \propto \frac{1}{P}$). This is because the more a gas is compressed, the more collisions there are between gas particles and the walls of the container. Therefore, pressure increases, because more force is exerted on the walls of the container.



As the number of moles of gas increases, volume increases if temperature and pressure remain constant. So, volume is directly proportional to the number of moles ($V \propto n$). This is very logical because if you blow into a balloon, the volume increases as the balloon expands.

As temperature increases, volume increases if pressure and the amount of gas remain the same. So, volume is directly proportional to temperature ($V \propto T$). For example, if you place the balloon over a flame, the balloon expands, and it may eventually pop. This is because the gas particles gained kinetic energy as the temperature increased, which caused more collisions between the gas particles and the edge of the balloon. This caused the volume to increase.

1.3.3 Ideal gas equation

Ideal gases obey kinetic theory and the gas laws. Most gases act like ideal gases at particularly high temperatures and low pressures. However, at very high pressures and low temperatures, gases tend to not behave like ideal gases.

The behaviour of ideal gases under different conditions is described by the **ideal gas equation**:

$$PV = nRT$$

Where P is pressure (in kPa), V is volume (in L), n is number of moles, T is temperature (K), and R is the ideal gas constant ($8.31 \text{ J K}^{-1} \text{ mol}^{-1}$).

This equation considers the relationships that we considered above:

$$V \propto T$$

$$V \propto n$$

$$V \propto \frac{1}{P}$$

SAMPLE :

Calculate the pressure of 5 moles of oxygen gas if the volume is 20 L at 40°C :

- First, we need to convert the temperature into Kelvin:

$$\begin{aligned}\text{Temperature in K} &= \text{temperature in } {}^\circ\text{C} + 273 \\ &= 40 + 273 \\ &= 313 \text{ K}\end{aligned}$$

- Now, we can substitute the values into the ideal gas equation and rearrange to find the pressure:

$$\begin{aligned}P \times 20 &= 5 \times 8.31 \times 313 \\ P &= \frac{13005.15}{20} \\ &= 650.26 \text{ kPa}\end{aligned}$$

In some questions, you may be given the **molar volume (V_m)** instead of the volume of the gas. Molar volume is the volume that 1 mole of any gas occupies at a specific temperature and pressure. For example, 1 mole of any ideal gas is 22.7 L at **standard temperature and pressure (STP)**, which is 0°C and 100 kPa. Therefore, the molar volume is 22.7 L/mol for any gas at STP. To calculate the volume of a particular amount of gas from the molar volume use the formula below.

$$V_m = \frac{V}{n}$$

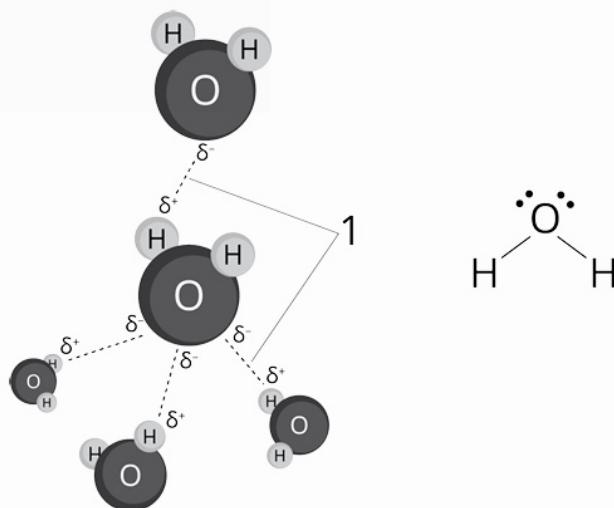
Section 2

Aqueous solutions and acidity

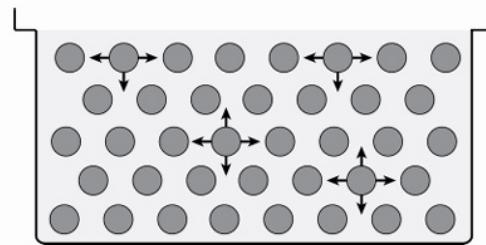
2.1 Aqueous solutions and molarity

2.1.1 Unique properties of water

Water has unique properties due to its ability to form hydrogen bonding. This is because water forms strong intermolecular forces as each water molecule can form **four hydrogen bonds**. A hydrogen bond is formed when a lone pair on an oxygen atom is attracted to a hydrogen atom's partial positive charge. Remember that the oxygen atom in a water molecule has two lone pairs on electrons, as shown below.



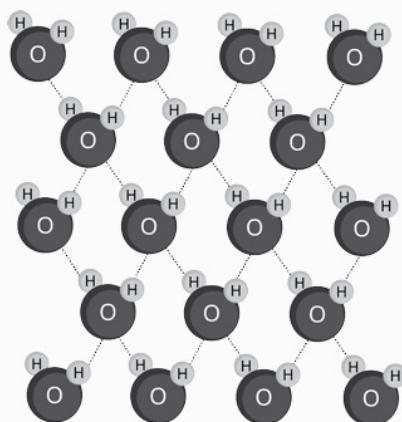
Water is unique because it has a **high surface tension**, which is a liquid's resistance to increasing its surface area. This is why insects can walk across water. This is due to the strong hydrogen bonds between neighbouring water molecules on the surface. At the surface, water molecules have an overall downwards force. By contrast, water molecules in the middle of the liquid experience no net force. You can see this in the diagram.



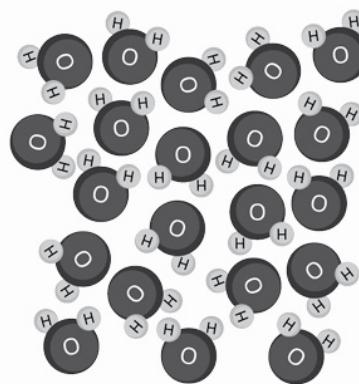
Water also has a **very high boiling point**. This is because the strong hydrogen bonds require a large amount of energy to break, so the particles in liquid water can separate and become water vapour. The table below compares water's boiling point to other group 16 hydrides. You can see that water's boiling point is far higher than the other compounds' boiling points.

Compound	Boiling point (°C)
H ₂ O	100
H ₂ S	-60.7
H ₂ Se	-41.5
H ₂ Te	-2.2
H ₂ Po	36.1

Water is unusual because it has a higher density as a liquid than it does as a solid. **Density** is defined as mass per unit volume, as it describes a substance's compactness. This is because ice is made up of a lattice of water molecules, causing ice to have a greater volume than liquid water. This is represented visually in the diagram below.



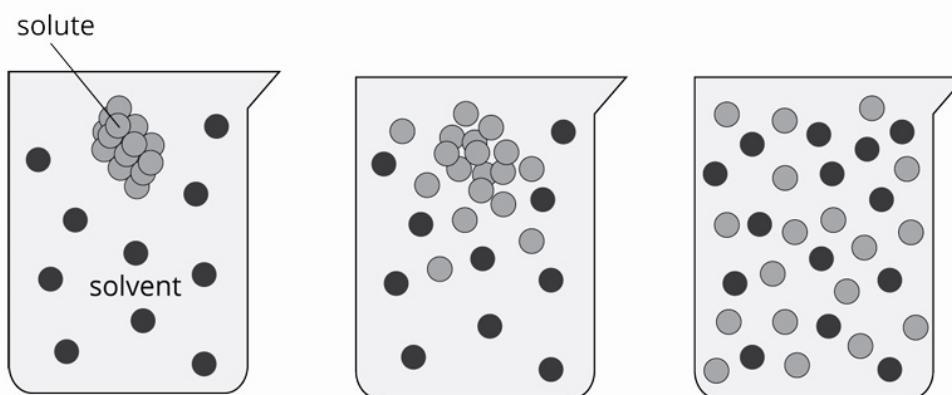
Water molecules in solid ice



Water molecules in liquid water

2.1.2 Solutions

You probably have already learnt about solutions previously at school, but let's review what a solution is. A **solution** is made up of a **solute** dissolved in a **solvent** (which is usually a liquid). Water is a common solvent used to create solutions. **Dissolution** is the process of dissolving one substance into another. If both substances are liquids, you can say that they are **miscible**. When a solute dissolves, the intermolecular forces holding the solvent together need to be broken. If water is the solvent, these forces are hydrogen bonds. The intermolecular forces holding the solute together also need to be broken, so new intermolecular forces can form between the solute and solvent particles.



A substance is soluble if the solute–solute and solvent–solvent intermolecular forces are weaker than the solute–solvent interactions. Therefore, 'like dissolves like', as a solute will dissolve in a solvent if their intermolecular forces are similar. So, polar solutes dissolve in polar solvents, and non-polar solutes dissolve in non-polar solvents.

2.1.3 Concentration

Concentration is the amount of solute dissolved in a certain volume of solution. For example, when you are making cordial, you would describe the drink as concentrated if you've added a lot of cordial in relation to the amount of water. If there is a lot more water than cordial in the glass, you would describe it as dilute.

Concentration can be expressed in many different units. First, let's look at how concentration can be described as the **mass of solute per litre of solution**. This is calculated using the formula below.

$$\text{concentration (g/L)} = \frac{\text{mass of solute (g)}}{\text{volume of solution (L)}}$$

2.1 Aqueous solutions and molarity

Another unit used to describe concentration is **parts per million** (ppm). This is calculated using the formula below:

$$\text{concentration (ppm)} = \frac{\text{mass of solute (mg)}}{\text{mass of solution (kg)}}$$

It's important to remember that 1 g = 1,000 mg.

SAMPLE :

A solution contains 0.002 g of sodium chloride dissolved in 1,500 g of solution. Find the concentration of the solution in ppm:

$$\begin{aligned}\text{concentration (ppm)} &= \frac{\text{mass of solute (mg)}}{\text{mass of solution (kg)}} \\ &= \frac{2 \text{ mg}}{1.5 \text{ kg}} \\ &= 1.33 \text{ ppm}\end{aligned}$$

Make sure that you remember to convert the units in the question to the appropriate units for the formula!

In chemistry, **molarity** is commonly used to describe concentration. Molarity is a solution's concentration in moles per litre (M). This is calculated using the formula below.

$$n = CV$$

where n is the number of moles, C is the concentration (M), and V is the volume (L).

Note that 1 L of solution containing 1 mole of solute has a molarity of 1 M.

SAMPLE :

Find the concentration in mol/L of a solution containing 20 mg of NaCl dissolved in 100 mL of solution:

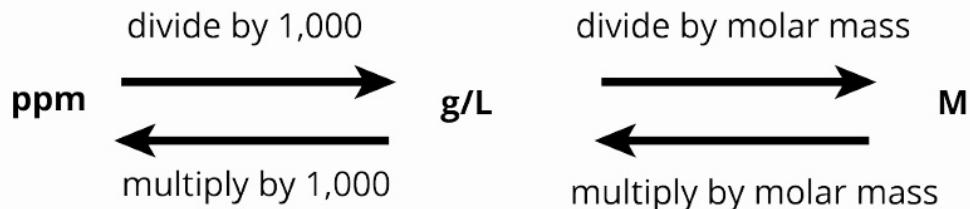
1. First, we need to calculate the number of moles of NaCl:

$$\begin{aligned}n &= \frac{m}{M} \\ &= \frac{20 \times 10^{-3}}{58.44} \\ &= 3.42 \times 10^{-4} \text{ mol}\end{aligned}$$

2. Now, we can substitute the values into the formula for molarity:

$$\begin{aligned}C &= \frac{n}{V} \\ &= \frac{3.42 \times 10^{-4}}{0.1} \\ &= 3.42 \times 10^{-3} \text{ M}\end{aligned}$$

The methods to convert between M, g/L, and ppm are given below.



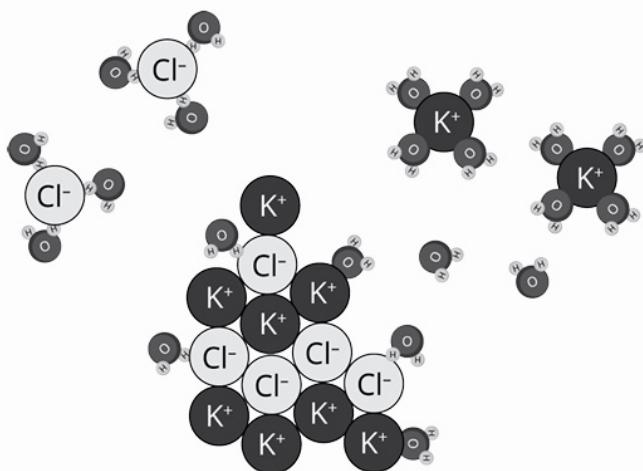
2.1.4 Unsaturated, saturated, supersaturated solutions

There are three types of solutions that you need to be familiar with: unsaturated, saturated, and supersaturated solutions.

No more solute can be added to **saturated solutions** at a particular temperature. Whereas, more solute can be added to **unsaturated solutions**, as they contain less solute than is required to make the solution saturated. Meanwhile, **supersaturated solutions** are unstable as they contain more dissolved solute than a saturated solution. If disturbed, then some solute will separate from the solvent to form a solid. This process is called **crystallisation**, and it is how crystals are formed.

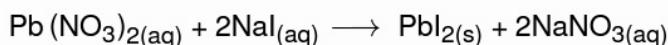
2.2 Identifying ions in solution

Most ionic compounds are soluble in water. For example, potassium chloride readily dissolves in water. When this occurs, the hydrogen bonds in water are broken, and the lattice of K^+ and Cl^- ions is broken down as the ionic bonds are broken. This occurs as the positively charged end of a water molecule is attracted to the negatively charged Cl^- anion. The negatively charged end of a water molecule is attracted to the positively charged K^+ cation. These interactions are called **ion-dipole attractions**. If these attractions are strong enough, the ions in the ionic compound can be drawn out of the lattice. This causes **dissociation**, which is when ions can move freely in solution.

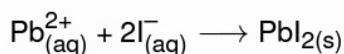


Not all ionic compounds are soluble in water. To determine which ionic compounds are soluble in water, a **solubility table** can be used. This is given in the QCAA formula and data booklet. You can also remember that all ionic compounds containing group 1 metals, nitrates (NO_3^-), ammonium (NH_4^+), and ethanoate (CH_3COO^-) are always soluble in water. For example, using the formula sheet, you can see that barium sulphate (BaSO_4) is insoluble in water.

Precipitation reactions occur when ions in solution react to produce a new compound that is insoluble in water. This is called the **precipitate**. For example, aqueous lead (II) nitrate reacts with sodium iodide to produce lead (II) iodide and sodium nitrate. Using the formula sheet, you can see that sodium nitrate is soluble in water; however, lead (II) iodide is not. Therefore, lead (II) iodide is the precipitate. The chemical equation is given below:

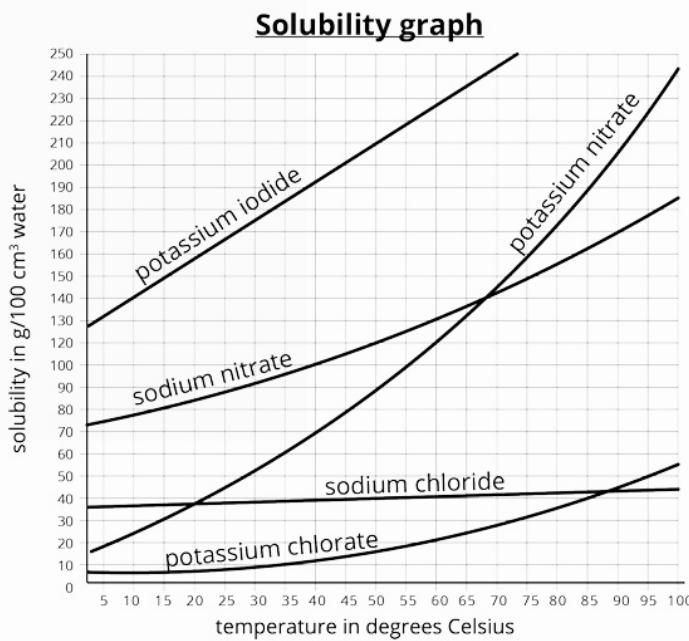


The reaction above is the **full equation**. **Ionic equations** are a special type of equation which only shows the reacting substances. These equations don't include **spectator ions** which are not involved in forming the precipitate. For this reaction, the ionic equation is given below, which omits the nitrate and sodium ions:



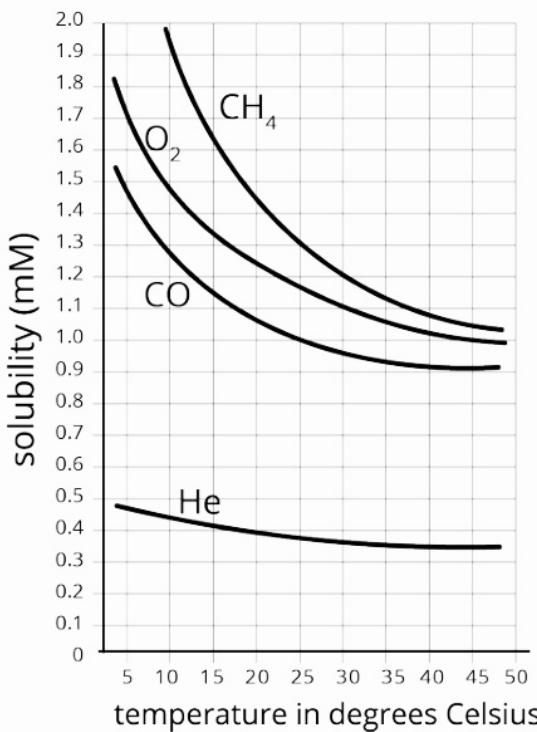
2.3 Solubility curves

Most solids become more soluble in liquid when temperature increases. This is because the particles in the solid have more energy to overcome the forces of attraction between them. This relationship between solubility and temperature can be shown by **solubility curves**.



Each point on a solubility curve represents the maximum amount of solute that can dissolve in 100 g of water. So, the graph represents saturated solutions. Therefore, unsaturated solutions would contain less solute than what is given on the graph, and supersaturated solutions would contain more solute than what is stated.

Gases actually do the opposite thing to solids! Gases become less soluble as temperature increases. This is also shown by solubility curves:



2.4 pH

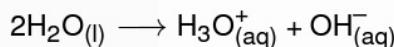
2.4.1 What are acids and bases?

In Year 11 chemistry, we look at the **Arrhenius model** for acids and bases. However, in Year 12 you will look at a different model.

KEY POINT :

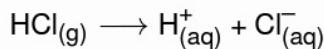
The Arrhenius model defines an **acid** as a substance that dissociates in water to produce hydrogen ions (H^+). A **base** is defined as a substance that dissociates in water to produce hydroxide ions (OH^-).

To a very small degree, pure water ionises itself to produce hydronium ions (H_3O^+) and hydroxide ions (OH^-). This is called the **self-ionisation of water** and the chemical equation is shown below:

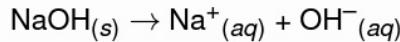


This means that all aqueous solutions have both hydronium and hydroxide ions. However, as the amount of one of these ions increases, the amount of the other decreases. Therefore, since acid dissociates in water to produce hydrogen ions, the concentration of hydronium ions increases and the concentration of hydroxide ions decreases.

Some examples of acids include hydrochloric acid (HCl), sulfuric acid (H_2SO_4), and nitric acid (HNO_3). When dissolved in water, all these compounds produce hydrogen ions. For example, the chemical equation for hydrochloric acid gas being bubbled through water is given below:



Some bases include sodium hydroxide (NaOH), ammonia (NH_3), and calcium hydroxide ($\text{Ca}(\text{OH})_2$). When these compounds dissolve in water, hydroxide ions are produced. For example, the chemical equation for solid sodium hydroxide dissolving in water is given below:



2.4.2 pH

pH is a measure of the concentration of hydrogen ions in solution, which ranges from 0 to 14. pH of acids is between 0 and 7 and pH of bases is between 7 and 14. Neutral solutions have a pH of 7.

$$\text{pH} = -\log_{10} (\text{concentration of } \text{H}_3\text{O}^+)$$

The equation above can be re-written as: $[\text{H}_3\text{O}^+] = 10^{-\text{pH}}$. Note that square brackets are often used to represent concentration.

Indicators can be used to measure the pH of a solution, as acids and bases change colour when certain chemicals are present. For example, acids turn litmus paper red, whereas bases turn litmus paper blue. Universal indicator is commonly used in experiments because it is a mixture of indicators so it can change into many colours.

Strong acids have a lower pH than **weak acids**. This is because strong acids fully dissociate in water to form H^+ , whereas weak acids only partially dissociate.

Similarly, the pH values of **strong bases** are higher than **weak bases**. This is because strong bases fully dissociate in water to form OH^- ions, whereas weak bases only partially dissociate. Make sure you don't get the strength of acids and bases confused with concentration!

2.4.3 Diluting acids and bases

Acids and bases can be diluted by adding water to them. To calculate the effect of adding water on the concentration of acids and bases the following equation can be used:

$$C_1 V_1 = C_2 V_2$$

where C_1 and V_1 are the initial concentration (M) and volume (L) of acid/base, and C_2 and V_2 are the final concentration and volume of the solution after water has been added.

Note that this equation can be used to dilute other solutions (not just acids and bases).

SAMPLE :

Let's calculate how much water is required to dilute 15 mL of 3 M NaOH solution to 2 M:

$$C_1 = 3 \text{ M}$$

$$V_1 = 0.015 \text{ L}$$

$$C_2 = 2 \text{ M}$$

$$C_1 V_1 = C_2 V_2$$

$$3 \times 0.015 = 2 \times V_2$$

$$V_2 = \frac{3 \times 0.015}{2}$$

$$= 0.0225 \text{ L}$$

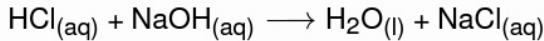
So, the final volume of the diluted base is 22.5 mL. Therefore, 7.5 mL of water needs to be added to dilute the base.

2.5 Reaction of acids

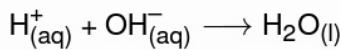
Now that we know what an acid is, let's look at how it reacts with different substances:

1. Acid + reactive metal \rightarrow salt + hydrogen gas
 - Reactive metals include magnesium, calcium, zinc, and iron but not silver, copper or gold.
 - e.g. $2\text{HCl}_{(\text{aq})} + \text{Mg}_{(\text{s})} \rightarrow \text{MgCl}_{2(\text{aq})} + \text{H}_{2(\text{g})}$
2. Acid + metal hydroxide (base) \rightarrow salt + water
 - We've already covered this reaction. It is sometimes called a neutralisation reaction because the acid and base are neutralised, as the H^+ and OH^- react to form water.
 - e.g. $\text{H}_2\text{SO}_{4(\text{aq})} + 2\text{NaOH}_{(\text{aq})} \rightarrow \text{Na}_2\text{SO}_{4(\text{aq})} + 2\text{H}_2\text{O}_{(\text{l})}$
3. Acid + metal carbonate \rightarrow salt + water + carbon dioxide
 - The carbonate anion can be found the formula book (CO_3^{2-}).
 - e.g. $\text{CaCO}_{3(\text{aq})} + 2\text{HCl}_{(\text{aq})} \rightarrow \text{CO}_{2(\text{g})} + \text{H}_2\text{O}_{(\text{l})} + \text{CaCl}_{2(\text{aq})}$

The reactions above can also be written as ionic equations. For example, the neutralisation reaction:



can be written as:



The sodium and chloride ions are omitted from the ionic equation because they are spectator ions.

Section 3

Rates of chemical reactions

3.1 Rates of reaction

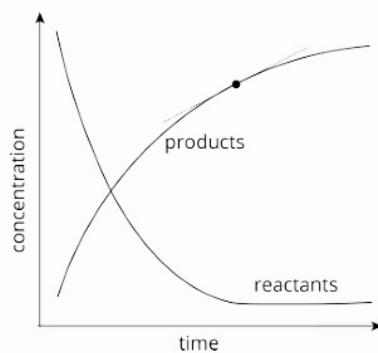
3.1.1 Factors that affect rates of reaction

The **rate of reaction** is how fast a reaction occurs. The higher the rate of reaction, the quicker the reaction will be. Rate of reaction is affected by many factors: concentration, temperature, surface area, pressure, and the presence of a catalyst. I've summarised how these factors affect reaction rate and I've included an explanation of why this occurs using collision theory in the table below. **Collision theory** states that for reactions to occur, the reactants need to collide to form products. The reactants need to collide with the correct orientation and sufficient energy to break the reactants' bonds. The higher the frequency of successful collisions, the higher the reaction rate, as products are formed more quickly.

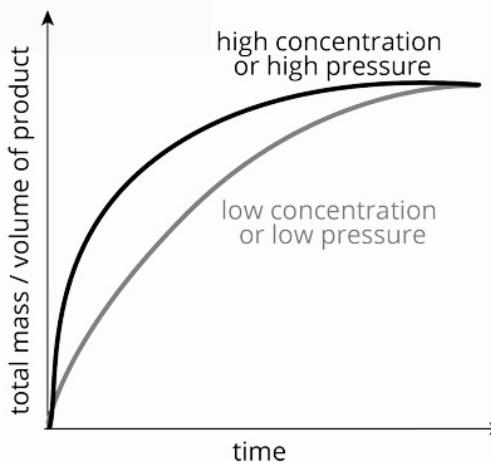
	Effect on rate of reaction	Explanation using collision theory
Increasing concentration	Increase	This increases the number of particles per unit volume. Therefore, collisions between reactants are more probable, so products are formed quicker.
Increasing temperature	Increase	This increases the average kinetic energy and hence the average speed of the particles. So, there are more collisions and products are formed more quickly.
Increasing surface area	Increase	When the reactants are solids, only the particles on the surface react. Therefore, if more particles can react, then more collisions occur, and products are formed quicker.
Increasing pressure	Increase	When pressure is increased by either adding more gas or decreasing volume, the concentration of the reactants increases. This increases the frequency of collisions, causing products to form quicker. Note that pressure only affects reaction rate if the reactants are gases.
Presence of a catalyst	Increase	Catalysts decrease the activation energy required for a reaction to take place by providing an alternative reaction pathway. Activation energy is the minimum energy that a collision between reactants must have. By decreasing the activation energy, catalysts increase the proportion of successful collisions, as more reactant particles have sufficient energy to react. Therefore, products are formed quicker. Note that catalysts don't appear in chemical equations as reactants or products, as they are not broken down in reactions. Catalysts include metal nanoparticles and enzymes.

3.1.2 Graphical representations of rate of reaction

The relationships described in the table on the previous page can also be shown graphically. The graph on the right shows how the concentration of reactants decreases, and the concentration of products increases as a reaction progresses. The rate of reaction at any time during the reaction is given by the slope of the graph for the products. This is the **instantaneous rate of reaction** because it occurs at a particular time.

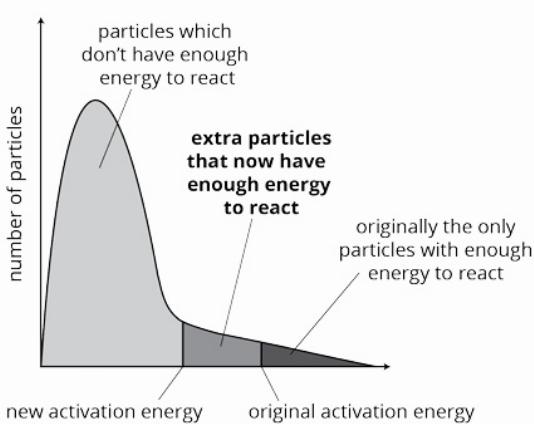
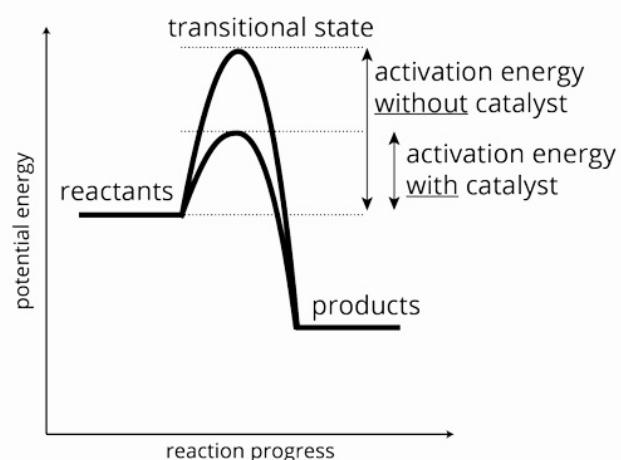


Similarly, the graph below shows how a high concentration and high pressure increase reaction rate. This is shown because the slope (which shows reaction rate) is steeper for the reaction with high concentration or high pressure. This is also shown as this reaction reaches the final mass of products before the other reaction does.



It's important to recognise that catalysts lower the activation energy, but they don't affect the enthalpy change of the reaction. This is shown in the energy diagram on the right.

The **transitional state** is the arrangement of atoms when the activation energy is absorbed. At this stage, bonds are both breaking and forming. Therefore, this arrangement of atoms is unstable.



The effect of catalysts can be also shown by a Maxwell Boltzmann distribution such as the one on the left. This shows that with catalysts present, the activation energy lowers. Therefore, there are more particles with sufficient energy to react than there are when no catalyst is present.

Part III

Advice for studying Chemistry

Assignment advice

This year, you will likely complete two assignments: a student experiment and a research investigation. You will probably have a choice of what experiment or research you complete, so try to make it as simple for yourself as possible. Don't overcomplicate it, and definitely don't choose research articles that you don't understand. Choose an experiment that is easy to conduct and comprehensible.

My main advice is to **read the ISMG** (instrument-specific marking guide, available in the QCAA Chemistry syllabus document on QCAA's website) and make sure that you **understand all the cognitive verbs** in the criteria. This way you will know what you are being marked on. I recommend that after you complete your assignment that you go back through it and check that you have covered every criterion on the ISMG.

To help you understand the ISMG, **look at the exemplars on the QCAA website**. After you have received your marked assignment from your teacher, record any mistakes that you have made, so next year you can improve on them. If you don't understand where you went wrong, talk to your teachers about it, because it's really important that you fully understand how you can do better in Year 12.

Studying tips

My main tip for studying is to **keep up with the class work**. I recommend that you try to make notes and do textbook questions on the topics that you learn in Chemistry every week. This way you can make sure that you understand each topic you complete in class and you can ask your teachers any questions you may have about the content. This way when you come to studying for your exam at the end of the year, you will already have a good understanding of chemistry and good notes as well. It's very important to plan ahead, so try to have a **study timetable** and **set aside a certain amount of time each week for Chemistry**.

When you make your notes, try to make them as concise as possible. Highlight key words and have good headings. This will help you memorise the information.

Doing lots of **practice questions** is extremely important, to make sure that you fully understand all the topics. When you do practice questions closer to the exam, try to not use your notes or textbook. This way you can notice which topics you don't know as well, so you can focus on these a bit more.

It's also important that you **memorise the key definitions and ideas in Chemistry**, so it's a good idea to get a friend or a family member to test you on your notes.

Exam advice

In the exam, the most important thing is to stay calm. Make sure you **highlight the key words in the question**. It's really annoying when you get a question wrong, not because you didn't know the content, but because you didn't read the question properly. In particular, **highlight any units** in the questions, because some questions will try to trick you by including values in units that need to be converted before using them in a formula. Make sure that you know the units for each variable in the formulas. These are mainly SI units, which include kg, m, L, and K. However, in mole calculations, grams are used.

Finally, before you go into the exam, **have a strategy**. Make sure that you think about how you are going to use perusal time. For example, some people like to read through the whole exam and mentally note questions that look harder and may take them more time. It's important that you are aware of how much time you have, and make sure that you don't get stuck on questions. I recommend leaving any questions that you find really hard until the end. This way you don't waste time on them, when you could be getting easy marks for questions that you do know how to do.

I hope you all do well on your exam and have enjoyed Chemistry this year!