ATAR Notes

CHEMISTRY

COMPLETE COURSE NOTES

NSW Year 12
2019-2020

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Preface

Welcome to the Year 12 Chemistry curriculum. Before jumping straight into the content of the course, I thought it would be a good idea to look more broadly at this new and improved syllabus, at the structure of HSC Chemistry, and at Year 12 generally.

I completed Year 12 in 2014, so when I did Chemistry, it was a completely difficult curriculum! One of the main complaints that Year 12 students have is the lack of resources available for the new syllabus. Hopefully these notes will help minimise that issue. However, another complaint that you'll probably have throughout the year is just how damn hard the Chemistry course is!

If you think this course is difficult, you are correct. In the old curriculum, students focused on the impact of chemistry-related concepts on society and the environment. In the new curriculum, you focus on actual chemistry! Personally, I think this change is for the best, but it definitely makes your life a bit more difficult.

So, when you're struggling to understand certain aspects of the Year 12 syllabus, try not to stress. Everyone else is struggling as well, because the content you're learning is objectively difficult. Whilst I didn't do all of the new Year 12 curriculum when I was in high school, I did do a lot of it at university. The content you're learning is *university level!*

There are two different skills you should be learning in Year 12 Chemistry. The first skill is obvious: an appreciation for the chemical sciences, and an ability to apply your learning to real-life problems. The second skill is less obvious: an ability to study difficult content, learn it, and use the content to answer questions.

The second skill is, in many ways, more important that the first. Obviously, it's important that you understand Chemistry. However, if you begin to hone your study skills and your ability to quickly and efficiently process difficult content, you'll find it much easier to improve your marks for Chemistry, as well as all of your other subjects! By improving these skills now, you're setting yourself up to have a much better Year 12 experience.

At the end of these notes, I've discussed some more specific study techniques, as well as general considerations you should be keeping in mind whilst studying Year 12 Chemistry. Make sure to check them out!

Finally, I want to wish you the best of luck with your study of Year 12 Chemistry. As I've said, this is a fairly tough subject. And everyone at ATAR Notes wants to make sure you have all the tools at your disposal necessary to best learn the content. These notes are just the beginning; we run free lectures throughout the year, and if you head over to atarnotes.com you can access all kinds of free help and resources.

Good luck!

— Jake Silove

Part I

Module 5: Equilibrium and Acid Reactions

Topic 1

Static and Dynamic Equilibrium

SYLLABUS :
Inquiry question: What happens when chemical reactions do not go through to completion?

1.1 Reversible and irreversible reactions

SYLLABUS :
Conduct practical investigations to analyse the reversibility of chemical reactions, for example:

- Cobalt(II) chloride hydrated and dehydrated
- Iron(III) nitrate and potassium thiocyanate
- Burning magnesium
- Burning steel wool

In the Year 11 course, you will recall that you studied many, many chemical equations and reactions. They would have all taken the form:

$$A + B \longrightarrow C + D$$

In words, the above equation states that when substance A reacts with substance B (the **reactants**), substance C and D are created (the **products**). You know how to **balance** these chemical equations, you know that you need to include the states of each chemical, and you know how to use them to work out the final product of a given reaction. Here, though, we want to investigate a seemingly unimportant part of the equation; the arrow. Every equation you looked at in the preliminary course would have used the arrow ⟶. However, this arrow doesn't actually represent every reaction. Instead, it represents **irreversible** reactions. In other words, we use a regular left-to-right arrow when the reactants combine to form the products, and the products are unable to turn back into the reactants!

Let's contrast this with other types of chemical reactions, called **reversible** chemical reactions. In these cases, the reactants combine to form the products, but it is also possible for the products to turn back into the reactants! The arrow we use for these reactions looks like: ⇋.

KEY POINT :
⟶ should be used to represent **irreversible** chemical reactions.

⇋ should be used to represent **reversible** chemical reactions.

In this course, you will learn to identify whether a particular equation is reversible or irreversible. One of the interesting aspects of reversible reactions is that, when you change the conditions of the reactions, you will get a different proportion of reactions to products.

Let's look at some examples of reversible chemical reactions, as per the syllabus dot points.

Cobalt(II) chloride hydrated and dehydrated

Hydration or dehydration reactions are exactly what they sound like; the addition, or removal, of water molecules (H_2O_l). The hydration, and subsequent dehydration of cobalt(II) chloride is an interesting example of reversible reactions, due to the physically observable difference between the states.

Cobalt(II) chloride has the chemical equation $CoCl_{2(s)}$. However, upon hydration, it becomes cobalt(II) chloride hexahydrate. This sounds complicated, but hexahydrate just means 'hydrated by six water molecules'. The chemical formula for the hydration of a molecule is simple; we put a dot after the chemical molecule, and just write out however many water molecules have been used for hydration.

Thus, our reversible chemical reaction will be:

$$CoCl_{2(aq)} + 6H_2O_{(l)} \leftrightharpoons CoCl_2 \cdot 6H_2O_{(l)}$$

You will likely do this experiment in the lab, as it is quite a pretty one. Whilst the hexahydrate (hydrated) form is purple, the anhydrous (dehydrated) form is a sky blue colour. So, by adding water to cobalt(II) chloride, the substance will turn from sky blue to purple. As the substance dries (i.e. during dehydration) the colour will turn back to sky blue! We can physically observe the reversibility of some chemical reactions.

KEY POINT :
Anhydrous means substances that do not contain water molecules.

—hydrate means substances that contain some number of water molecules. The prefix determines how many there are (e.g. hexahydrate means six!)

Due to the obvious colour change, this substance can be used to test for the presence of water. Can you think about an example of where testing for the presence of water may be beneficial for chemists?

Iron(III) nitrate and potassium thiocyanate

The relevant ionic equation for this reaction is:

$$Fe^{+3}_{(aq)} + NCS^{-}_{(aq)} \leftrightharpoons FeNCS^{+2}_{(aq)}$$

For this experiment, the iron ions are yellow in solution, and the potassium thiocyanate is clear. However, upon reaction, the two substances form the blood red iron thiocyanate. As above, this can be a useful detection method, and is interesting from a reversible reaction perspective!

Burning magnesium

Not all chemical changes are reversible. Have you ever tried un-scrambling an egg? No, because that would be completely insane. How about turning combusted material back into the fuel that it began as? Well, if we could do that, we would have an infinitely-renewable source of energy, which could prevent the effects of climate change!

Reactions such as rusting and combustion are not reversible; they are irreversible. As a matter of technicality, I suppose it is possible to try and collect all the products of a reaction, pump lots of energy into the system and try to make them re-combine... just like it may be technically possible to pick apart a scrambled egg and put everything back into its unscrambled form. But to think like that would just be nonsense. So, given how difficult that process is, it just makes sense to call it an **irreversible reaction.**

The burning of magnesium is an irreversible reaction. We use the regular arrow to represent this:

$$2Mg_{(s)} + O_2 \longrightarrow 2MgO_{(s)} + \text{ energy}$$

If you want, you can think about the irreversibility of reaction by the amount of energy this reaction releases. Many reactions (we can them exothermic) release energy. By putting the same amount of energy back into the system, we can sometimes reverse this reaction. However, in combustion reactions, so much energy is released that it would be impractical to try and reverse the reaction.

KEY POINT :
Combustion reactions are irreversible, and so require the use of $\longrightarrow$.

Burning steel wool

The burning of steel wool is irreversible, as the above explanation elaborates on. Steel wool is effectively iron, and so the equation is:

$$2Fe_{(s)} + \frac{3}{2}O_{2(g)} \longrightarrow Fe_2O_{3(s)}$$

If you don't quite understand how we are allowed to use $\frac{3}{2}$ in our equation, go back to the Year 11 course and recall that the number in front of each substance is merely a representation of the number of moles necessary to cause the reaction to move forward. $\frac{3}{2}$ is just the number of moles of oxygen necessary to burn 2 moles of iron!

1.2 Types of equilibrium, and open and closed systems

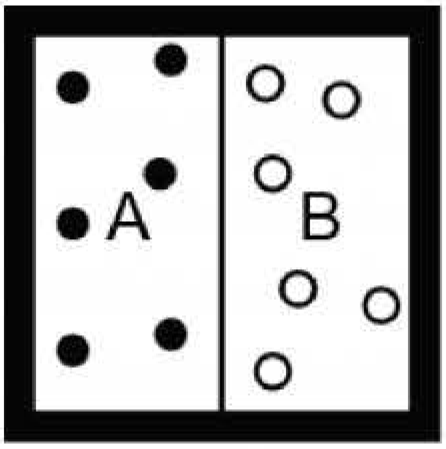
SYLLABUS :
Model static and dynamic equilibrium and analyse the differences between open and closed system.

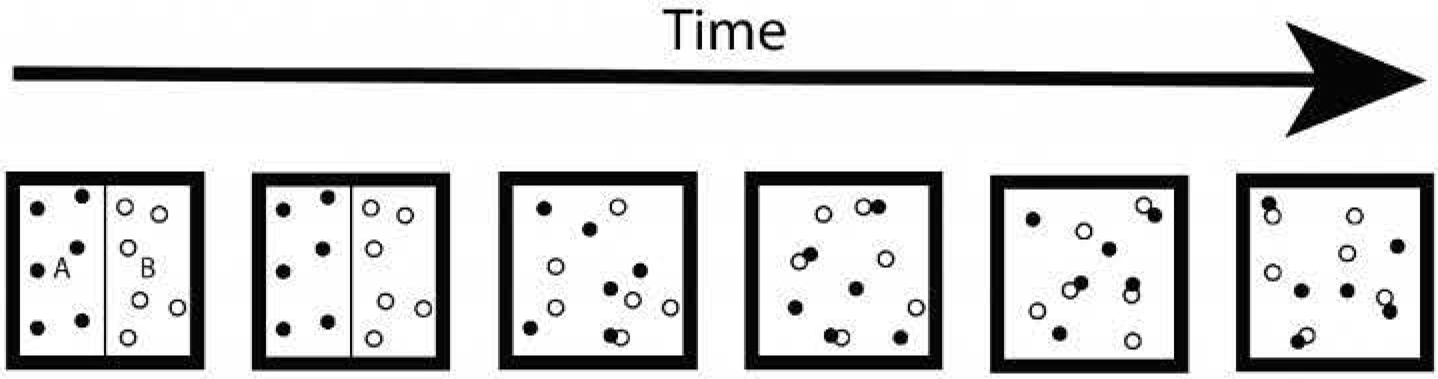
Reversible reactions, over time, will reach a state called an **equilibrium** state. This is the point at which the concentration of **products** and the concentration of **reactants** stays **approximately equal.** To explain this, let's look at a quick example:

$$A_{(g)} + B_{(g)} \leftrightharpoons C_{(g)}$$

● + ○ ⇋ ●○

Let's start with a vessel with a wall down the middle separating gases A and B.
At this stage, we can say that the system is in equilibrium. This is because the 'A' molecules do not change over time, and nor do 'B' molecules. If you looked at this system again in 5 minutes, you would see the exact same thing!
Now, let's remove the wall and see what happens.





At what point does this system reach equilibrium? Between the first two frames, nothing really happens. However, the gas hasn't had time to expand. We see this expansion in the third frame, including a single reaction between A and B to form the molecule C. By the fourth frame, we have three molecules of C. This concentration stays constant in frame five and six. So, equilibrium is reached in frame four!

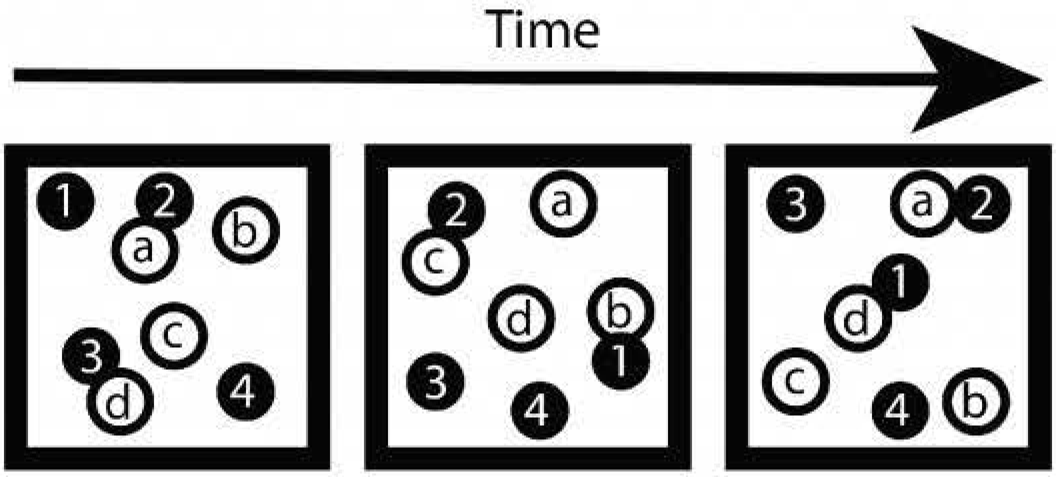
This is a pretty basic example of an otherwise complex notion. Hopefully it has given you a good idea of what an equilibrium is; over time, a system will shift, until it reaches some stable state. In future dot points, we'll find out how to calculate how long this will take!

1.2.1 Dynamic equilibrium

Dynamic equilibrium will be reached **when the rate of the forward reaction is equal to the rate of backwards reaction.** Recall that the point of a reversible reaction is that the reactants can form products, and the products can form reactants. When an equilibrium is reached, the concentration of each reactant and product stays the same. However, the actual particles do not stay stable. In a dynamic equilibrium, they still shift back and forward, but at a rate that is equal.

Let's look at a slightly simplified version of the previous example. This time, let's label each atom. The system has reached equilibrium, and so the concentration of A, B, and C will stay the same. However, the forward reaction ($A_{(g)} + B_{(g)} \longrightarrow C_{(g)}$) will still occur, at the exact same rate as the backwards reaction ($C_{(g)} \longrightarrow A_{(g)} + B_{(g)}$) does! So, we observe the atoms 'switching places'.

The 'A' molecules are labelled 1, 2, 3, and 4, and the 'B' molecules are a, b, c, and d.



Note that the concentration of reacted and unreacted molecules remains constant, but the actual particles themselves are oscillating backwards and forwards! This is what is known as a **dynamic equilibrium.**

1.2.2 Static equilibrium

A static equilibrium, on the other hand, is an equilibrium in which all particles are at rest. Unlike dynamic equilibriums, they do not oscillate between reacted and unreacted states; they remain constant over time.

An example is a normal, '⟶' reaction (irreversible) in which the products cannot turn back into the reactants. As such, the reaction will 'go to completion', with no rate of backwards reaction, until all reactants have turned into products.

KEY POINT :

Dynamic equilibrium is where rates of forwards and backwards reactions are **equal.**

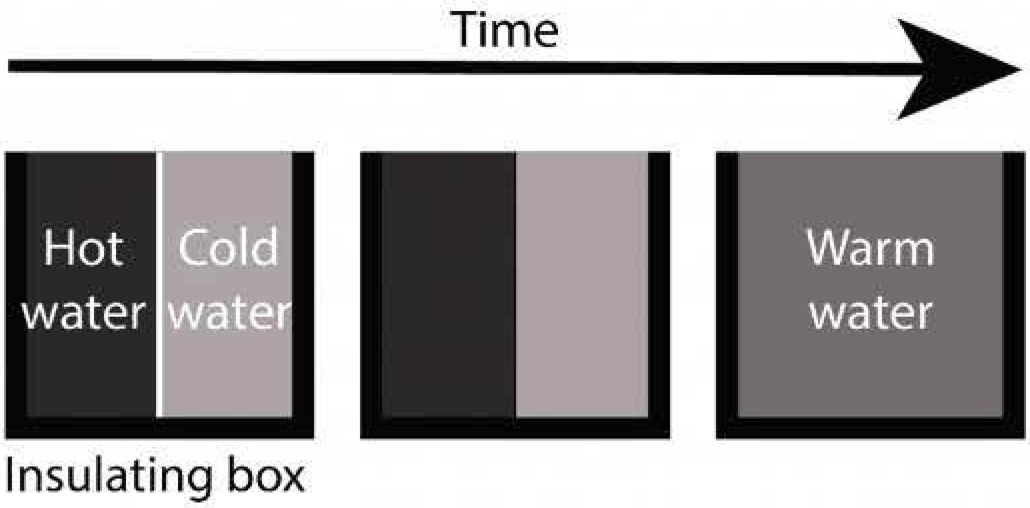
Static equilibrium is where rates of forwards and backwards reactions are **zero.**

Because of these definitions, you can think of static equilibriums as a specific *type* of dynamic equilibrium!

1.2.3 Open and closed systems

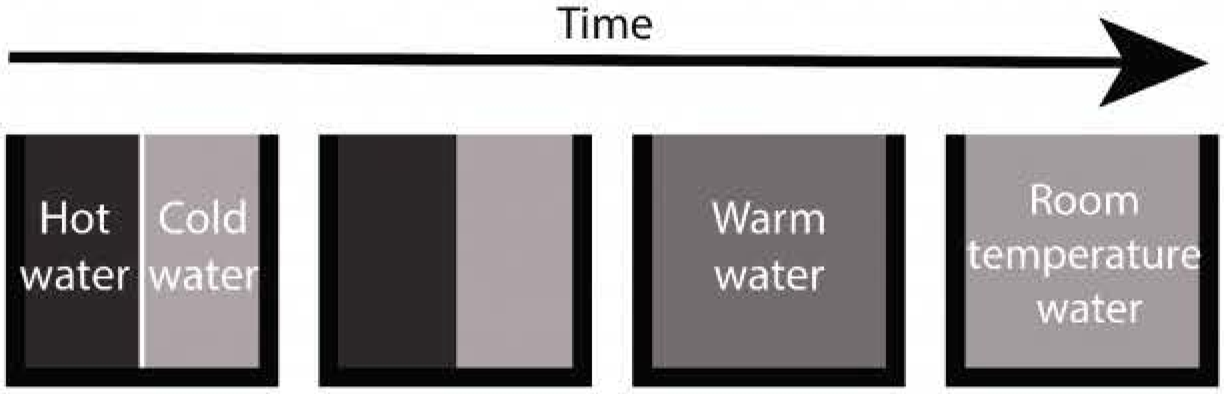
Open systems are systems in which the reaction is able to interact with its surrounding. Closed systems are systems in which the reaction in unable to interact with its surroundings. So the above example ($A_{(g)} + B_{(g)} \leftrightharpoons C_{(g)}$) when confined to a box is a closed system!

Where a system is closed, an equilibrium will be reached by reference to the reactants contained within that system. However, when the system is open, the external environment (e.g. its temperature, gas composition, pressure) will have an impact on equilibrium being reached. The most obvious example here is combining hot and cold water.



In an insulating container, hot and cold water will combine to form warm water. The insulating box serves to 'close' the system; there is no way for the temperature of the water to be affected by the external, ambient temperature.

Now, let's see what happens if the system becomes 'open'. We can represent this as 'opening the lid' of the box.



Now that the box is open, the equilibrium reached is dependent on the external environment! Eventually, the water will reach room temperature.

KEY POINT :

Open systems: are able to interact with the surrounding environment.

Closed systems: are unable to interact with the surrounding environment.

1.3 Non-equilibrium systems

SYLLABUS :

Analyse examples of non-equilibrium systems in terms of the effect of entropy and enthalpy, for example:

- Combustion reactions
- Photosynthesis

We've already described systems in equilibrium (reversible reactions), and the fact that only some chemical reactions are reversible. **How can we explain why some reactions are reversible, and others are irreversible?**

Theoretically, in closed systems, all reactions are reversible. If you just 'put back' what you 'got out' of the reactions, you should be able to re-form the reactant. However, in practice, the vast majority of chemical reactions are done in open systems. This just means the reaction is done in the open air, allowing material and energy to escape.

A reaction becomes irreversible when a product is **irreversibly lost to the environment.** So, where a substance is burned, a large amount of energy is produced (in the form of heat and light). This energy is lost to the environment and cannot be recovered.

You would have learned about **entropy** and **enthalpy** in Year 11 – most combustion reactions result in an increase of entropy, although this is not always the case. However, where entropy has been increased, there is no way to decrease the entropy again unless there is another (greater) increase in entropy elsewhere.

That may sound confusing. Basically, the only way to 'recover' the chemical reactants would be to provide the system with more energy than was lost and increase the total entropy.

This is similar for photosynthesis, which is described by the chemical reaction:

$$6CO_{2(g)} + 6H_2O_{(l)} \longrightarrow C_6H_{12}O_{6(aq)}$$

The process stores chemical energy in the chemical bonds of the product. Work must be done in order to break those bonds and release the energy. Rather than a reversible reaction, this is simply a second reaction which can be used to release energy.

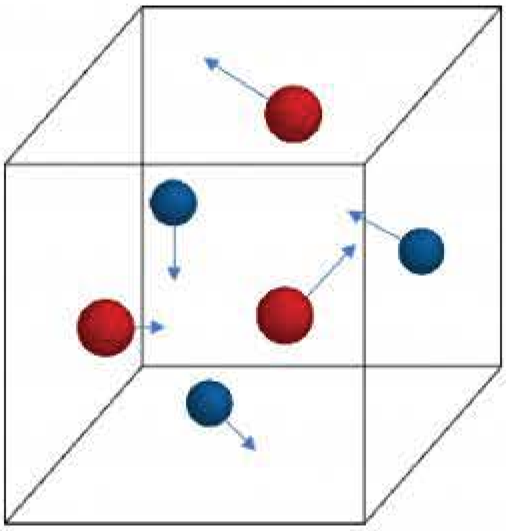
1.4 Collision theory and reaction rate

SYLLABUS :
Investigate the relationship between collision theory and reaction rate in order to analyse chemical equilibrium reactions.

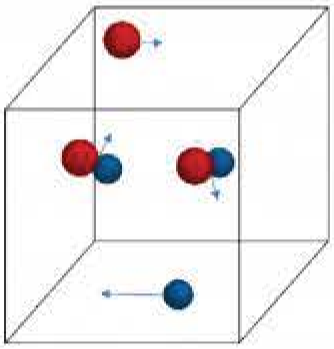
Recall your study of collision theory in the Year 11 course. Collision theory is a model of gas interaction that accounts for the various factors relevant to the rate of reaction.

Collision theory models gases as spherical balls, bouncing around in a vessel. When certain spheres hit other spheres with a certain kinetic energy (and thus, at a certain speed) they react, creating a new compound. Let's imagine a simple gas interaction: $A_{(g)} + B_{(g)} \longrightarrow AB_{(g)}$.

In collision theory, this will look something like the following. Imagine A is the big atom, and B is the smaller atom:



The gases are flying around, in straight lines, and bouncing off the walls. Sometimes, they will bounce into each other. If they have a high enough kinetic energy, and are oriented in the right direction, a reaction may occur! So, at some later point in time, our box could look like the following.



Two reactions have occurred, leaving one molecule of each atom left over. Given enough time, they may react as well, or they may not. There are a few important concepts within reaction theory that we will look at next. First, the molecules must have **sufficient energy.** Secondly, the molecules must actually **collide.** Thirdly, the molecules must collide at the **right orientation.**

1.4.1 Activation energy

For a reaction to occur, the reactants must overcome some initial activation energy. If the particles don't have enough energy, they will not react, even if they bump into each other. We define activation energy as the **minimum energy required** so that the particles, when they collide, with react to form a product. Why is there a minimum energy? When a product forms, chemical bonds within A and B must be broken to form the compound, AB. This requires energy. The higher the energy of the particles (i.e. the faster they are moving), the more likely a reaction is to occur.

1.4.2 Collisions

For a reaction to occur, the gases must actually collide. This collision must have enough energy to overcome the activation chemistry. However, if we want more reactions to occur, we don't just require more kinetic energy. We also need to ensure *heaps* of collisions occur. To do this, we can do one of two things: **increase the concentration** of the reactants (i.e. put more spheres in the box), or **increase the speed** of the reactants. We can increase the speed of gas molecules by heating them up.

1.4.3 Molecular orientation

So far, we have discussed the importance of a collision occurring with sufficient energy. However, even if a collision does occur, the actual reaction may not go forward. Remember, we started by approximating the gases as spheres. However, in real life, they are complex molecules. The reaction might require bonds to break in a particular area of the compound. In that case, the inbound sphere must hit that exact location.

Topic 2

Factors that Affect Equilibrium

SYLLABUS :
Inquiry question: What factors affect equilibrium and how?

2.1 Le Chatelier's principle

SYLLABUS :
Investigate the effects of temperature, concentration, volume and/or pressure on a system at equilibrium and explain how Le Chatelier's principle can be used to predict such effects, for example:

- Heating cobalt(II) chloride hydrate
- Interaction between nitrogen dioxide and dinitrogen tetroxide
- Iron(III) thiocyanate and varying concentration of ions

Le Chatelier's principle is the most important chemical law related to equilibrium. You'll be writing this out a billion times throughout your HSC, so it's worth memorising it now. Le Chatelier's principle states that:

> *"A system at equilibrium, when introduced to a change,*
> *will shift in equilibrium to minimise that change."*

What does that mean? After a reversible reaction has reached equilibrium, if there are no changes to that system, nothing will change over time. However, if you introduce some sort of **change** to that system, **the equilibrium will shift** in order to 'fix' or 'turn back' that change! Let's look at the three major ways in which this occurs. Before we do that, however, we have to remember a few things.

First, recall the definitions for **exothermic** and **endothermic** reactions. Some reactions require extra energy to occur. Others are able to release energy, often in the form of heat. These kinds of reactions are called endothermic and exothermic reactions.

KEY POINT :
Endothermic reactions absorb heat to move forward.

Exothermic reactions release heat to move forward.

When we represent a chemical reaction, you may have seen a value on the right hand side denoted by ΔH^0. This is a measure of the energy released or absorbed. Where $\Delta H^0 > 0$, the reaction is **endothermic** (as the particles are absorbing heat). Where $\Delta H^0 < 0$, the reaction is **exothermic** (as the particles are releasing heat).

Secondly, recall that one **mole** of gas will always occupy the **same volume.** In other words, one mole of carbon dioxide will occupy the same volume as one mole of hydrogen gas. This volume is **dependent on temperature.** At 25°C, one mole of gas occupies approximately 24.79 L of space.

Now, we can investigate the all-important Le Chatelier's principle. Let's use this example:

$$A_{(g)} + B_{(g)} \leftrightharpoons AB_{(g)} \qquad \Delta H^0 = 37\text{kJ}$$

Concentration

First, we investigate what happens when we change the concentration of either the reactants or the products. If we **increase the concentration of the reactants,** then by Le Chatelier's principle, the equilibrium will shift to minimise that change. In other words, the equilibrium will want to **decrease the amount of reactants present,** and thus **shift away from the reactants.** We call this **'shifting to the right,'** as more products are produced. In the previous example, an increase in the concentration of gases A and B will cause a shift towards the gas AB.

If we **increase the concentration of the products,** then by Le Chatelier's principle, the equilibrium will shift to minimise that change. In other words, the equilibrium will want to **decrease the amount of products present,** and this **shift away from the products.** We call this **'shifting to the left',** as more reactants are produced. In the previous example, an increase in the concentration of AB will cause a shift towards the gases A and B.

KEY POINT :

An **increase in reactants** causes a shift to the **right (towards products).**

An **increase in products** causes a shift to the **left (towards reactants).**

Pressure

Next, we investigate what happens when we increase or decrease the pressure of a system. If we **increase the pressure** of a system, then by Le Chatelier's principle, the equilibrium will shift so as to minimise that change. In other words, the equilibrium will want to **lower the pressure.** Remember that each mole of gas occupies the same volume. So, to lower the pressure, the equilibrium will **shift towards the side with fewer moles of gas.** In the previous example, an increase in pressure will cause the equilibrium to shift to the right (as the products contain 1 mole of gas, but the reactants contain 2 moles of gas).

If we **decrease the pressure** of a system, then by Le Chatelier's principle, the equilibrium will shift so as to minimise that change. In other words, the equilibrium will want to **raise the pressure.** Remember that each mole of gas occupies the same volume. So, to raise the pressure, the equilibrium will **shift towards the side with more moles of gas.** In the previous example, a decrease in pressure will cause the equilibrium to shift to the left (as the products contain 1 mole of gas, but the reactants contain 2 moles of gas).

KEY POINT :

An **increase in pressure** causes a shift to the side with **fewer** moles of gas.

An **decrease in pressure** causes a shift to the side with **more** moles of gas.

Temperature

Finally, we investigate what happens when we increase or decrease the temperature of a system, If we **increase the temperature** of a system, then by Le Chatelier's principle, the equilibrium will shift so as to minimise that change. In other words, the equilibrium will want to **lower the temperature.** Remember that exothermic reactions release energy when moving to the right, and endothermic reactions absorb energy when moving to the right. So, to **lower** the temperature, **exothermic** reactions will move to the **left** (towards the reactants), and **endothermic** reactions will move to the **right** (towards the products). In the previous example, we know that the system is endothermic (as it absorbs heat in order to move forward). As such, to lower the temperature the equilibrium will shift to the right (towards the products).

If we **decrease the temperature** of a system, then by Le Chatelier's principle, the equilibrium will shift so as to minimise that change. In other words, the equilibrium will want to **raise the temperature.** Remember that exothermic reactions release energy when moving to the right, and endothermic reactions absorb energy when moving to the right. So, to **raise** the temperature, **exothermic** reactions will move to the **right** (towards the products), and **endothermic** reactions will move to the **left** (towards the reactants). In the previous example, we know that the system is endothermic (as it absorbs heat in order to move forward). As such, to raise the temperature the equilibrium will shift to the left (towards the reactants).

KEY POINT :
An increase in temperature will cause **exothermic** reactions to move to the **left,** and **endothermic** reactions to move to the **right.**

A decrease in temperature will cause **exothermic** reactions to move to the **right,** and **endothermic** reactions to move to the **left.**

Heating cobalt(II) chloride hydrate

Recall from earlier that the equilibrium state of cobalt(II) chloride is:

$$CoCl_{2(aq)} + 6H_2O_{(l)} \leftrightharpoons CoCl_2 \cdot 6H_2O_{(l)}$$

However, it is now important to know whether this process is endothermic or exothermic. We can actually work this one out; we know that warming up the substance will evaporate the water, and so dehydrate the system. Since $CoCl_2 \cdot 6H_2O_{(l)}$ is the hydrated form of cobalt(II) chloride, and an increase in energy will cause a shift to the left, the system must be exothermic (i.e. $\Delta H^0 < 0$).

Now that we know the process is exothermic, we know from Le Chatelier's principle that by increasing the temperature, we will cause a shift to the left (towards the dehydrated reactants). We also know, from a previous dot point, that the dehydrated form is sky blue, whereas the hydrated form is purple. So heating the mixture will make the solution turn blue!

Nitrogen dioxide and dinitrogen tetroxide

An equilibrium between nitrogen dioxide and dinitrogen tetroxide is formed as follows:

$$2NO_{2(g)} \leftrightharpoons N_2O_{4(g)} \qquad \Delta H^0 = -57.20 \text{ kJ/mol}$$

This is a great way to investigate all three applications of Le Chatelier's principle!

First, we look towards the easiest application: concentration. By increasing the concentration of nitrogen dioxide in the system, Le Chatelier's principle says that the equilibrium will shift to minimise that change. So, the equilibrium will 'use up' the nitrogen dioxide by turning into dinitrogen tetroxide! This causes a shift to the right.

Now let's look at pressure. Say we increase the pressure of the vessel in which the reaction is occurring. What will happen? Knowing that gases occupy the same volume per mole, we know that the best way to decrease the pressure (as Le Chatelier's principle tells us we should be doing) is to shift towards the side with the fewest moles of gas. In this case, that is the right hand side, and so the equilibrium will shift towards the products.

Finally, we can investigate temperature. This equilibrium is exothermic, as nitrogen dioxide releases energy when reacting to form dinitrogen tetroxide. So, an increase in temperature will cause the equilibrium to shift to the left; that is, towards the products.

Let's say we wanted to increase the yield of this reaction as much as possible (that is, attempt to produce as much dinitrogen tetroxide as possible). How would we do this? Well, first we would constantly increase the concentration of nitrogen dioxide, and lower the concentration of the products. Next, we would decrease the temperature as much as possible, to cause a shift to the right. Finally, we would increase the pressure as much as is feasible, also causing a shift to the right.

KEY POINT :
In all of this, it is super important to keep economic factors in mind. Although it would benefit the yield, high pressures are expensive to maintain, and dangerous. We must balance safety and economy against chemical principles and yield.

Iron(III) thiocyanate and varying concentration of ions

The equilibrium produced by the mixture of these two substances is as follows:

$$Fe^{+3}_{(aq)} + NCS^{-}_{(aq)} \leftrightharpoons FeNCS^{+2}_{(aq)}$$

Here, you can (in the lab) clearly investigate the effect of La Chatelier's principle. The reactants are yellow or clear, whilst the products are a deep, blood red. By adding various concentrations of reactants and products, you can watch the equilibrium shift backwards and forwards!

2.2 Equilibrium and collision theory

SYLLABUS :
Explain the overall observations about equilibrium in terms of the collision theory.

Le Chatelier's principle can also be explained using collision theory.

- **Concentration:** as you increase the concentration of the reactants, there are more molecules whizzing around the reaction vessel. As such, there is a higher likelihood of two molecules colliding with each other, and thus causing a reaction to take place. This is pretty intuitive; the more molecules, the more reactions!
- **Pressure:** when we increase the pressure of a gas, what we really mean is that the density of gas molecules has increased. Similar to concentration, this results in a higher likelihood of collisions taking place. As a result, the reaction rate will increase, and the equilibrium will shift.
- **Temperature:** an increase in temperature really means that the individual molecules will have more energy, and thus be zooming around more quickly (i.e. with a higher kinetic energy). Where particles interact, they often have to overcome some activation energy in order to react. This increased kinetic energy will increase the likelihood of a reaction being successful upon collision.

2.3 Activation energy and heat

SYLLABUS :
Examine how activation energy and heat of reaction affect the position of equilibrium.

The **activation energy** is the **energy required to break chemical bonds,** and **cause a reaction** to take place. Where the activation energy is high, few molecules will have sufficient energy to cause a reaction to take place. So, increasing kinetic energy (i.e. heating up the system) will help overcome this activation energy, causing a shift towards the product of that reaction.

Topic 3

Calculating the Equilibrium Constant K_{eq}

SYLLABUS :
Inquiry question: How can the position of equilibrium be described and what does the equilibrium constant represent?

3.1 The equilibrium expression

SYLLABUS :
Deduce the equilibrium expression (in terms of K_{eq}) for homogeneous reactions occurring in solution.

Before deducing anything, we need to know what a homogenous reaction is. A **homogenous reaction** is any reaction which occurs in a **single phase,** meaning that all reactants and products are of the **same state** (solid, liquid, or gas). An example of a homogenous reaction might be:

$$2No_{2(g)} \leftrightharpoons N_2O_{4(g)}$$

On the other hand, we have combustion reactions, such as the one below, which contains two phases (solids and gases). Thus, this is a **heterogeneous** reaction.

$$2Fe_{(s)} + \frac{3}{2}O_{2(g)} \longrightarrow Fe_2O_{3(s)}$$

We have already discussed the importance of equilibriums, ways to *qualitatively* describe equilibriums, and methods to predict the change in equilibriums. Now we are ready to develop of **quantitative method** of describing equilibriums!

This is where the equilibrium constant K_{eq} comes in. Chemists need a universally accepted way to define where the equilibrium of a particular reaction sits, and this is that method!

To find the equilibrium constant, we only need to know the concentration of each substance at equilibrium. We multiply the concentration of each product together, and divide that by each reactant multiplied together. Mathematically, this is described by:

$$aA + bB \leftrightharpoons cC + dD$$
$$K_{eq} = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

The upper case letters are all **chemicals,** whereas the lower case letters are the **molar ratios** relevant to the chemical equation. The square brackets are used to remind us that this is the concentration at equilibrium, not just any constant.

Why is each concentration to the power of the number of moles relevant in this reaction? Well, let's look at a very simple system in equilibrium.

$$2A_{(g)} \leftrightharpoons 3B_{(g)}$$

Using our equation, we can write out the equilibrium constant as:

$$K_{eq} = \frac{[B]^3}{[C]^2}$$

However, we could also rewrite our initial equation as the following:

$$A_{(g)} + A_{(g)} \leftrightharpoons B_{(g)} + B_{(g)} + B_{(g)}$$

Now, remember that the whole point of the equilibrium constant is that we multiply the products, and divide by the reactants multiplied together.

$$K_{eq} = \frac{[B][B][B]}{[A][A]} = \frac{[B]^3}{[A]^2} \quad \text{... as expected!}$$

3.2 Using K_{eq}

SYLLABUS :
Perform calculations to find the value of K_{eq} and concentrations of substances within an equilibrium system, and use these values to make predictions on the direction in which a reaction may proceed.

Now, it's time to look at a number of examples to put our new-found equation to use. Whilst it may look complicated, it's actually deceptively simple.

SAMPLE :
Question: Hydrogen and nitrogen can form ammonia in equilibrium. Write out an expression for the equilibrium constant.

To write out the equilibrium constant for any system, we must start by writing out a balanced chemical equation. We know that the equilibrium involves two reactants and a product, so

$$N_{2(g)} + 3H_{2(g)} \leftrightharpoons 2NH_{3(g)}$$

Now, to write out the equilibrium constant, we just use the formula above!

$$K_{eq} = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

SAMPLE :
Question: At equilibrium, the concentration of ammonia is 2×10^{-2} mol L^{-1}, while the concentration of nitrogen is 3.1×10^{-3} mol L^{-1} and the concentration of hydrogen is 1.2×10^{-2}mol L^{-1}. Find the equilibrium constant.

Here, just plug the values given to us into the equation we found above:

$$K_{eq} = \frac{[NH_3]^2}{[N_2][H_2]^3} = \frac{(2 \times 10^{-2})^2}{(3.1 \times 10^{-3})(1.2 \times 10^{-2})^2} = 7.5 \times 10^4 \text{ (correct to two significant figures)}$$

It's worth noting here that when $K_{eq} < 1$, the mixture will contain mostly reactants. However, where $K_{eq} > 1$, the mixture will contain mostly products. In this case, there will be barely any reactants, and a lot of products!

Now, let's look to some more difficult questions. These are far more common than those above, and take quite a few more steps. However, once you've done one or two, you'll be able to tackle anything the HSC can throw at you!

SAMPLE :

Question: The following chemical equation reaches equilibrium in a closed system:

$$\mathbf{ClNO_{2(g)} + NO_{(g)} \leftrightharpoons NO_{2(g)} + ClNO_{(g)}}$$

A 1.0 L glass container was filled with 0.1 moles of $ClNO_{2(g)}$. At equilibrium, there were 0.03 moles of $ClNO_{2(g)}$ and 0.01 moles of $NO_{(g)}$. Calculate the equilibrium constant for this reaction.

To find the equilibrium constant, we need to calculate the concentration of each substance at equilibrium. The equation is going to be:

$$K_{eq} = \frac{[NO_2][ClNo]}{[ClNo_2][NO]}$$

When answering questions like this, it may be useful for you to write out a table. This is the recommended method in university chemistry; see if it works for you!

We have three stages of reaction; the initial state, the change in state, and the final state. So, we could write a table like this:

	$ClNo_{2(g)}$	$NO_{(g)}$	$NO_{2(g)}$	$ClNO_{(g)}$
Initial	0.1 moles	?	0	0
Change	?	?	?	?
Final	0.03 moles	0.01 moles	?	?

Note that the reactants will start at some molar value, and the products will start at zero (as the reaction is yet to take place).

Now, our job is to calculate everything in the 'final' row. To do this, we will need to fill in every cell labelled '?'. Firstly, we note that the moles of $ClNO_2$ has decreased (changed) by 0.1 – 0.01 = 0.07 moles. As the molar ratio of this equation is 1:1:1:1 (i.e. it takes one mole of both reactants to form one mole of each product), we know that this change will be the same for all other substances. However, the sign (i.e. whether the substance increases in concentration or decreases in concentration) will be different.

	$ClNo_{2(g)}$	$NO_{(g)}$	$NO_{2(g)}$	$ClNO_{(g)}$
Initial	0.1 moles	?	0	0
Change	– 0.07 moles	– 0.07 moles	+ 0.07 moles	+ 0.07 moles
Final	0.03 moles	0.01 moles		

Now we are able to fill in the final row:

	$ClNo_{2(g)}$	$NO_{(g)}$	$NO_{2(g)}$	$ClNO_{(g)}$
Initial	0.1 moles	0.08 moles	0	0
Change	– 0.07 moles	– 0.07 moles	+ 0.07 moles	+ 0.07 moles
Final	0.03 moles	0.01 moles	0.07 moles	0.07 moles

Great! So we can now calculate equilibrium constants. However, recall that the square brackets indicate that we are interested in concentration. So, we need to calculate the concentration of the reactants and products at equilibrium. To do so, we employ the formula:

$$C = \frac{n}{V} \quad \text{where the volume is 1 L}$$

	$ClNo_{2(g)}$	$NO_{(g)}$	$NO_{2(g)}$	$ClNO_{(g)}$
Initial	0.1 moles	0.08 moles	0	0
Change	– 0.07 moles	– 0.07 moles	+ 0.07 moles	+ 0.07 moles
Final	0.03 mol / L	0.01 mol / L	0.07 mol / L	0.07 mol / L

$$K_{eq} = \frac{[NO_2][ClNo]}{[ClNo_2][NO]} = \frac{[0.07][0.07]}{[0.03][0.01]} = 16.3$$

SAMPLE :

Question: Hydrogen iodide gas decomposes into hydrogen gas and iodine gas. A 5.0L glass container was filled with 1.8 moles of hydrogen iodide gas. Upon reaching equilibrium, there were 0.55 moles of iodine gas present in the container. Calculate the equilibrium constant.

First, we must write out a balanced chemical equation. This is absolutely crucial to answer equilibrium constant question; the molar ratio determines the nature of the reaction, and factors directly into the equation for the equilibrium constant. Using the information provided:

$$2HI_{(g)} \leftrightharpoons H_{2(g)} + I_{2(g)}$$

Now, we write out our table from an earlier question.

	$HI_{(g)}$	$H_{2(g)}$	$I_{2(g)}$
Initial	1.8 moles	0	0
Change	?	?	?
Final	?	?	0.55 moles

We can quickly work out the change in moles of iodine gas as 0.55 – 0 = 0.55 moles. Given the molar ration, the production of 0.55 moles of iodine gas will also produce 0.55 moles of hydrogen gas, and require the reaction of 2 × 0.55 = 1.1 moles of hydrogen iodide (due to the molar ratio). So:

	$HI_{(g)}$	$H_{2(g)}$	$I_{2(g)}$
Initial	1.8 moles	0	0
Change	– 1.1 moles	0.55 moles	0.55 moles
Final	0.14 mol / L	0.11 mol / L	0.11 moles

So, our equilibrium constant is:

$$K_{eq} = \frac{[H_2][I_2]}{[HI]^2} = \frac{(0.11)(0.11)}{(0.14)^2} = 0.62 \text{ (correct to two significant figures)}$$

SAMPLE :

Question: Consider a mixture of gases in a closed 0.75 L vessel:

Gas	Quantity (mol)
CH_4	2.00
H_2O	0.75
CO	1.25
H_2	0.75

The chemical reaction occurring is described by the equation:

$$CH_{4(g)} + H_2O_{(g)} \leftrightharpoons CO_{(g)} + 3H_{2(g)}$$

If the equilibrium constant is 0.55, is the system at equilibrium? If not, in which direction will the equilibrium move?

First, let's work out whether the system is at equilibrium. We can do this by writing out the equation for the equilibrium constant.

$$K_{eq} = \frac{[H_2]^3[CO]}{[H_2O][CH_4]}$$

Now, the question asks whether the system is currently at equilibrium. So, we work out the concentration of each gas (by dividing by the volume of the vessel) and plug these values into the above equation. Note that I haven't used K_{eq}, as we don't actually know whether the system is at equilibrium!

$$K = \frac{\left[\frac{0.75}{0.75}\right]^3\left[\frac{1.25}{0.75}\right]}{\left[\frac{0.75}{0.75}\right]\left[\frac{2.00}{0.75}\right]} = 0.625$$

As the equilibrium constant given is not equal to the value calculated above, the system is not at equilibrium. Remember that the higher the value of this constant is, the 'more' products there will be (as the value is the products divided by the reactants). Since 0.625 > 0.55, the system will move backwards, towards the reactants, in order to reach equilibrium.

SAMPLE :

Question: An equilibrium is established between hydrogen gas and hydrogen ions in a 5.0 L vessel. Some number of moles of hydrogen gas were put into this vessel, and at equilibrium 1×10^{-2} **moles of hydrogen ions are present. If the equilibrium constant for this system is** 3.3×10^{-4}**, how many moles of hydrogen gas were in the vessel initially?**

At first glance, this question doesn't seem to contain enough information. Whilst we can calculate the change in moles of hydrogen gas, how are we supposed to find the absolute initial molar value? Well, let's just stick with our usual method and see what happens.

First, we write out our chemical equation:

$$H_{2(g)} \leftrightharpoons 2H^+_{(g)}$$

Now, let's write out our usual table (more for consistency than anything else):

	$H_{2(g)}$	$H^+_{(g)}$
Initial	?	0
Change	?	?
Final	?	1×10^{-2} moles

As the products began at a concentration of zero, we can quickly calculate the change in moles. In addition, given the molar ratio is 1:2, we can calculate the change in moles of hydrogen gas.

	$H_{2(g)}$	$H^+_{(g)}$
Initial	?	0
Change	-1×10^{-2} moles	$+1 \times 10^{-2}$ moles
Final	?	1×10^{-2} moles

Now, let's quickly write out the equation for the equilibrium constant of this system. Note that we are really just putting everything we know onto paper to see what comes out. This is generally a good method when it comes to mathematical questions!

$$K_{eq} = \frac{[H^+]^2}{[H_2]}$$

So, what do we do now? Well, we can calculate the final concentration of hydrogen ions ($[H^+] = \frac{1\times10^{-2}}{5.0} = 2 \times 10^{-3}$ moles/L). And, we have the equilibrium constant. So, the only unknown in the above equation is the final concentration of the reactants! We can easily solve for this value:

$$\begin{aligned} K_{eq} &= \frac{[H^+]^2}{[H_2]} \\ 3.3 \times 10^{-4} &= \frac{(2 \times 10^{-3})^2}{[H_2]} \\ [H_2] &= 0.01212 \text{ mol/L} \end{aligned}$$

We know the change in moles for hydrogen gas, but only in terms of moles. So, let's work out the number of moles present in the final equilibrium (as opposed to the concentration).

$$\begin{aligned} C &= \frac{n}{V} \\ n &= 0.01212 \text{ mol/L} \times 5 \\ &= 0.061 \text{ moles} \end{aligned}$$

The change in moles is 1×10^{-2} moles. So, the answer to this question is:

$$0.061 + 1 \times 10^{-2} = 0.071 \text{ moles}$$

SAMPLE :

An equilibrium is reached between nitrogen dioxide and dinitrogen tetroxide in a 2.0 L container. Initially, there are 4.5 moles of nitrogen dioxide and 2.25 moles of dinitrogen tetroxide. After reaching equilibrium, there are A moles of nitrogen dioxide and B moles of nitrogen tetroxide. The equilibrium constant of this system is 1.044. Find the values of A and B.

Again, this question seems complicated, and involves a technique we haven't used so far. Let's perform our usual steps. First, write out out chemical equation:

$$2NO_{2(g)} \rightleftharpoons N_2O_{4(g)}$$

Now, let's use our table technique:

	$NO_{2(g)}$	$N_2O_{4(g)}$
Initial	4.5 moles	2.25 moles
Change	?	?
Final	**A**	**B**

Is this as far as we can go? Not quite. Remember that our molar ratio tells us a lot; for example, it says that for every 2 moles of nitrogen dioxide used, 1 mole of dinitrogen tetroxide will be produced. So, if we define the 'change' in dinitrogen tetroxide as 'x', then we know that the 'change' in nitrogen dioxide will be '$-2x$' (i.e. reduced by twice as many moles). Let's put this into our table:

	$NO_{2(g)}$	$N_2O_{4(g)}$
Initial	4.5 moles	2.25 moles
Change	$-2x$ moles	$+x$ moles
Final	**A**	**B**

Now we are able to replace the values A and B with some other value relative to the initial molar quantity, and the change in molar quantity:

	$NO_{2(g)}$	$N_2O_{4(g)}$
Initial	4.5 moles	2.25 moles
Change	$-2x$ moles	$+x$ moles
Final	$4.5 - 2x$ moles	$2.25 + x$ moles

Both of our final concentrations are now in terms of x, and we are given the equilibrium constant. Let's now try to work with our constant to solve for x (which will give us all the information we need)!

Remember to convert to concentration!

$$K_{eq} = \frac{[N_2O_4]}{[NO_2]^2} = \frac{\left(\frac{2.25+x}{2}\right)}{\left(\frac{4.5-2x}{2}\right)^2} = \frac{\left(\frac{2.25+x}{2}\right)}{(2.25-x)^2}$$

Now, we can use the equilibrium constant provided to us to make this an equation with only one variable to solve.

$$1.044 = \frac{\left(\frac{2.25+x}{2}\right)}{(2.25-x)^2}$$

Now, the syllabus doesn't clearly specify that you *need* to know this, but this methodology is still very useful, so I'll demonstrate it here.

$$1.044(2.25-x)^2 = \frac{2.25+x}{2}$$

$$2.088(2.25^2 - 2 \times 2.25 \times x + x^2) = 2.25 + x$$

$$2.088x^2 - 10.396x + 8.3205 = 0$$

From here, we invoke the quadratic equation, which states that:

$$\text{For: } ax^2 + bx + c = 0 \qquad x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

We can now solve for x :

$$x = \frac{10.396 \pm \sqrt{(-10.396)^2 - 4(2.088)(8.3205)}}{2(2.088)} = 3.98,\ 1.00$$

Now we have two values for x, so how do we decide which one is right? Let's fill in the table with the values we've found:

	$NO_{2(g)}$	$N_2O_{4(g)}$
Initial	4.5 moles	2.25 moles
Change	$-2 \times$ (3.98 or 1.00) moles	+(3.98 or 1.00) moles
Final	$4.5 - 2 \times$ (3.98 or 1.00) moles	2.25 + (3.98 or 1.00) moles

Can you see the issue? If we use the value of x = 3.98 moles, then the final value of nitrogen dioxide will be negative! So, $4.5 - 2 \times 3.96 = -3.46$ moles. You can't have a negative number of moles, so, the 'real' value for x will be x = 1.00 moles, and our table becomes:

	$NO_{2(g)}$	$N_2O_{4(g)}$
Initial	4.5 moles	2.25 moles
Change	-2×1.00 moles	+1.00 moles
Final	$4.5 - 2 \times 1.00$ moles	2.25 + 1.00 moles

And now we've answered the question, and can write with certainty that:

$$A = 4.5 - 2 \times 1 = 2.5 \text{ moles} \qquad B = 2.25 + 1 = 3.25 \text{ moles}$$

That's definitely quite a difficult question, but hopefully you can learn something from it! In particular, attributing an unknown value with the variable x can often be a useful representation.

KEY POINT :
To answer K_{eq} questions:

1. Write out a **balanced chemical equation.**
2. Write out the **general expression** for the K_{eq} of the system.
3. Write out a **table** (remember **ICE:** Initial, Change, Final/Equilibrium)
4. Fill in the table, find all values, and **solve for the desired quantity.**

3.3 Effect of temperature on the value of K_{eq}

SYLLABUS :
Qualitatively analyse the effect of temperature on the value of K_{eq}.

In the same way as we used Le Chatelier's principle to determine the impact a change in the system will have on the equilibrium, so too can we analyse equilibrium in terms of the equilibrium constant. However, whilst change in concentration and pressure will not alter the equilibrium constant itself, temperature will.

KEY POINT :
Only changes in **temperature** will change the value of K_{eq}.

The equilibrium constant will change with temperature, but the direction of change will depend on whether the reaction is exothermic or endothermic. The reasoning behind this is quite tricky, and definitely beyond this course given that the syllabus only asks you to '*qualitatively* analyse.' So, let's just learn the rule!

KEY POINT :
Increasing the temperature in an **exothermic** reaction will **decrease** K_{eq}.
Increasing the temperature in an **endothermic** reaction will **increase** K_{eq}.

SYLLABUS :
Explore the use of K_{eq} for different types of chemical reactions, including but not limited to:

- Dissociation of ionic solutions
- Dissociation of acids and bases

So far, we've focussed on using the equilibrium constant for arbitrary chemical equilibriums. However, we can use the equilibrium constant to describe more generally two kinds of dissociation reactions; the dissociation of ionic solutions, and the dissociation of acids and bases.

Dissociation of ionic solutions

To find the equilibrium constant for the dissociation of ionic solutions, we employ the same formal definition as always. Let's take as an example the dissociation of table salt, $NaCl_{(s)}$. The equilibrium describing this reaction is:

$$\text{NaCl\_(aq)} \leftrightharpoons Na^{+}_{(aq)} + Cl^{-}_{(aq)}$$

In solution, water molecules (due to their polarity) rip apart the two atoms, overcoming the intramolecular force. We can use the equilibrium constant to define the extent to which ionisation occurs. The equation will be:

$$K_{eq} = \frac{[Na^{+}][Cl^{-}]}{[NaCl]}$$

There are plenty of ways to experimentally detect the concentration of ions, including conducting an electrical probe test (this determines the electrical conductivity of a liquid, which will be dependant on the concentration of ions present). Given a set of information, we can quickly calculate the equilibrium constant for this system!

Dissociation of acids and bases

Personally, I think this part of the dot point is placed in a less-than-ideal part of the curriculum. Up until now, you won't have thought too much about acids and bases conceptually. But we'll introduce some concepts that we will go into more detail with later in these notes.

When it comes to dissociation of acids and the constant associated with this process, chemists tend to use different notation. Instead of K_{eq}, the constant reached is known as the **acid constant,** or K_a. However, this is calculated in exactly the same way as the equilibrium constant.

When an acid is placed in solution, dissociation occurs. As we will learn later on in this course, **acids are proton donors.** Recall that a hydrogen atom is just one proton, circled by one electron. So, a hydrogen ion will be a hydrogen atom with one electron removed (i.e. just a proton!). Hence, when acids donate protons, they are losing a hydrogen ion.

This general dissociation equation can be written as:

$$HA + H_2O_{(l)} \leftrightharpoons A^- + H_3O_{(aq)}$$

...where HA is some acid, which donates a proton (i.e. hydrogen ion) to water, to form a hydronium molecule (we will look at this in more depth later on in the course).

Now, we could write the acid dissociation constant as follows:

$$K_a = \frac{[A^-][H_3O]}{[HA][H_2O]}$$

However, we can reduce this formula a fair bit when we consider certain aspects of this particular interaction. Firstly, the concentration of hydronium molecules is really just the concentration of hydrogen ions floating around (which have been donated by the acid upon dissociation). Secondly, the concentration of water will effectively be 1. This is because there is *so* much more water in any soluble system than the acid, or dissociated products. So, without risking inaccuracy, we can write the generally accepted definition:

$$K_a = \frac{[A^-][H^+]}{[HA]}$$

Topic 4

Solution Equilibria

SYLLABUS :
Inquiry question: How does solubility relate to chemical equilibrium?

When we place salt in water, it dissolves. **Dissolving** is a physical change representing the 'pulling apart' of individual molecules or compounds. In effect, this 'pulling apart' has to overcome intermolecular (or intramolecular) forces in order to disperse the molecules throughout the medium. Here, we will discuss the dissolution of ionic compounds in water, however the same principle applies more generally to dissolution.

4.1 Polarity

SYLLABUS :
Describe and analyse the processes involved in the dissolution of ionic compounds in water.

To understand why certain substances dissolve, we need to understand polarity. You will recall the notion of polarity from the preliminary course, but let's do a quick refresher here.

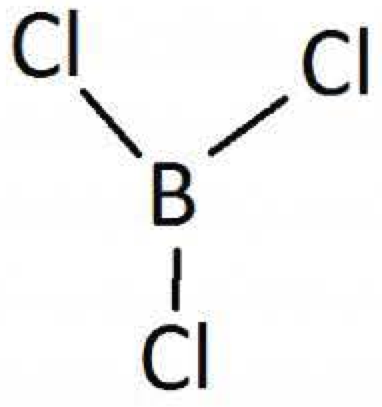
Polarity is essentially a measure of charge distribution. A **polar substance** is one with a **positive end** and a **negative end.** A **non-polar substance** is one that is **neutral everywhere.**

We can use **delta-notation** to show the polarity of a substance, and to explain the kind of bonding that occurs. For instance, fluorine is very electronegative, and hydrogen is not very electronegative (it can be thought of as positive). Thus, we can write the compound hydrofluoric acid as:

$$\delta^{+}\mathrm{H}-\mathrm{F}\;\delta^{-}$$

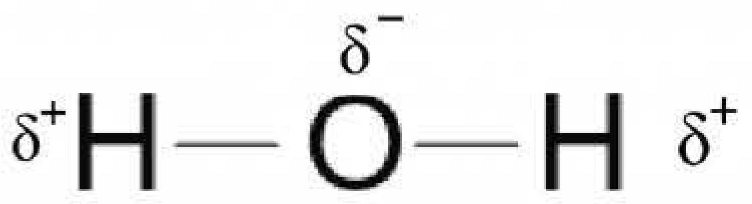
...where δ^- indicates the negative side, and δ^+ indicates the positive side. This substance is clearly polar.

We can use this to model the shape of molecules. We expect charges to **repel;** two negative charges will want to get as far away from each other as possible! For instance, let's look at boron trichloride. Chlorine is very electronegative, and boron is weakly electronegative.



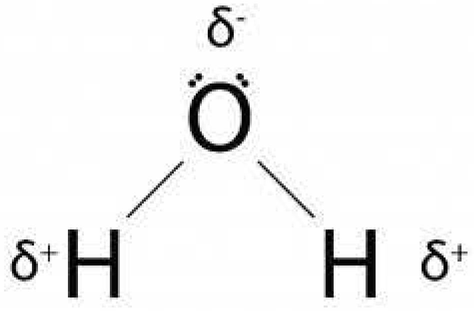
We expect the chlorine atoms to spread out as much as possible, and thus the angle between each atom will be 120°.

Let's apply our understanding of polarity into water molecules, $H_2O_{(l)}$. Hydrogen is positively charged, and oxygen is negatively charged, so you would think that it is 'linear' as follows:



$$\delta^{+}\mathrm{H}-\overset{\delta^{-}}{\mathrm{O}}-\mathrm{H}\;\delta^{+}$$

However, recall that oxygen has six valence electrons. One valence electron is used in the bond between hydrogen and oxygen, and so only two of the six electrons are 'in use'. This means there are two pairs of electrons floating around the oxygen molecule. This will have an affect on the molecular shape!

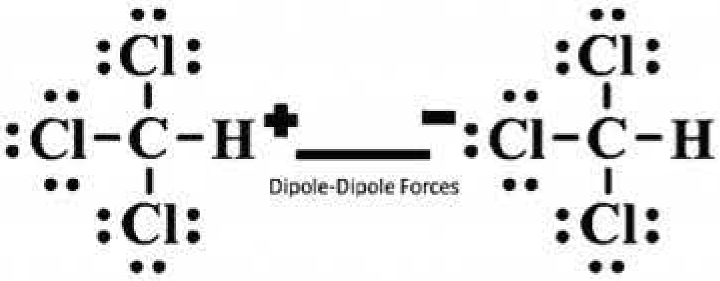


So, water molecules have a charge distribution. Thus, as you probably know, water is polar.

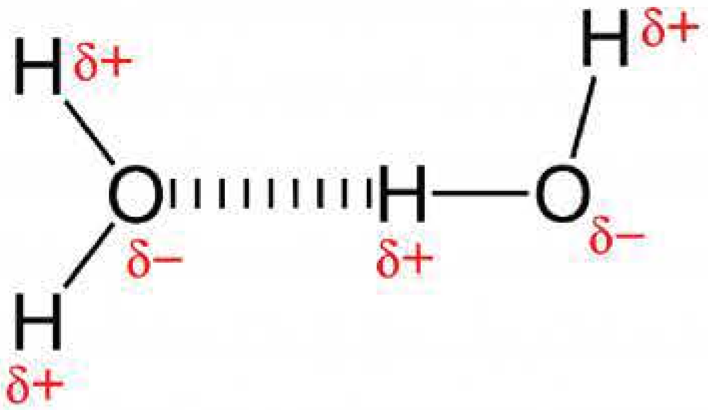
4.2 Types of intermolecular forces

Let's quickly remind ourselves about the **three kinds of intermolecular forces.** Firstly, there are **dispersion intermolecular forces.** These are the hardest to conceptualise, and are also the weakest intermolecular force. As electrons zoom around, at some point there may be more electrons in one place than in another. This creates a net negative charge, which means there must be a net positive charge somewhere else. These are very, very small dipoles (areas of net charge) which cause the molecules to line up such that one negative dipole attracts another positive dipole.

Secondly, there are **dipole-dipole intermolecular forces.** These are stronger than dispersion forces, and occur when polar substances (that is, substances that *always* have a positive dipole and a negative dipole) line up such that a negative dipole attracts a positive dipole, and vice versa. The following is an example of this:



Finally, there are **hydrogen bonds.** These are exactly the same as dipole-dipole intermolecular forces, however are much stronger. They occur when a positive dipole (hydrogen) lines up with extremely electronegative negative dipoles (you can remember this as FON: fluorine, oxygen, or nitrogen!). So, a hydrogen bond between two water molecules looks like this:



How does all of this connect with the dissolution of ionic substances in water? Well, polar ionic substances will be dissolved in water because the positive ions will be drawn towards the negative dipole of a water molecule, and negative ions will be drawn towards the positive dipole of a water molecule!

4.3 Application questions

SYLLABUS :
Investigate the use of solubility equilibria by Aboriginal and Torres Strait Islander Peoples when removing toxicity from foods, for example:

- Toxins in cycad fruit

Solubility equilibria are just like regular equilibria; they are a method of determining how much of a given substance has been dissolved in some substances such as water. For example, for the following equilibrium:

$$NaCl_{(aq)} \leftrightharpoons Na^+_{(aq)} + Cl^-_{(aq)}$$

...our equilibrium constant is:

$$K_{eq} = \frac{[Na^+][Cl^-]}{[NaCl]}$$

The formal idea of chemical equilibria is reasonably recent, however communities have been exploiting these principles for tens of thousands of years. A really interesting example is the cycad fruit, found in Australia, which is highly toxic and carcinogenic.

Cycad fruit was an important food resource for indigenous peoples for at least 10,000 years. However, without a method of reducing toxicity, there is no way that the indigenous peoples could have survived such a prolific use of such a poisonous fruit!

One method used to decrease toxicity was to cut open the fruit and leave the fruit in water. Ideally, the fruit would be left in a running stream or river. This can be explained by solubility equilibria! Say that we placed the fruit in a bowl of water. The equilibrium achieved will be:

$$Toxin_{(s)} \leftrightharpoons Toxin_{(aq)}$$

...where the solid toxin is contained in the fruit, and the aqueous toxin has been dissolved into the water (leaving most of the fruit behind). Over time, an equilibrium will be reached. Removing the fruit from the water would leave behind the aqueous toxins. But that isn't the end of the story!

Let's imagine that we add more water to the bowl, or change the water in the bowl. All of a sudden, the concentration of aqueous toxin will decrease. However, by Le Chatelier's principle, a change in a system will cause the equilibrium to shift in order to minimise that change. So, the system will want to increase the amount of aqueous toxin, and the equilibrium will shift to the right (again, decreasing the toxicity of the fruit!). So, by having a constantly moving/changing water source, the equilibrium will always be shifting to the right. The longer you wait, the less toxic the fruit will be!

SYLLABUS :
Conduct an investigation to determine solubility rules, and predict and analyse the composition of substances when two ionic solutions are mixed, for example:

- Potassium chloride and silver nitrate
- Potassium iodide and lead nitrate
- Sodium sulphate and barium nitrate

For the most part, this dot point is concerned with a practical investigation that you will complete in class. However, there is a shortcut in remembering the solubility rules themselves which you can use in an assessment task!

The trick is to remember the following:

Key Point :

NAGSAG, PMS, and **CASTROBeAR**

- **N**itrates (NO_3^{2-}) are always soluble
- **A**cetates ($C_2H_3O_2^-$) are always soluble
- **G**roup 1 ions are always soluble
- **S**ulphates (SO_4^{2-}) are always soluble
- **A**mmonium (NH_4^+) is always soluble
- **G**roup 7 ions are always soluble

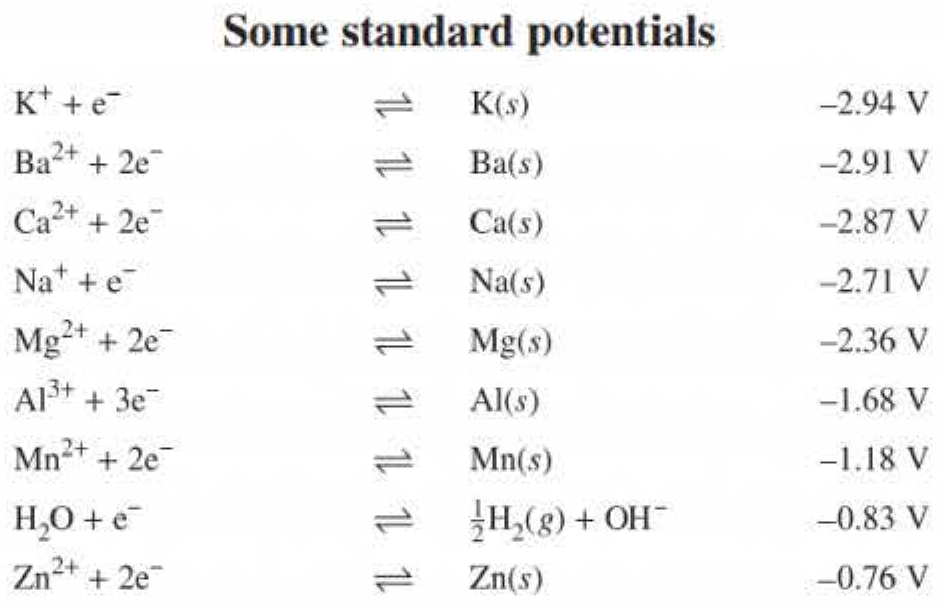
Exceptions:

- **PMS:** **P**b^{2+}, **M**ercury, **S**ilver
- **CASTROBeAR:** **Ca**lcium (Ca^{2+}), **Stro**ntium (St^{2+}), **Ba**rium (Ba^{2+})

From these mnemonic devices, you can answer all solubility questions.

For instance, silver nitrate and silver acetate are both soluble, however silver sulphate and silver chloride are not. First, figure out what the 'general' rule (**NAGSAG**) for a given substance is. Then, check for any exceptions (i.e. **PMS** or **CASTROBeAR**).

Now, let's look to the dot point itself. We need to remember that more active metals will produce less active metals in solution. The activity series is established by the table of standard potentials you have access to:



Some standard potentials

$K^+ + e^-$	$\rightleftharpoons$	$K(s)$	–2.94 V
$Ba^{2+} + 2e^-$	$\rightleftharpoons$	$Ba(s)$	–2.91 V
$Ca^{2+} + 2e^-$	$\rightleftharpoons$	$Ca(s)$	–2.87 V
$Na^+ + e^-$	$\rightleftharpoons$	$Na(s)$	–2.71 V
$Mg^{2+} + 2e^-$	$\rightleftharpoons$	$Mg(s)$	–2.36 V
$Al^{3+} + 3e^-$	$\rightleftharpoons$	$Al(s)$	–1.68 V
$Mn^{2+} + 2e^-$	$\rightleftharpoons$	$Mn(s)$	–1.18 V
$H_2O + e^-$	$\rightleftharpoons$	$\frac{1}{2}H_2(g) + OH^-$	–0.83 V
$Zn^{2+} + 2e^-$	$\rightleftharpoons$	$Zn(s)$	–0.76 V

The higher a substance is in the list, the more reactive it is.

Key Point :

Recall that **more active metals** will replace **less active metals** in solution.

Potassium chloride and silver nitrate

First, let's establish whether potassium chloride and/or lead nitrate are soluble. Chloride is a Group 7 metal, and by NAGSAG, all group 7 metals are soluble. However, the PMS exception applies, which states that potassium, mercury, and silver compounds involving a Group 7 substance are *not* soluble. So, potassium chloride is *not* soluble. From NAGSAG, all nitrates are soluble, so silver nitrate is soluble. So our equation is the addition of a solid (potassium chloride) to an aqueous solution (silver nitrate). Will displacement occur? To work that out, we check whether potassium is more reactive than silver. According to our table of standard potentials, potassium is much higher up than silver (not included on the extracted list above). The rule is therefore that **more active metals will replace less active metals in solution.** Our equilibrium becomes: $KCl_{(s)} + AgNO_{3(aq)} \leftrightharpoons AgCl_{(s)} + KNO_{3(aq)}$.

Sodium sulphate and barium nitrate

As all nitrates are soluble, barium nitrate is of course soluble. Sulphates, by NAGSAG, are soluble. However, the PMS and CASTROBeAR apply. So, sodium sulphate is not soluble. However, barium is higher on the activity series than sodium. So, no displacement will occur, and no equilibrium will be reached; both substances will just sit there, inactive.

4.4 Solubility products

SYLLABUS :
Derive equilibrium expressions for saturated solutions in terms of K_{sp} and calculate the solubility of an ionic substance from its K_{sp} value.

The solubility product is slightly different to the equilibrium constant. Rather than dividing by the concentration of the reactants, we assume that the dissolution percentage is quite small, and so the denominator is approximately equal to 1. So, for the equilibrium: $AB_{(s)} \leftrightharpoons A^-_{(aq)} + B^+_{(aq)}$, we have an equilibrium constant of $K_{eq} = \frac{[A^-][B^+]}{[AB]}$. However, our solubility product will just be $K_{sp} = [A^-][B^+]$.

For solubility equilibria, we can use K_{sp} to calculate the solubility of a given substance. Let's look at $AgCl_{(s)} \leftrightharpoons Ag^+_{(aq)} + Cl^-_{(aq)}$ as an example.

This dissolution has solubility product of 2×10^{-10} at room temperature. So what is the solubility of silver chloride? We can write the solubility product as $K_{sp} = [Ag^+][Cl^-] = 2 \times 10^{-10}$.

How does this help us? Because the reaction produces one mole of each product, we can say that the concentration of both silver and chlorine ions are equal. So:

$$2 \times 10^{-10} = [Ag^+]^2 = [Cl^-]^2$$
$$[Ag^+]^2 = [Cl^-] = \sqrt{2 \times 10^{-10}} \text{ mol/L}$$

We want to derive the **solubility** of this substance. Solubility is calculated in grams per mL, and should be **dependent on the reactant, not the product.**

Let's pretend with have 1 L of water, and we place 1 mol of silver nitrate into it.

We know that $\frac{\sqrt{2 \times 10^{-10}} \text{ mol}}{2} \times 1 \text{ L} = \sqrt{2 \times 10^{-10}}$ mol of silver nitrate will be dissolved, because that is how much of the product is produced. We then calculate the number of grams of reactant has been dissolved:

$$\text{Moles} = \frac{\text{mass}}{\text{molar mass}}$$
$$\text{mass} = \sqrt{2 \times 10^{-10}} \times 143.32$$
$$= 0.002 \text{ g}$$

Solubility of a given substance will be equal to the grams dissolved, divided by the volume of water dissolved into. Therefore:

$$\text{solubility} = \frac{0.002 \text{ g}}{1 \text{ L}} = 0.002 \text{ g/L}$$

Now, if you google the solubility of silver chloride, you'll find that the answer is exactly the same as what we got above!

4.5 Precipitates

SYLLABUS :
Predict the formation of a precipitate given the standard reference values for K_{sp}.

A precipitate will form when the quantity of a soluble substance exceeds a **saturation point.** For instance, using silver chloride as an example from above, if you put more than 0.002 g/L of silver chloride in water, then the excess will remain as a solid!

We can predict whether a precipitate will be formed using the same methodology as above. Let's look at another example: the interaction between potassium iodide and lead (II) nitrate.

First, let's establish whether a displacement reaction will occur at all. By NAGSAG, potassium iodide is insoluble (solid), and lead (II) nitrate is an aqueous solution. Say we place 1.3 g of potassium iodide in 100 mL of 0.11 mol/L lead nitrate solution. At room temperature, lead iodide has a $K_{sp} = 4.4 \times 10^{-9}$. Will a precipitate form? If so, how much precipitate will form?

We start by writing out an overall chemical equation. Note that potassium is more reactive than lead, a displacement reaction will occur. Thus: $2KI_{(s)} + Pb(NO_3)_{2(aq)} \leftrightharpoons PbI_{2(s)} + 2KNO_{3(aq)}$.

First, we need to work out how much lead iodine is produced. We begin by calculating the number of moles of each of the reactants.

$$n_{KI} = \frac{m}{MM} = \frac{1.3}{166.0028} = 0.0078 \text{ moles}$$
$$n_{Pb(NO_3)_2} = CV = 0.11 \times 0.1 = 0.011 \text{ moles}$$

For every 1 mole of lead nitrate used up, 2 moles of potassium iodide will be used up. So, if all 0.011 moles of lead nitrate were to be used up, we would need 0.022 moles of potassium iodide. However, we only have 0.0078 moles! So, 0.0078 moles of potassium iodide will be used up (leaving zero moles), 0.0039 moles of lead nitrate will be used up (leaving 0.0079 moles), 0.0039 moles of lead iodide will be produced, and 0.0078 moles of potassium nitrate will be produced!

The only solid in this final system is the lead iodide, of which there is 0.0039 moles. The solubility equilibrium will be $PbI_{2(s)} \leftrightharpoons Pb^{2+}_{(aq)} + 2I^{-}_{(aq)}$. The relationship for the K_{sp} is $K_{sp} = [Pb^{2+}][I^{-}]^2 = 4.4 \times 10^{-9}$.

The concentration of iodide will be double the concentration of lead (because the molar ratio is 1:2). So:

$$[Pb^{2+}]\left(2[Pb^{2+}]\right)^2 = \left(\frac{1}{2}[I^-]\right)[I^-]^2 = 4.4 \times 10^{-9}$$
$$[Pb^{2+}] = \sqrt[3]{1.1 \times 10^{-9}} \text{ mol/L}$$
$$[I^-] = 2\sqrt[3]{1.1 \times 10^{-9}} \text{ mol/L}$$

So, as above, we know that the moles of lead iodide which may be dissolved into the system is:

$$\sqrt[3]{1.1 \times 10^{-9}} \text{ mol/L} \times 0.1 \text{ L} = 0.000103 \text{ moles}$$

But, we calculated that there would be 0.0039 moles of lead iodide produced! So, to find the number of moles 'left over' as a precipitate after dissolution, we simply subtract the amount dissolved from the total amount:

$$n_{PbI_2 \text{ precipitate}} = 0.0039 - 0.000103 = 0.003797 \text{ moles}$$

Key Point :

To determine if a precipitate forms:

1. Find the **limiting reagent.**
2. Calculate the **solubility.**
3. Determine whether there is a **precipitate remaining.**

Part II

Module 6: Acid/Base Reactions

Topic 1

Properties of Acids and Bases

SYLLABUS :
Inquiry question: What is an acid and what is a base?

Our understanding of acids and bases has changed a lot over the last few hundred years. As we'll see in this module, even the definitions of acids and bases have changed drastically.

1.1 Defining and naming acids and bases

SYLLABUS :
Investigate the correct IUPAC nomenclature and properties of common inorganic acids and bases.

Naming acids and bases can sometimes be tricky, and for the most part, you will just have to rote learn their names. Some important acids and bases are listed below.

Name	Chemical Formula
Hydrofluoric acid	HF
Hydrochloric acid	HCl
Nitrous acid	HNO_2
Nitric acid	HNO_3
Sulphurous acid	H_2SO_3
Sulphuric acid	H_2SO_4
Phosphoric acid	H_3PO_4
Carbonic acid	H_2CO_3
Acetic acid	CH_3COOH
Citric acid	$C_6H_8O_7$
Sodium hydroxide	$NaOH$
Barium hydroxide	$Ba(OH)_2$
Ammonia	NH_3

KEY POINT :
The current definition of **acids** is that they are **proton donors.**

Conversely, **bases** are **proton acceptors.**

Recall that a hydrogen atom is just one proton, circled by one electron. So, a hydrogen ion will be a hydrogen atom with one electron removed (i.e. just a proton!) So, when acids donate protons, they are losing a hydrogen ion.

1.2 Products of acid reactions

SYLLABUS :
Predict the products of acid reactions and write balanced equations to represent:

- Acids and bases
- Acids and carbonates
- Acids and metals

You will regularly be asked to predict the result when mixing particular products, and to write about the relevant chemical equation. From these equations, we will soon be able to calculate expected pH, neutralisation points, and much more!

1.2.1 Acids and bases

Recall that acids are proton donors, and bases are proton acceptors. When mixing together acids and bases, there is a general rule that you can use to predict the outcome.

KEY POINT :
Acid + Base ⟶ Salt + Water

Let's look at some examples. The most common acid-base reaction that you will look at in this course is the **mixture of hydrochloric acid and sodium hydroxide.** We know that this will produce a salt and water:

$$NaOH_{(aq)} + HCl_{(aq)} \longrightarrow \text{Salt} + \text{Water}$$

The acid (hydrochloric) will lose a proton (hydrogen ion), produce a chlorine ion. The base (sodium hydroxide) will gain a proton (hydrogen ion), by ionising to form a sodium ion and a hydroxide ion.

The acid and base equilibrium relevant to these two molecules is:

$$NaOH_{(aq)} \leftrightharpoons Na^{+} + OH^{-}$$
$$HCl_{(aq)} \leftrightharpoons H^{+} + Cl^{-}$$

Together, we can see that the salt produced will be sodium chloride (i.e. table salt).

$$NaOH_{(aq)} + HCl_{(aq)} \longrightarrow H_2O_{(l)} + NaCl_{(aq)}$$

Next, let's look at the mixture of hydrochloric acid with barium hydroxide:

$$2HCl_{(aq)} + Ba(OH)_{2(aq)} \longrightarrow 2H_2O_{(l)} + BaCl_{2(aq)}$$

Notice that barium chloride is aqueous, as per our NAGSAG solubility rules.

1.2.2 Acids and carbonates

We also have a rule which helps to determine the products in an acid and carbonate reaction.

KEY POINT :
Acid + Carbonate ⟶ Salt + Water + Carbon dioxide

An example is the mixture of hydrochloric acid and sodium carbonate. You can usually determine the salt present based on a combination of the anion and cation in the two reactants.

$$2HCl_{(aq)} + Na_2CO_{3(aq)} \longrightarrow 2NaCl_{(aq)} + H_2O_{(l)} + CO_{2(g)}$$

Another example is the mixture of sulphuric acid and copper carbonate:

$$H_2SO_{4(aq)} + CuCO_{3(aq)} \longrightarrow CuSO_{4(aq)} + H_2O_{(l)} + CO_{2(g)}$$

1.2.3 Acids and metals

The rule with respect to the mixture of acids and metals is as follows.

KEY POINT :
Acid + Metal ⟶ Salt + Hydrogen gas

When you add an acid to a metal, you'll notice that it 'fizzles', and lets off a gas. That gas is hydrogen gas, as the above rule suggests!

Let's look at the addition of sulphuric acid to magnesium:

$$H_2SO_{4(aq)} + Mg_{(s)} \longrightarrow MgSO_{4(s)} + H_{2(g)}$$

We can see the release of hydrogen gas, and the production of a metal salt. If you remember the reaction rules above, you'll have no trouble finding the products of particular chemical equations!

1.3 Neutralisation

SYLLABUS :
Investigate applications of neutralisation reactions in everyday life and industrial processes.

Neutralisation reactions can be seen in many parts of our society. It's important that we are able to change (by either increasing or decreasing) the pH of any given substance, depending on what we need the substance to be. For example, neutralisation reactions are important to treat water that is either too basic or too acidic. This might be to balance out the pH of a river, drinking water, or just your backyard pool!

Industrial processes use neutralisation reactions as well. Most obviously, the industrial treatment of drinking water uses several neutralisation stages. This is particular important, as the water will often be made alkaline (basic) for the cleaning process, and must be neutralised (i.e. pH = 7) before it is consumed.

Neutralisation reactions occur naturally in the human body. In the stomach, acids break apart the food you consume in order to release nutrients. However, when those nutrients move into the intestines, an alkaline environment is actually preferred in order to most benefit absorption. The pancreas produces a basic substance to neutralise the stomach acid.

1.3.1 Enthalpy of neutralisation reactions

SYLLABUS :
Conduct a practical investigation to measure the enthalpy of neutralisation.

The important thing to get out of this dot point is that **all neutralisation reactions are, in essence, the same!** *Every* reaction where an acid and a base interact comes down to a basic (no pun intended) relationship:

$$H^+ + OH^- \longrightarrow H_2O_{(l)}$$

This reaction happens to be exothermic, and has an enthalpy change of –57.30 kJ/mol. That means that *every* neutralisation reaction will **release** heat.

KEY POINT :
Neutralisation reactions are always **exothermic.**

1.4 Changing definitions of acids and bases

SYLLABUS :

Explore the changes in definition and models of an acid and a base over time to explain the limitations of each model, including but not limited to:

- Arrhenius' theory
- Brønsted-Lowry theory

The way we have defined acids and bases has changed drastically over the years. The below is a very brief summary of some of these theories, including the one used in the HSC Chemistry syllabus (Brønsted-Lowry theory).

Theory	Description	Limitations
Lavoisier	After experimenting on oxides of non-metals, Lavoisier came to the conclusion that acids were substances that contained oxygen atoms.	Many acids, such as hydrochloric acid, do not contain oxygen.
Davy	Noting that hydrochloric acid did not contain oxygen, whilst still acting as an acid, Davy proposed that acids contained replaceable hydrogen atoms.	This theory didn't really have any substance to it. There was no explanation as to when, or how, the molecules interacted.
Arrhenius	Rather than a substance with replaceable hydrogen atoms, Arrhenius suggested that acids would ionise in water to produce hydrogen ions (H^+), and bases would ionise in water to produce hydroxide ions (OH^-).	The Arrhenius definition only works for acids and bases in aqueous solutions. Some acids and bases operate as such *without* needing to be in solution. Still, this theory is very similar to the modern theories.
Brønsted-Lowry	This theory removed the need for ionisation at all, and thus can explain acids which were acidic without being introduced to water. This definition is that acids are **proton donors** and bases are **proton acceptors.** Recall that a hydrogen ion is a hydrogen atom with one electron removed. As a hydrogen atom has only one electron and one proton, a hydrogen ion is the same as a single proton!	This theory still requires hydrogen atoms to be present within a molecule to be classified as an acid or a base. However, some molecules (e.g. BF_3) act as an acid without containing any hydrogen. Though this is beyond the HSC syllabus, it is worth noting that the more modern Lewis definition resolves this issue.

Topic 2

Using Brønsted-Lowry Theory

SYLLABUS :
Inquiry question: What is the role of water in solutions of acids and bases?

Performing quantitative analysis of chemical solutions and reactions is something that you will be very familiar with by now. Let's apply our mathematical skills to acid-base reactions!

2.1 Calculating concentrations

SYLLABUS :
Calculate pH, pOH, hydrogen ion concentration, and hydroxide ion concentration for a range of solutions.

The only formulas that you will need are:

$$C = \frac{n}{V}$$
$$n = \frac{m}{MM}$$
$$pH = -\log[H^+]$$
$$pOH = -\log[OH^-] = 14 - pH = 14 + \log[H^+]$$

Since pH is a logarithmic function, a change of pH of 1 indicates a ten-times increase/decrease in concentration of hydroxide/hydrogen ions! Let's look to a few simple examples of finding the pH of a substance, and then one difficult one.

SAMPLE :
Question: The concentration of hydrogen ions in a solution is 0.02 **mol/L. What is the pH and pOH of this solution?**

When it comes to maths questions, it's always a good idea to write out the information we are given. In this case, we know that:

$$[H^+] = 0.02 \text{ mol/L}$$

We want to calculate the pH and the pOH of this solution. The relevant formulas (given the fact that we know the concentration of hydrogen ions) are:

$$ph = -\log[H^+]$$
$$pOH = 14 + \log[H^+]$$

Substituting in the value given, we get:

$$ph = -\log(0.02) = 1.7$$
$$pOH = 14 + \log(0.02) = 12.3$$

A good sanity check is the see whether the pH and pOH add up to 14; here, they clearly do!

SAMPLE :

Question: The pOH of a solution is 13.3. What is the concentration of hydrogen ions in the solution?

This question is going to require a bit of mathematics on our part. We know that:

$$\text{pOH} = 14 + \log[\text{H}^+] = 13.3$$

But how do we go from $14 + \log[\text{H}^+] = 13.3$ to some value of $[\text{H}^+]$? We need to understand logarithms!

For a relationship: $10^x = y$, we can say that $\log(y) = x$. This result means '10 to the power of x is equal to y. Rearranging the above equation, we get:

$$\log[\text{H}^+] = -0.7$$

This means '10 to the power of –0.7 equals $[\text{H}^+]$.' So, our answer is:

$$[\text{H}^+]=10^{-0.7} = 0.20\text{ mol/L}$$

SAMPLE :

Question: 2.30 g of NaOH is dissolved into 250 mL of solution. What is the pH of the solution?

This question actually requires more theory than it seems. We can't answer this question without understanding what a **strong and weak acid or base** is. The definitions are pretty simple:

KEY POINT :

A **strong acid/base** ionises completely in solution.

A **weak acid/base** does not ionise completely in solution.

This means that, for example, hydrochloric acid is a strong acid. So we can write:

$$\text{HCl}_{(aq)} \longrightarrow \text{H}^+ + \text{Cl}^-$$

All of the acid, in solution, ionises. However, most acids or bases are weak. An example is acetic acid (CH_3COOH). So we can write:

$$\text{CH}_3\text{COOH} \leftrightharpoons \text{CH}_3\text{COO}^- + \text{H}^+$$

Notice the difference in the arrow? In the **strong acid** case, we use a **right arrow** to indicate that *all* of the acid is ionised. In the **weak acid** case, we use an **equilibrium arrow** to indicate that there is some equilibrium between the ionised molecules, and the 'main' acid.

If you get a question asking about the pH of a solution of a weak acid/base, you can't actually answer it without knowing how much of the substance has ionised. Think about it; the solution may contain 1 mole of acid, but how much of that acid actually turned into hydrogen ions?

However, when you get a question involving strong acids/bases, you know that *all* of it ionised. So, we can answer that question. Conveniently, NaOH (the base in this question) is a strong base.

Let's return to the question. First, we need to find the moles of NaOH added to water using:

$$n = \frac{m}{MM} = \frac{2.3}{39.998} = 0.0575\text{ moles}$$

Next, we find the concentration of the base in solution using:

$$C = \frac{n}{V} = \frac{0.0575}{0.25} = 0.230 \text{ mol/L}$$

Make sure to use SI units (i.e. litres) instead of non-SI units (i.e. mL).

Now we have the concentration of NaOH in solution, and since we know it is a strong base, we can say that:

$$NaOH_{(aq)} \longrightarrow Na^+ + OH^-$$

Thus, the concentration of NaOH will be *equal* to the concentration of OH^-. Therefore:

$$[OH^-] = 0.230 \text{ mol/L}$$

We want to find the pH. So we use our formulas from before:

$$pOH = -\log[OH^-] = 14 - pH$$

Solving this will take two steps. First, we find $-\log[OH^-]$. Then, we set it equal to $14 - pH$.

$$-\log(0.230) = 0.639 = 14 - pH$$
$$pH = 13.362$$

So the solution is super basic (ha ha)!

SAMPLE :

Question: 125 mL of 0.30 mol/L HCl is added to 25 mL 1.3 mol/L NaOH. What is the pH of the solution?

This is a tricky question! It requires a fair few steps, and a good understanding of the molecular interactions of the two substances.

Remember that every neutralisation reaction is just a hydrogen atom chemically reacting with a hydroxide ion. When this happens, water is produced, which has a neutral pH.

Our job is just to find the concentration of hydroxide ions, or hydrogen ions. From there, everything else is easy. But how do we get there when we have both acids and bases floating around?

Recall from the Year 11 course the concept of the **limiting reagent.**

The limiting reagent is the substance that runs out first in the reaction, thus halting the reaction. Let's investigate an arbitrary chemical reaction:

$$A_{(s)} + B_{(s)} \longrightarrow C_{(s)}$$

Substances A and B react to form substance C. Note that the molar ratio tells us how the reaction will occur; one mole of A will react with one mole of B to form one mole of C. A different equation, reading:

$$3A_{(s)} + 2B_{(aq)} \longrightarrow C_{(s)}$$

...would show that three moles of A react with 2 moles of B to form one mole of C. This understanding can help us decide which is the limiting reagent. Take the previous equation as an example. If we begin with 9 moles of A and 9 moles of B, which will run out first? For every three moles of A, 2 moles of B react. So, when 6 moles of A react, 4 moles of B will react. Finally, when 9 moles of A reacts, 6 moles of B will be used up. However, we've now run out of substance A; thus, it is the limiting reagent.

We use the same concept in acid-base interactions. When all is said and done, either the acid or the base will be 'left over'. Our first step will *always* be to write out a balanced chemical reaction.

$$NaOH_{(aq)} + HCl_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)}$$

Our second step is to work out how many moles of hydroxide ions are present in the solution, and how many moles of hydrogen ions are present in solution:

$$C = \frac{n}{V}$$
$$n_{HCl} = 0.30 \times 0.125 = 0.0375 \text{ moles}$$
$$n_{NaOH} = 1.3 \times 0.025 = 0.0325 \text{ moles}$$

So, we have more moles of HCl. The chemical equation above shows that one mole of sodium hydroxide will react with one mole of hydrochloric acid. By analogy, one mole of the hydroxide ions produced by sodium hydroxide will react with one mole of the hydrogen ions produced by hydrochloric acid. Since there are more moles of HCl, all hydroxide molecules will be used up. We will have, left over:

$$n_{H^+} = 0.0375 - 0.0325 = 0.0050 \text{ moles}$$

If you're confused, just think through what we've done. We worked out that NaOH was the limiting reagent. Then, we worked out how many moles of HCl was left over after reaction. Finally, we noted that HCl was a strong acid, and so will ionise completely. This leaves us with 0.0050 moles of solution. But don't forget we're looking for a pH, which requires a concentration of hydrogen ions. We can find that easily:

$$C_{H^+} = \frac{n}{V} = \frac{0.0050}{0.125 + 0.025} = 0.0392 \text{ mol/L}$$

Recall that the volume will be the total volume of the solution (i.e. the sum of the two solutions we have mixed together). Therefore, the pH of our solution turns out to be:

$$pH = -\log[H^+] = -\log(0.0392) = 1.41$$

2.2 Ionic equations

SYLLABUS :
Write ionic equations to represent the dissociation of acids and bases in water, conjugate acid/base pairs in solution and amphiprotic nature of some salts, for example:

- Sodium hydrogen carbonate
- Potassium dihydrogen phosphate

2.2.1 Dissociation of acids and bases in water

We've looked at this pretty extensively in the previous section. To summarise, when an acid or base is in solution, it will ionise. Generally speaking, this will produce either hydrogen ions (acids) or hydroxide ions (bases) as per the Arrhenius definition of acids and bases. The only real question is whether to use an arrow (for strong acids or bases) or an equilibrium symbol (for weak acids or bases). Finally, some acids or bases release more than one proton. This will be looked at later in the course.

2.2.2 Conjugate acid/base pairs in solution

Conjugate acid/base pairs are pairs of acids and bases which differ by a single hydrogen atom (proton). When an acid donates a proton, one of the products will be a base. When a base receives a proton, the product will be an acid!

The two most obvious examples are ammonium (a weak acid) and a fluoride ion (a weak base). When ammonium donates a proton (i.e. loses a hydrogen ion), the chemical equation will be:

$$NH^+_{4(aq)} \longrightarrow H^+ + NH_{3(aq)}$$

So ammonia is produced, which we know is a weak base! In simple terms, the above equation shows that:

$$\text{acid} \longrightarrow \text{proton} + \text{base}$$

When a fluoride ion receives a proton, the chemical equation will be:

$$F^- + H^+ \longrightarrow HF$$

We also happen to know that hydrogen fluoride is an acid, so we can write:

$$\text{base} + \text{proton} \longrightarrow \text{acid}$$

This is a generally accepted principle of acid/base conjugates in solution. We call the base/acid 'produced' the 'conjugate' of the acid/base 'used up'. In neutralisation reactions, we can generalise the above principles as follows.

KEY POINT :
Acid + Base ⇋ Conjugate Base + Conjugate Acid

2.2.3 Amphiprotic nature

An amphiprotic substance is one that can act as both a proton donor and a proton acceptor, depending on which substance it is introduced to. The most obvious one of these is water, as depending on whether it is introduced to an acid or a base, it will either accept or donate protons. For example:

- Water (H_2O):
  - As a proton donor: $Cl^- + H_2O_{(l)} \longrightarrow HCl_{(aq)} + OH^-$
  - As a proton acceptor: $NH^+_{4(aq)} + H_2O_{(l)} \longrightarrow NH_{3(aq)} + H_3O^+_{(aq)}$
- Hydrogen carbonate (HCO_3^-):
  - As a proton donor: $HCO_3^- \longrightarrow CO_3^{2-} + H^+$
  - As a proton acceptor: $HCO_3^- + H^+ \longrightarrow H_2CO_{3(aq)}$
- Dihydrogen phosphate ($H_2PO_4^-$):
  - As a proton donor: $H_2PO_4^- \longrightarrow HPO_4^{2-}$
  - As a proton acceptor: $H_2PO_4^- + H^+ \longrightarrow H_3PO_{4(aq)}$

2.3 Calculating pH

SYLLABUS :
Calculate the pH of the resultant solution when solutions of acids and/or bases are diluted or mixed.

We've actually already looked at a question which exactly relates to this dot point. The general technique is to:

1. Write out the balance chemical equation.
2. Find out how many moles of acid/base you have.
3. Work out the limiting reagent.
4. Calculate the pH of the resulting solution.

Practice makes perfect, so let's look at a new example!

SAMPLE :

Question: 250 mL of 0.33 mol/L HNO_3 (a strong acid) is diluted into 500 mL of 0.1 mol/L KOH (a strong base). What is the resulting pH?

Let's follow the steps above. First, we need to write out the chemical equation. Don't be scared; remember that an acid and a base will always produce a salt and water!

$$HNO_{3(aq)} + KOH_{(aq)} \longrightarrow H_2O_{(l)} + KNO_{3(aq)}$$

Now we need to work out how many moles of the acid and the base we have.

$$\begin{aligned} n &= CV \\ n_{acid} &= 0.25 \times 0.33 = 0.0825 \text{ mol} \\ n_{base} &= 0.5 \times 0.1 = 0.05 \text{ mol} \end{aligned}$$

Clearly, we have more moles of acid than moles of base. So, the base is the limiting reagent! We will be left with 0.0825 – 0.05 = 0.0325 mol of acid after neutralisation occurs. So, our final concentration of hydrogen ions is:

$$[H^+] = \frac{0.0325}{0.5 + 0.25} = 0.043 \text{ mol/L}$$

Substituting this into our pH equation, we get:

$$pH = -\log[H^+] = -\log(0.043) = 1.36$$

Topic 3

Quantitative Analysis

SYLLABUS :
Inquiry question: How are solutions of acids and bases analysed?

3.1 Titration and neutralisation

3.1.1 The titration process

SYLLABUS :
Conduct practical investigations to analyse the concentration of an unknown acid or base by titration.

Whilst this is obviously a practical dot point, I thought it best to summarise the titration process. **Titration** is the method of identifying the unknown concentration of an acid or a base.

First, you will need to prepare a standard solution.

During the HSC Chemistry course, you will prepare so many standard solutions and perform so many titrations that you will be able to discuss this topic in your sleep!

- **Standard solution:** the point of a standard solution is to obtain something with perfectly (or as accurately as possible) known pH/concentration. This is done by dissolving a known mass of solid substance into distilled water, and diluting it to a known volume. The requirements of the substance to be dissolved are:
  - The substance is a water-soluble solid (generally powder)
  - The substance is very pure
  - The substance has a known chemical formula
  - The substance is hydrophobic (meaning it does not absorb water present in the atmosphere, which would make its mass change over time)
  - The substance is stable (meaning it does not react in air)
- The steps to prepare a standard solution are as follows:
  1. Carefully measure the desired quantity of the solid substance on an electronic balance. To determine the desired quantity, decide on the ideal concentration, and use: concentration $= \dfrac{\text{moles}}{\text{litres}}$, and moles $= \dfrac{\text{mass}}{\text{molar mass}}$.
  2. Fully dissolve the solid in a small beaker using distilled water.
  3. Transfer the solution into a volumetric flask. Make sure to wash the beaker and any mixing devices into the volumetric flask, so as not to lose any of the solution.
  4. Fill the volumetric flask until just below the calibration line, and then add distilled water drop-wise until the meniscus is at the calibration line.
  5. Hold your thumb over the stopper and turn the volumetric flask upside down once or twice to ensure the solution is homogenous.

The point of a titration is to discover the concentration of an unknown solution. Now that we have a standard solution, with an accurately-known concentration, we can do this. By slowly adding one solution to the other until they are neutral, we can calculate the unknown concentration.

Before actually performing the titration, it is important to select the correct indicator. By adding a few drops of **indicator** to the solutions, it becomes easy to tell when the pH changes and the solution becomes neutral. First, test small quantities of each solution to find their pH. This tells you which indicator to use; if they are a strong acid and a strong base, you want an indicator with an equivalence around 7 (i.e. bromothymol blue).

If they are a strong acid and a weak base, you want a neutralisation point of around 4–5. An appropriate indicator is one with an endpoint at approximately 4–5, such as methyl orange. Similarly, if the solutions are a weak acid and a strong base, you'll want to use an indicator such as phenolphthalein.

To perform the titration itself, you add a known quantity of the standard solution to a conical flask and place it below a burette filled with the unknown liquid. After adding a few drops of the appropriate liquid, solution is slowly dripped into the standard solution. When the colour changes, the quantity added is recorded. This process is repeated until at least 3 tests are within 0.1 mL of each other.

Once the quantity is determined, the chemical equation can be used to calculate the unknown concentration. And to determine the concentration, we need to do some maths. Let's look at an example:

SAMPLE :

Question: A standard solution of NaOH (1.1 moles/L) is titrated against a solution of HCl with unknown concentration. On average, 25 mL of standard solution requires 27.8 mL of titre. What is the concentration of the HCl solution?

First, it is important to recognise the chemical formula relevant to the question. In this case:

$$NaOH_{(aq)} + HCl_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)}$$

Recall that an **acid + base ⟶ salt + water.** Clearly, the molar relationship of neutralisation is 1:1 (i.e. one mole of sodium hydroxide is required to neutralise one mole of hydrochloric acid).

Now, we need to figure out how many moles of NaOH were in each test. As we know:

$$\text{concentration} = \frac{\text{moles}}{\text{litre}}$$

Therefore, for the NaOH solution:

$$1.1 = \frac{\text{moles}}{0.025}$$
$$\text{moles} = 0.0275$$

As the molar relationship is 1:1, the same number of moles of HCl must have been added to the solution in order to neutralise it. Therefore, for the HCl solution:

$$\text{concentration} = \frac{0.0275}{0.0278} = 0.989$$

Therefore, to two significant figures, the concentration of the HCl solution is 0.99 moles/L.

Topic 3 – Quantitative Analysis

3.1.2 Titration curves

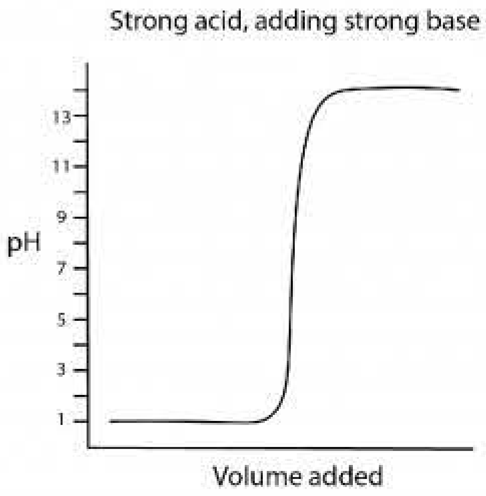
SYLLABUS :

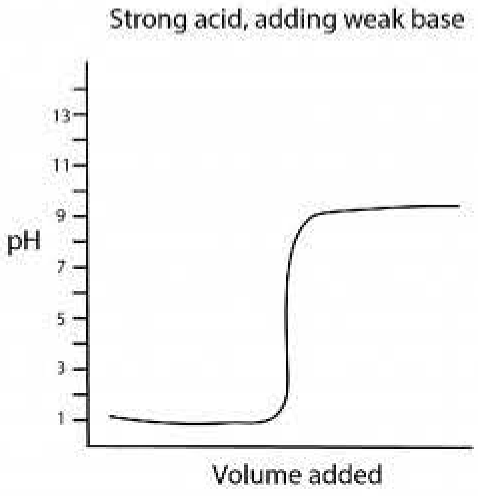
Investigate titration curves and conductivity graphs to analyse data to indicate characteristic reaction profiles, for example:

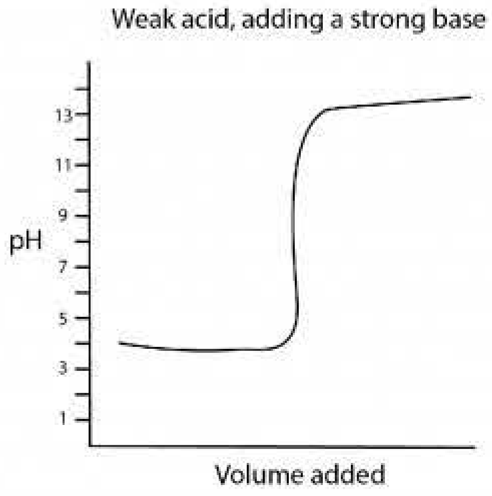
- Strong acid/strong base
- Strong acid/weak base
- Weak acid/strong base

A titration curve is a plot of pH vs the amount of titrant added. If you start off with an acidic substance, and continuously add base, you expect the pH to increase over time. The converse will be true for a basic substance and the addition of an acid.

The following are some model titration curves, based on the type of titration you might be performing.







You'll notice that the strong acid/strong base plot is fairly symmetrical, whilst the strong acid/weak base is centred on the acidic portion of the pH scale, and the weak acid/strong base is centred on the basic portion of the pH scale. We call that centre point the **equivalence point** of a titration; basically, it describes the **neutralisation point** where all hydrogen ions are neutralised by hydroxide ions.

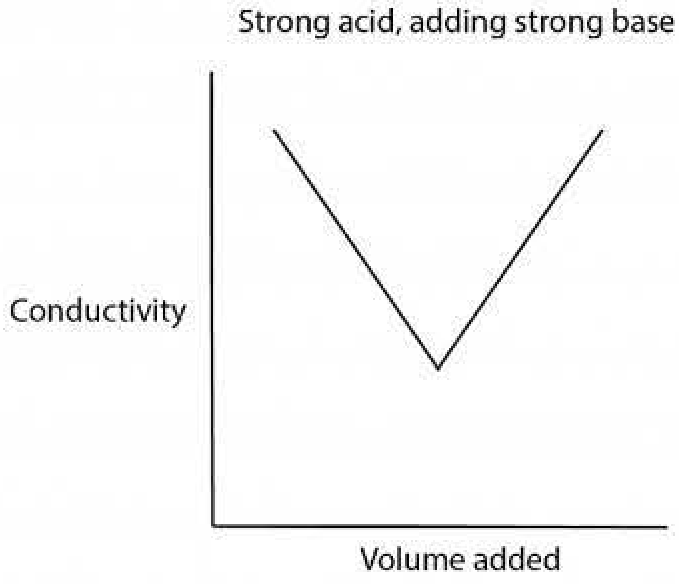
This equivalence point will occur around a pH of 7 for the strong acid/strong base titration, around 3–5 for the strong acid/weak base titration, and around 9–11 for the weak acid/strong base titration.

3.1.3 Conductivity graphs

Another way to test the neutralisation of a solution is through a conductivity probe. Conductivity of electricity in solution is dependent on the concentration of ions in that solution. So, the more ions, the more conductive a solution will be. By testing conductivity, we can get an idea of the solution's neutralisation.

When a solution is very acidic or very basic, there will be *loads* of hydrogen or hydroxide ions floating around. These ions conduct current. But, at the neutralisation point, there will be far fewer ions present. As such, conductivity will decrease until it reaches a 'minimum,' then increase again as you continue to add ionising solution.

The following is a model conductivity curve. You can clearly see where neutralisation occurs.



Note that the plot will not reach 'zero' conductivity, as the solution will contain impurities which will be capable of carrying current.

3.2 The dissociation constant and strong/weak acids

SYLLABUS :
Calculate and apply the dissociation constant (K_a) and pK_a to determine the difference between strong and weak acids.

Recall from earlier that for chemical equation involving an acid HA: $HA + H_2O_{(l)} \leftrightharpoons A^- + H_3O_{(aq)}$, the acid constant is: $K_a = \frac{[A^-][H^+]}{[HA]}$. Similarly, we can simply define: $pKa = -\log(Ka)$. This is useful, as the strength of an acid (as with its pH) is situated on a logarithmic scale. Recall that strong acids ionise completely in solution, which weak acids do not. But this definition isn't as precise as it could be. We can do better than that! Let's calculate the K_a of a strong acid, like HCl. The relevant 'equilibrium' is going to be: $HCl \leftrightharpoons H^+ + Cl^-$. However, we know that the concentration of the acid (in un-ionised form) will be close to zero. This is because, as a strong acid, almost all of it will be ionised. Looking to our acid constant equation: $K_a = \frac{[A^-][H^+]}{[HA]}$.

We note that [HA] could *never* be zero, as then we would be dividing by zero, which is illegal and would cause the universe to implode. However, we can say that as [HA] gets *very close* to zero, K_a will get very, very large. Conversely, if [HA] was very large (i.e. very little of the acid actually ionised), the value of K_a would be quite small. So, we can draw conclusions regarding some sort of general rule.

KEY POINT :
A **high** K_a will indicate a **strong** acid.

A **low** K_a will indicate a **weak** acid.

Weak acids can be 'varyingly' weak, and some acids are much weaker than others, even though we describe both of them as 'weak'. The actual values of K_a are often very, very small. Here are some examples:

Acid	K_a
Sulphuric acid	1.2×10^{-2}
Nitrous acid	4.0×10^{-4}
Carbonic acid	4.8×10^{-11}

This is not particularly useful, because it's hard to actually visualise the acid constant values. This is why we use a logarithmic scale: $pKa = -\log(Ka)$.

Acid	pK_a
Sulphuric acid	1.92
Nitrous acid	3.40
Carbonic acid	10.32

As the acid gets weaker, the pK_a will get larger.

KEY POINT :
A **high** pK_a will indicate a **weak** acid.

A **low** pK_a will indicate a **strong** acid.

3.3 Analysis technique applications

SYLLABUS :

Explore acid/base analysis techniques that are applied:

- In industries
- By Aboriginal and Torres Strait Islander Peoples
- Using digital probes and instruments

Industrial application

Titrations are used fairly regularly in industry, although obviously in a bit more of an advanced way than you performed them in class!

Water treatment facilities must have an accurate idea of the pH of the water, so they can appropriately treat it for consumption. Often, water will begin acidic, due to acid rain run off or leeching. Then, the water will be made basic to promote flocculation of foreign substances. Finally, the water will need to be made neutral (or slightly alkaline).

Rain is often tested to determine its level of acidity. Acid rain is a huge problem, affecting natural life, waterways, crops, and human existence. Quickly and efficiently determining pH is crucial to minimise the otherwise devastating impact of acid rain.

The food industry often uses acid/base analysis techniques as this can be used to determine the amount of certain substances, such as sugar, salt, or vitamins, is present. This is hugely important for societal safety, and ensure products on our shelves are not a danger to those who might buy them.

Aboriginal and Torres Strait Islander application

This is likely an aspect of the syllabus dot point that your teacher will discuss in class. Suffice it to say that if a community is to survive tens of thousands of years, it will have some method of dealing with the concerns caused by acidic or basic substances. Whether or not a 'theory' of acids and bases was developed, we too often overlook the advanced scientific understanding Aboriginal and Torres Strait Islander communities had and have over the environment them lived in.

Digital probes and instruments

Throughout the year, you may have used pH probes or conductivity probes. These instruments use the conductivity of a solution to determine its pH. Recall that conductivity is a good measure of the concentration of ions in solution. By comparing to a water-only 'baseline,' such a probe is a quick and cheap way of determining the acidic/basic properties of a solution.

SYLLABUS :

Describe the importance of buffers in natural systems.

A buffer solution is a substance that resists changes in pH when another solution, such as an acid or a base, is added. They are vitally important in our bloodstream, as they ensure that our blood does not become too acidic or too basic when acids (such as carbon dioxide) or bases come into contact with it. Buffer solutions often contain weak acids and their conjugate bases. In blood, the buffer system is:

$$H^+_{(aq)} + HCO^-_{3(aq)} \leftrightharpoons H_2CO_{3(aq)} \leftrightharpoons H_2O_{(l)} + CO_{2(g)}$$

To incorporate your knowledge of acid/base reactions and Le Chatelier's principle into this dot point, ask yourself what happens once you add CO_2 (e.g. during exercise), and how that affects the equilibrium.

Part III

Module 7: Organic Chemistry

Topic 1

Nomenclature

SYLLABUS :
Inquiry question: How do we systematically name organic chemical compounds?

Naming organic chemicals is extremely important in Chemistry. We need a universal way of communicating what chemical we are talking about. The IUPAC naming conventions attempts to produce such a universal method!

1.1 Naming organic compounds

SYLLABUS :
Investigate the nomenclature of organic chemicals, up to C8, using IUPAC conventions, including simple methyl and ethyl branched chains, including:

- Alkanes
- Alkenes
- Alkynes
- Alcohols (primary, secondary and tertiary)
- Aldehydes and ketones
- Carboxylic acids
- Amines and amides
- Halogenated organic compounds

First, let's learn what an organic compound actually is! **Organic chemistry** is chemistry related to **substances containing carbon.** This is vitally important!

KEY POINT :
Organic compounds contain **carbon.**

When naming carbon-compounds, the number of carbon atoms is very important. You need to know the nomenclature of organic chemicals up to C8 (i.e. up to compounds containing 8 carbon atoms).

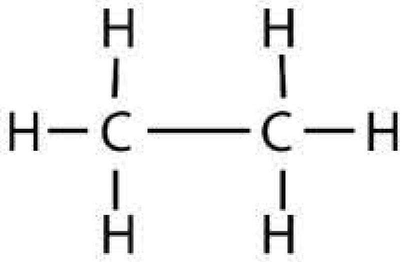
Number of Carbon Atoms	Nomenclature (Prefix)
1	Meth-
2	Eth-
3	Prop-
4	But-
5	Pent-
6	Hex-
7	Hept-
8	Oct-

KEY POINT :
The **number** of carbon atoms determines the **start** of the name of the compound.

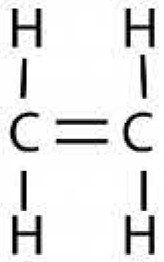
1.1.1 Alkanes

Alkanes are saturated organic compounds containing only carbon and hydrogen atoms. A compound is **saturated** when no additional atoms could possible bind to the compound; in other words, there can be **no double bonds** present between carbon atoms. For example:

Alkane

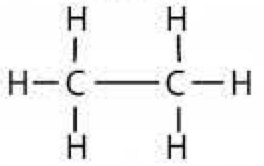


Not an Alkane

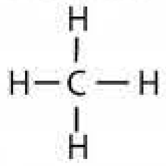


You can see that the alkane has single bonds, whilst the other substance (which we will soon find is an alkene) has a double bond. On the right are some examples of named alkanes. The most interesting one of these is the third. Note that, even though there is a weird bend going on, we simply count the number of carbons. We could shift the compound around a bit, and it would look like any other straight organic compound! The number of carbons will determine the start of the name, and the type of compound (here, the fact that it is an alkane) will determine the end of the name.

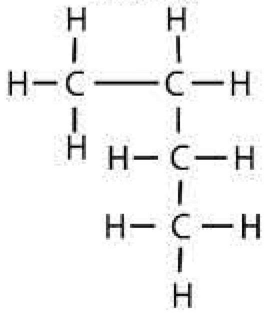
Ethane



Methane



Butane

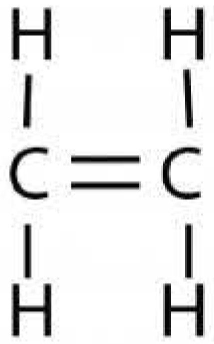


Key Point :
Alkanes contain only single-bonded carbon and hydrogen atoms, and end in **-ane.**

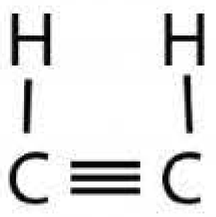
1.1.2 Alkenes

Alkenes are unsaturated organic compounds containing only carbon and hydrogen atoms, and at least one double-bonded carbon atom. A compound is **unsaturated** when additional atoms could possible bind to the compound; in other words, there **must be double bonds** (or triple bonds, etc.) present between carbon atoms. For example:

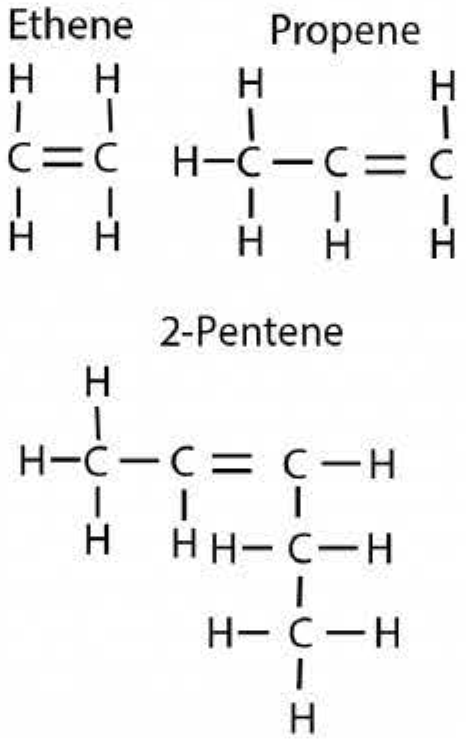
Alkene



Not an Alkene



You can see that the alkene has a double bond, whilst the other substance (which we will soon find is an alkyne) has a triple bond. On the right are some examples of named alkenes. As long as a compound has at least one double bond (and the rest are single bonds), it will be an alkene. However, it is important that we know where the double bond actually lies. The number in front of the pentene describes where the double bond is placed. To determine the number, we could count from both sides and see how many bonds from the end it takes to get to our double bond. From the 'top-left' carbon, the double bond is the second inter-carbon bond. From the 'bottom right' carbon atom, the double bond is the third inter-carbon bond. *We always choose the lowest number!*



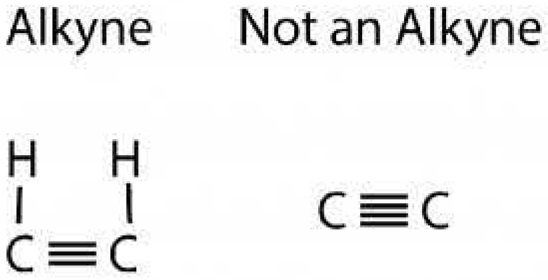
Key Point :
To name 'where' the bond of interest is, count from both sides and choose the **lowest** value.

Note that propene did not get a number, as no matter where we put the double bond, the answer would have been 1-propene.

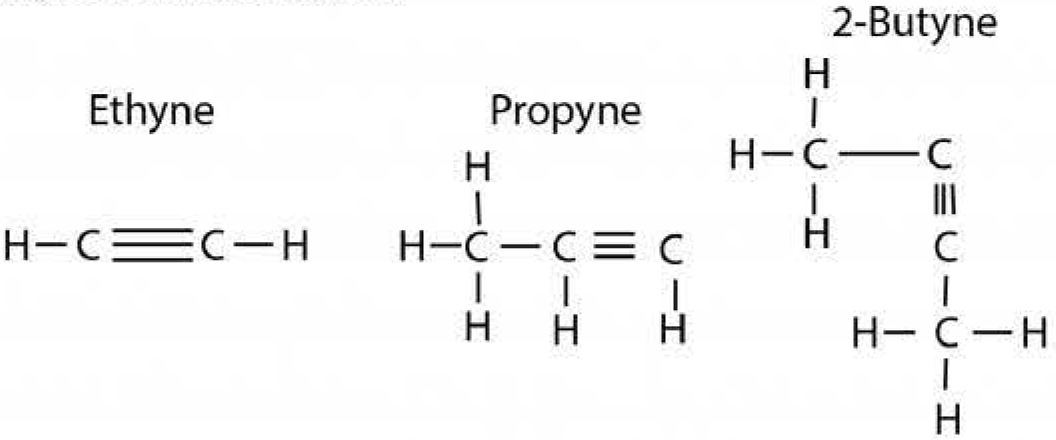
Key Point :
Alkenes contain only carbon and hydrogen atoms, and at least one set of double-bonded carbon atoms, and end in **-ene.**

1.1.3 Alkynes

Alkynes are unsaturated organic compounds containing only carbon and hydrogen atoms, and at least one triple-bonded carbon atom. For example:



Below are some examples of named alkynes:



As long as a compound has at least one triple bond (and the rest are single bonds), it will be an alkyne.

Key Point :
Alkynes contain only carbon and hydrogen atoms, and at least one set of triple-bonded carbon atoms, and end in **-yne.**

1.1.4 Alcohol

Alcohols are organic chemical compounds containing –OH groups. As you probably know, substances defined as alcohols will end in **-ol.**

Naming alcohols is a bit trickier than our previous naming conventions. This is due to the existence of primary, secondary and tertiary alcohols.

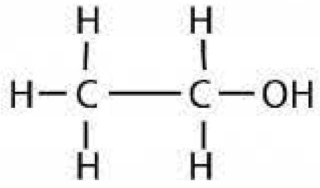
- **Primary alcohols:** are chemicals in which the carbon atom carrying the –OH group is only attached to a single alkyl group.
- **Secondary alcohols:** are chemicals in which the carbon atom carrying the –OH group is attached to two alkyl groups.
- **Tertiary alcohols:** are chemicals in which the carbon atom carrying the –OH group is attached to three alkyl groups.

The easiest way to understand this is to name a few compounds.

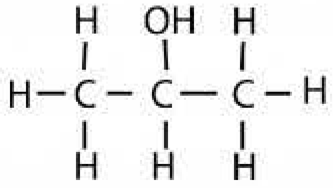
The first compound (ethanol) is a primary alcohol. The second compound (propan-2-ol) is a secondary alcohol. Note that the '2' in the –OH group was placed before the -ol suffix in the name of the compound. This makes it easier to identify that the '2' we care about is relevant specifically to the 'alcohol' aspect of the compound. Finally, the third compound (2-methlypropan-2-ol) is a tertiary alcohol. The 2-methyl aspect relates to the fact that the second carbon is where the methyl groups (i.e. –CH_3 groups) are attached.

An alternate naming convention for tertiary alcohols is 'tert-___ alcohol' where the blank is dependent on the total number of carbons. So, for instance, the alternate name of the third structure would be 'tert-butyl alchohol'.

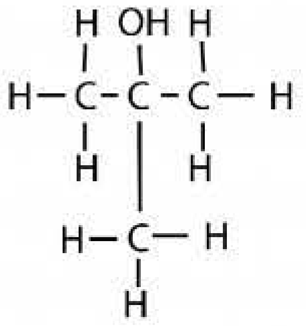
Ethanol



Propan-2-ol



2-methylpropan-2-ol

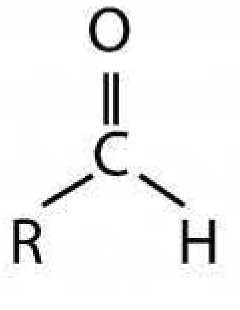


Key Point :
Alcohols contain an –OH group, and end in **-ol.**

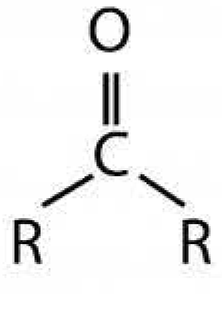
1.1.5 Aldehydes and ketones

Aldehydes and ketones are quite similar in form, but the naming conventions are a bit tricky. An aldehyde is where a carbon atom is double bonded to an oxygen, single bonded to a hydrogen, and then bonded to some further chemical group. Ketone, on the other hand, is double bonded to an oxygen, and has two separate single bonds to further chemical groups. Sound confusing? A diagram always helps!

Aldehyde



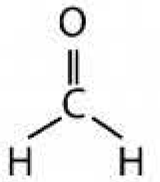
Ketone



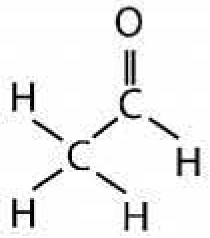
Here, 'R' just denotes some continuing carbon chain of some description. Compounds with an aldehyde group will end in **-al,** and compounds with a ketone group will end in an **-one.** Typically, aldehyde nomenclature doesn't use numbers to denote the location of the aldehyde, as they must naturally terminate on one side.

Below are some more examples to help you understand this nomenclature.

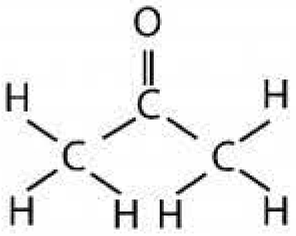
Methanal



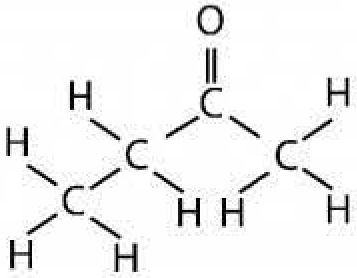
Ethanal



Propanone



2-Pentanone



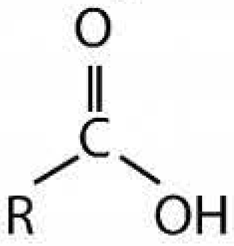
Key Point :
Aldehydes end in **-al.**

Ketones end in **-one.**

1.1.6 Carboxylic acids

Carboxylic acids are chemical compounds containing a –COOH group (a carbon atom, with a single double-bonded oxygen and a single bonded hydroxide group). You've already encountered one of these already acetic acid (which has the IUPAC name: ethanoic acid) with the chemical formula CH_3COOH. To name a carboxylic acid, all you have to do is count how many carbon atoms are present, and add an **-oic acid** on the end!

Carboxylic acid



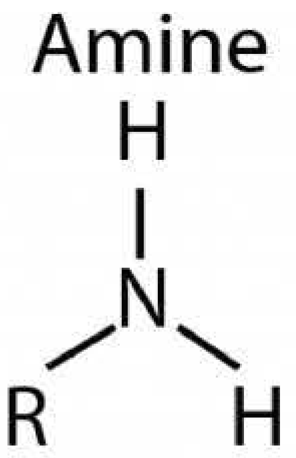
Number of atoms	Chemical formula	IUPAC name	Common name
1	HCOOH	Methanoic acid	Formic acid
2	CH_3COOH	Ethanoic acid	Acetic acid
3	C_2H_5COOH	Propanoic acid	Proionic acid
4	C_3H_7COOH	Butanoic acid	Butyric acid
5	C_4H_9COOH	Pentanoic acid	Valeric acid
6	$C_5H_{11}COOH$	Hexanoic acid	Caprioc acid
7	$C_6H_{13}COOH$	Heptanoic acid	Enanthic acid
8	$C_7H_{15}COOH$	Octanoic acid	Caprylic acid

Whilst you don't need to know the common names of the carboxylic acid, for those of you hoping to go on and study chemistry at university, you are more likely to use the common names than the IUPAC names!

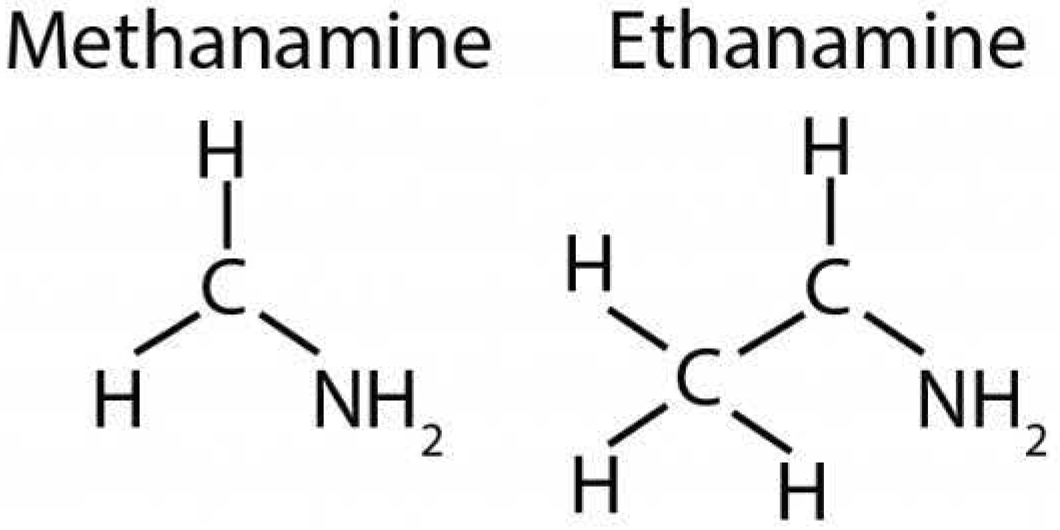
Key Point :
Carboxylic acids contain a –COOH group, and end in **-oic acid.**

1.1.7 Amines

An amide is a functional group which, added to an acid, will change an -oic acid into an **-amine.**



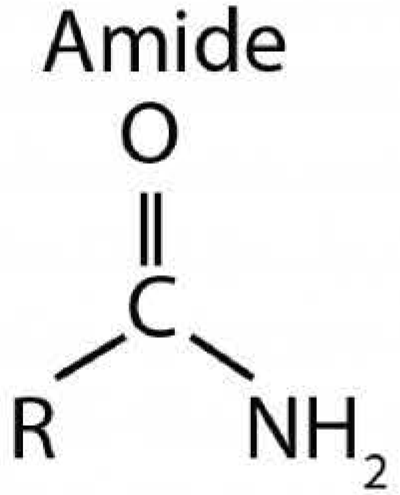
Some examples of compounds with amine groups are shown below. It's worth noting that, as with alcohols, amines have primary, secondary, and tertiary classes. However, as this isn't specified in the relevant dot point, it is likely you don't need to know that. The rules apply in much the same way as they do for alcohols.



Key Point :
Amines contain a $-CNH_2$ group, and end in **-amine.**

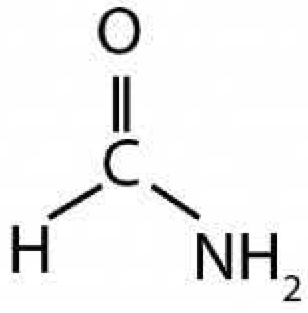
1.1.8 Amides

An amide is a functional group which, added to an acid, will change an -oic acid into an **-amide.** You can think of it as a carbon atom bonded to both an oxygen atom and an amine group.

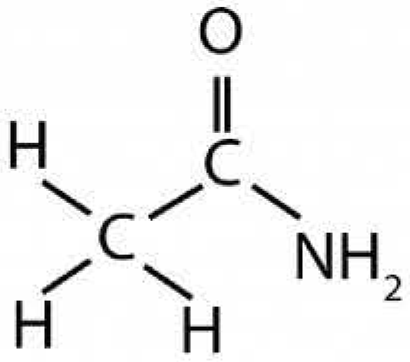


Some examples of compounds with amide groups are below. There are also primary, secondary, and tertiary classes of amides.

Methanamide



Ethanamide



Key Point :
Amides contain a $-CONH_2$ group, and end in **-amide.**

1.1.9 Halogenated organic compounds

Recall that **halogens** are **group seven** elements (e.g. chlorine, fluorine, bromine etc.)

When hydrocarbons (long chains of carbons, with only hydrogens bonded to them) react with halogens, the will often 'take on' one or more halogen atoms.

When a halogen bonds with an organic compound, its name will be placed at the start of the compound name. The prefixes are:

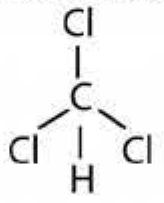
Halogen	Prefix
Fluorine	Fluoro-
Chlorine	Chloro-
Bromine	Bromo-
Iodine	Iodo-

To name halogenated organic compounds, we follow a number of steps:

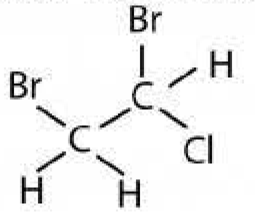
1. Count the longest carbon chain. This will be the name that comes after the prefix.
2. Locate the halogens, and count from the end which gives the lowest values. Do this with all the halogens or other functional groups.
3. If there are multiple of the same halogen, use prefix 'di', 'tri', etc.

Let's look at some examples:

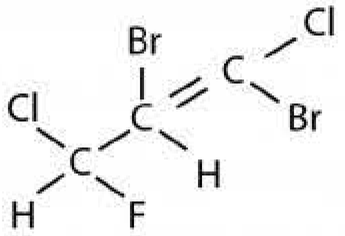
Trichloromethane



1,2,-dibromo-1-chloroethane



1,2,-dibromo-1,3,-dichloro-3,-fluoropropene



A few things are immediately obviously. Firstly, these things look tricky, but if you take them step-by-step are actually not to difficult. Secondly, there are a few rules that we are not familiar with that we're going to need to understand. Thirdly, the rules like '-ene' used for double-bonded carbons still apply.

The rules we don't yet know relate to **numbering.** Numbering is used to 'place' the halogens in the right place. The 'di', 'tri', etc. are related to the total number of a particular halogen in a compound. The 1-, 2-, etc. are used to precisely place the halogen. You must *always* count from a particular side. If you start counting in relation to one side of the carbon chain, all other values must be counted from that same side. So, how do you decide? The rule is that you **count from the side which gives the lowest total number.** So, you'll need to write out the two possible counting mechanisms (i.e. from the left, and from the right), and then add up all the numbers. Then you choose whichever value is the lowest!

Key Point :
Halogens have their own prefixes which are dependent on how many of them are present. When deciding which side to count from, use the side which gives the **lowest total value** when adding up the numbers.

There are *a lot* of naming conventions here. But while it may seem difficult to remember them all, it's actually fairly formulaic. If you're having trouble, try to write a really succinct set of colourful notes (even more succinct than these) which quickly explain when to use which rule. Maybe do it in table form, so you have a quick reference sheet!

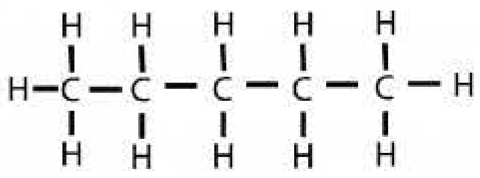
1.2 Structural isomers

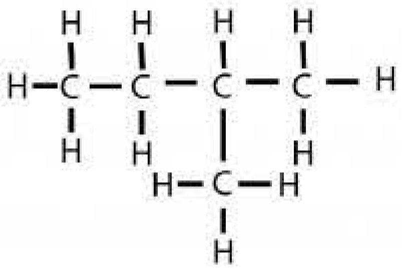
Syllabus :
Explore and distinguish the different types of structural isomers, including saturated and unsaturated hydrocarbons, including:

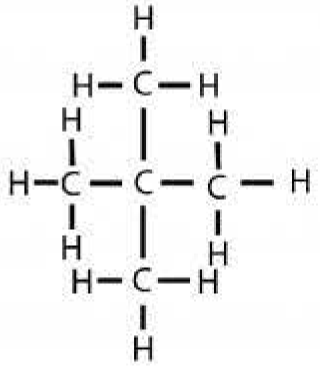
- Chain isomers
- Position isomers
- Functional group isomers

Over the past few syllabus dot points, we've investigated the correct ways to name various organic compounds. But why is it important that we distinguish between compounds, just based on where certain elements are placed? Shouldn't the molecule act in exactly the same way, as long as it contains the same elements?

It turns out that the **placement** of elements or functional groups in an organic molecule actually **changes its chemical properties.** This is due to the geometry of bonding; whilst each of the following contain the same elements (C_5H_{12}), you can imagine that the way in which they bond will be quite different.





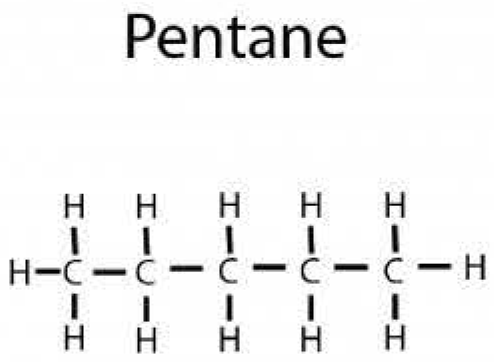


We call compounds that have the **same molecular formula** (e.g. C_5H_{12}) but **different structures** or bonding organisation **structural isomers.**

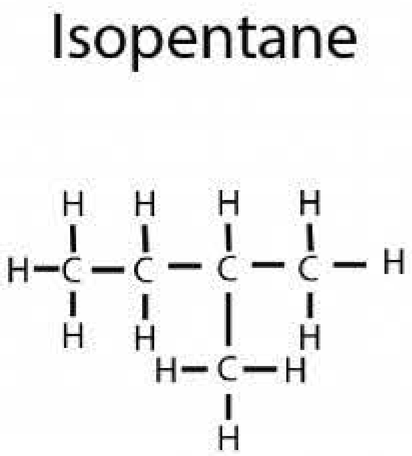
Key Point :
Two compounds will be **structural isomers** if their molecular formula is the same, but their bonding is different.

1.2.1 Chain isomers

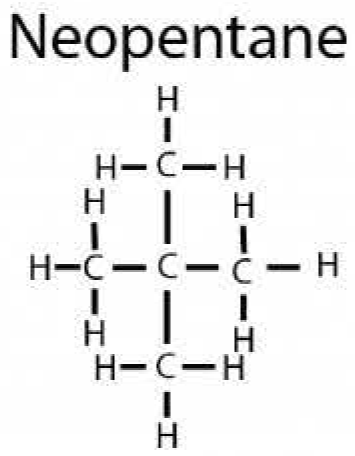
Chain isomers are isomers that differ based on the carbon backbone. The image below is a good example of chain isomerisation; the only thing that has changed is the way in which the carbon backbone is structured. For example, pentane (below) has three chain isomers. Whilst it is unlikely that you'll need to know their names, these chain isomers are:



Pentane



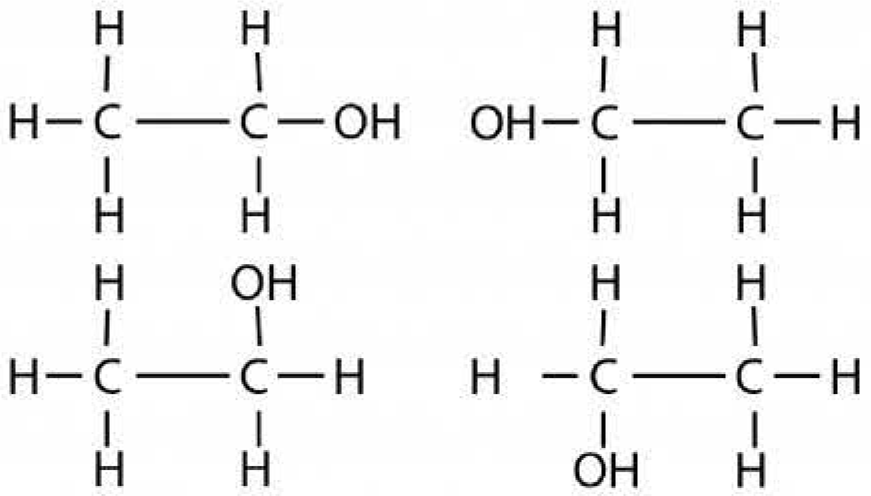
Isopentane



Neopentane

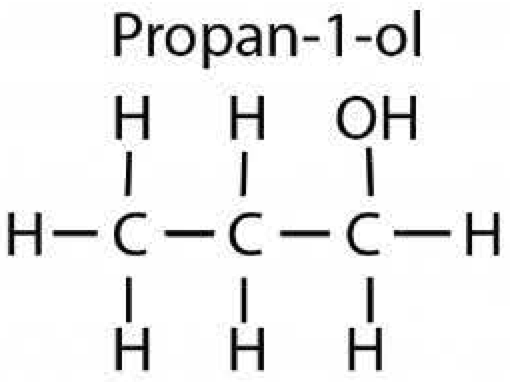
1.2.2 Position isomers

Position isomers are isomers that differ based on the location of the various functional groups. Remember that we can spatially 'flip' any compound, and it will still be exactly the same thing. For instance, all four of the following compounds are the same (i.e. not isomers), and also happen to be ethanol.

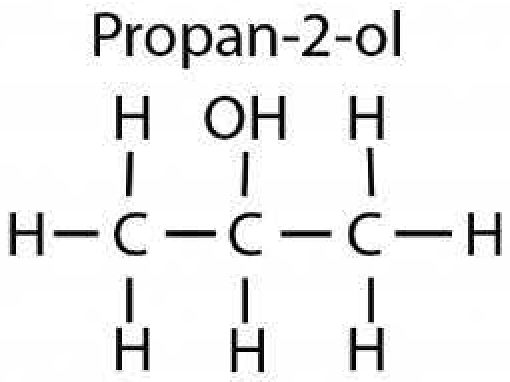


The top two compounds are identically, as they are just flipped along the vertical axis. The bottom two are the same as the top two, as the hydroxide (–OH) functional group is still bonded to the same position on the carbon backbone.

However, longer carbon chains have multiple position isomers. Propanol has two position isomers:



Propan-1-ol



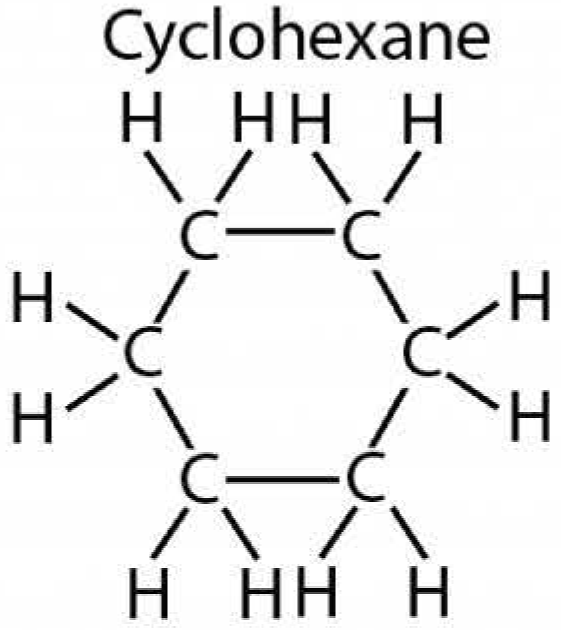
Propan-2-ol

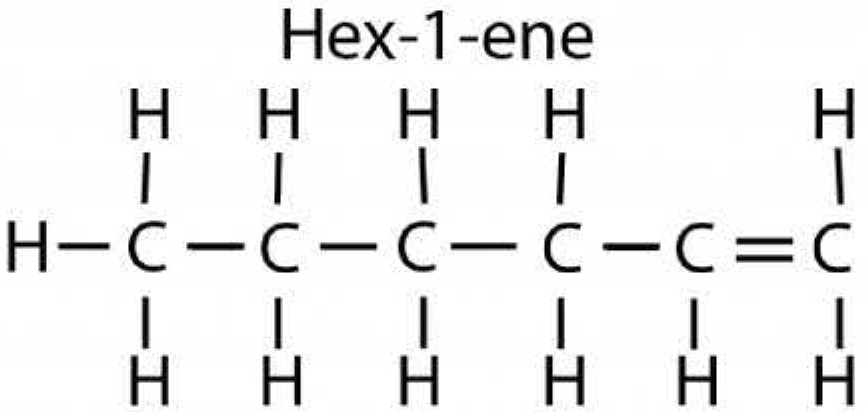
However, if you try to draw *another* position isomer, you'll see that each is just a repeat of one of the above! Give it a try and see what you get.

1.2.3 Functional group isomers

Functional group isomers sound like they should be the same as position isomers. However, rather than moving the functional groups around, this form of isomer changes the functional groups, whilst keeping the molecular formula the same.

Examples of functional group isomers are a bit trickier to come by, and a bit harder to see at first glance. The easiest example is cyclohexane, and hex-1-ene. Recall that hex-1-ene is an alkene (i.e. contains a double bonded carbon). Cyclohexane is similar to a regular hexane compound, however it 'rings' in on itself.





The two compounds have the same molecular formula (C_6H_{12}) but are in effect totally different. The functional group relevant to hex-1-ene is the double bond, whilst the significant aspect of cyclohexane is the ring shape. If anything, it's just coincidental that they have the same molecular formula!

Key Point :

Chain isomers vary based on the carbon backbone.

Position isomers vary based on the location of specific functional groups.

Functional group isomers vary based on changes to functional groups.

Topic 2

Hydrocarbons

SYLLABUS :
Inquiry question: How can hydrocarbons be classified based on their structure and reactivity?

2.1 Structure and bonding experiments

SYLLABUS :
Conduct an investigation to compare the properties of organic chemical compounds within a homologous series, and explain these differences in terms of bonding.

Whilst you'll obviously be doing this in class, it's worth quickly summarising some of the important outcomes of an experiment like this.

We can investigate the homologous series of standard straight-chained alkanes (i.e. with no branching). This series involves methane, ethane, propane, butane, pentane, and so on. The main physical difference between each of these is an addition of a $-CH_2-$, which means that each successive molecule has an increased molecular weight.

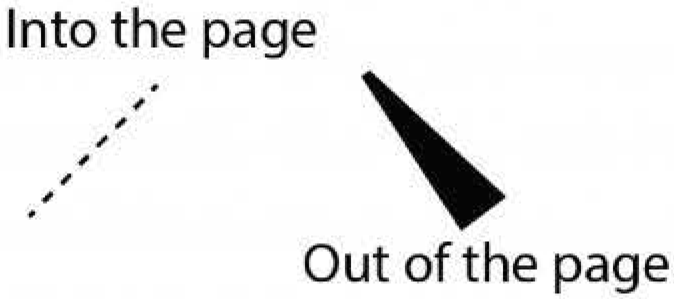
You will also find that boiling point increases with mass. This can be explained due to dispersion forces, which you learned about in the Year 11 course. As electrons zoom around, at some point there may be more electrons in one place than in another. This creates a net negative charge, which means there must be a net positive charge somewhere else. These are very small dipoles (areas of net charge) which cause the molecules to line up such that one negative dipole attracts another positive dipole. Larger molecules exhibit stronger dispersion force dipoles, and so their intermolecular forces are stronger. As such, it takes more energy to rip them apart. So, boiling heavier molecules will typically require more energy than boiling lighter molecules!

2.2 Bonding shapes

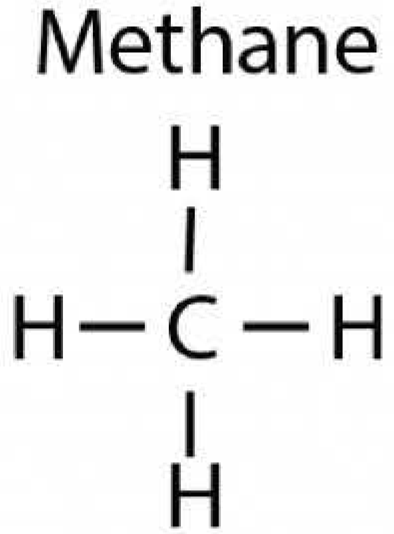
SYLLABUS :
Analyse the shape of molecules formed between carbon atoms when a single, double or triple bond is formed between them.

Up until now, we haven't really cared about the shape of molecules. However, in practice, molecular shape is extremely important to chemical bonding. This dot point serves as a brief introduction to the way we think about molecular shape.

First, we need to establish some terminology. Since molecules are three-dimensional, and a piece of paper is two-dimensional, we need to be able to represent a 3D model on a 2D plane. The usual convention is the used dotted and wedged bonds. Dotted bonds indicate that the bond goes 'backwards' (i.e. into the page), whilst wedged bonds indicate that the bond goes 'forwards' (i.e. out of the page).

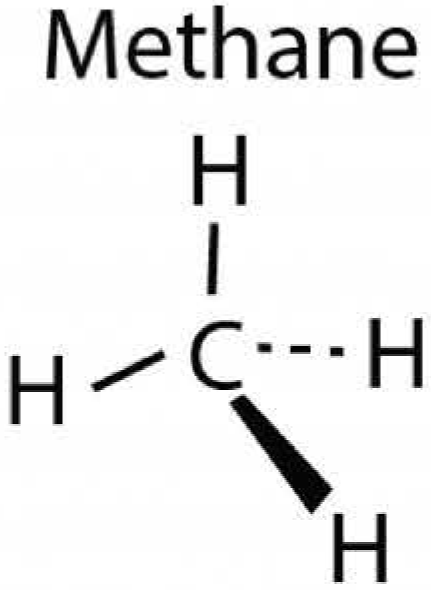


Let's put this into practice by looking to the simplest organic molecule we can thing of: methane. The chemical formula for methane, as hopefully you know by now, is CH_3. So far, we have been displaying it as:



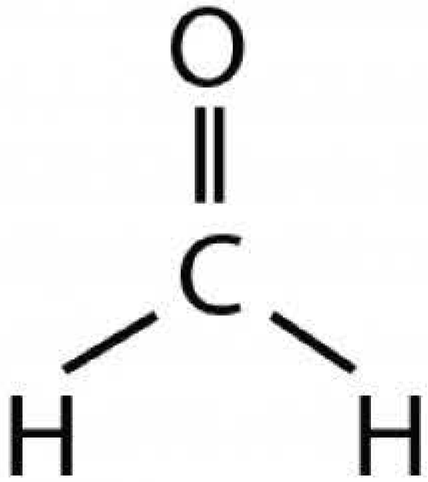
However, this doesn't really make sense. We've automatically understood that the hydrogen atoms will want to spread out as much as possible, due to the fact that chemical bonding is really just above orbiting electrons, and electrons (being negative) will want to repel each other as much as possible.

However, the hydrogen atoms above have only spread out in a two-dimensional plane. In reality, they will spread out across a three-dimensional plane, and form a **tetrahedral configuration.** The angle between each bond will be 109.5°, and the structural configuration will look more like the image on the right.



It's unlikely that you'll need to replicate anything complicated in an exam style environment; just know that the purpose of 3D models like the one above is to represent the need for atoms to spread out as much as they are able to.

Where carbon atoms have double bonded to a functional group (e.g. oxygen), then there will only be two remaining bonds. As there are effectively three things coming off the central carbon atom, a 2D triangle will actually be the most space efficient configuration. Thus, we get the image on the left.



This is the correct structural diagram, even though there are no three dimensional elements.

Finally, where a carbon atom has triple bonded with another functional group, it will only be able to bond to another atom by a single bond. As there are only two groups coming off the central carbon, the resulting molecule will be linear.

$$H-C\equiv C-H$$

2.3 Chemical properties

SYLLABUS :
Explain the properties within and between the homologous series of alkanes with reference to the intermolecular and intramolecular bonding present.

We've already described the change in boiling points within a homologous series. As the compound gets larger, the boiling point will increase due to strengthened intermolecular forces.

The chemical properties within a homologous groups are typically quite similar. This is because the functional groups (or lack thereof) are the same, as will be the approximate shape of the molecule. However, properties such as reactivity and rates of reaction may change within the homologous group.

The reactivity of alkanes is quite low. However, the smaller the alkane is, the more reactive it will be. Thus, methane is more reactive than octane.

SYLLABUS :
Describe the procedures required to safely handle and dispose of organic substances.

Handling organic substances safely is extremely important in the lab. Organic substances are often flammable, and toxic in high quantities. Ensuring distance between organic substances and the experimenters face and decrease fire risk by not using open flames when experimenting on organic substances, are just some of the procedures important to consider. In the event of a fire, having a fire blanket and fire extinguisher on hand is crucial.

Disposal of chemicals is an extremely important part of everyday chemistry. Ensuring safe procedures are developed and followed through on is vital to the continuing of chemical research in society.

Generally speaking, liquid organic chemicals cannot just be poured down the sink. Imagine pouring petrol or alcohol down the sink and having it pass through drainage and end up in the ocean. Doesn't sound very safe or environmentally friendly!

Instead, organic liquids need to be packaged in specialised containers which every lab will stock. These containers will typically be distinguished by whether or not the compound in question contains some form of halogen. These containers, when full, are sealed firmly and then generally transported by an independent company to disposal centres.

SYLLABUS :
Examine the environmental, economic, and sociocultural implications of obtaining and using hydrocarbons from the earth.

There are two main reasons that alternative sources of energy obtained typically from the petrochemical industry are necessary: fossil fuels are **non-renewable,** and they are **pollutants.**

Petrochemicals improve our quality of life as plastics and fuels. In the last century, our reliance on these has grown as they are easy to use and very cheap. However, the CSIRO estimates that our sources of petrochemicals will run out within 50–80 years. Fossil fuels account for about 95% of the world's energy demands and the consumption of fossil fuels has doubled over the past 20 years.

Petrochemicals used as fuels are destructive to the environment. The products of combustion (water and carbon dioxide) contribute to thermal pollution and climate change. Carbon (as dust) induces respiratory problems and lung cancer. Carbon monoxide is toxic when inhaled, and sulphur dioxide is also toxic and causes forest degradation in the form of acid rain.

Furthermore, **plastics** are non-biodegradable if they are man-made polymers. This decomposition problem increases the number and size of landfills, further contributing to environmental destruction.

Biomass is a possible alternative source. Biomass describes any organic material produced by living things (e.g. plant matter, animal waste, and organic refuse) and therefore is a renewable resource. The main advantage of biomass is that, when it combusts, it doesn't release any new carbon dioxide into the atmosphere. Rather, it recycles the existing carbon dioxide (when biomass is burnt, the carbon dioxide is taken in by living things and so becomes biomass again). Problems associated with biomass are the increased need to allocate farmland to produce the biomass, which increases soil erosion, salinity and over use of water resources. The cost of production for biomass is also quite substantial.

Solutions to these issues include the use of genetically modified crops to reduce the effects of producing biomass on the environment. Similarly, improvements to the current methods of producing ethanol (a fuel necessary to create biomass) will reduce costs of production.

Turning to the dot point specifically, let's list a few environmental, economic, and sociocultural implications of obtaining and using hydrocarbons.

Environmental impacts	**Economic impacts**	**Sociocultural impacts**
Extracting fuels from the Earth requires large, open-cut mines. This is hugely destructive to the natural environment.	Extracting fuels from the Earth has lead to the 'mining boom' in Australia, and has benefited the Australian economy for decades.	Mining has had devastating effects on many Aboriginal and Torres Strait Islander communities throughout the country. Since Mabo, the rightful owners of Australian land have been able to claim native title rights, unless these native title rights have been extinguished. In many cases, the sale of land to mining corporations and the subsequent mining activity extinguishes access to native title.
Burning fossil fuels produces greenhouse gases, which contributes significantly to climate change. The effects of climate change need not be discussed here; obviously, they are catastrophic.	Overreliance on the fossil fuel industry has stifled economic growth in many sectors of the Australian marketplace, most obviously in the renewable energy market.	We have become utterly dependant on energy generated by the extraction of organic fuels from the Earth. The value of these fuels has resulted in wars and sanctions globally.

Topic 3

Products of Reactions Involving Hydrocarbons

SYLLABUS :
Inquiry question: What are the products of reactions of hydrocarbons and how do they react?

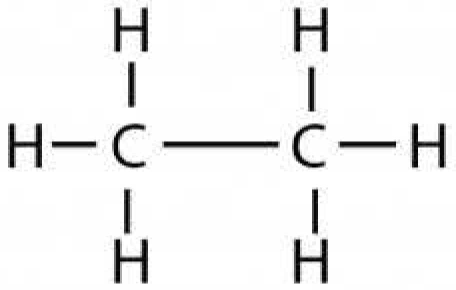
3.1 Unsaturated hydrocarbons

SYLLABUS :
Investigate, write equations and construct models to represent the reaction of unsaturated hydrocarbons when added to a range of chemicals, including but not limited to:

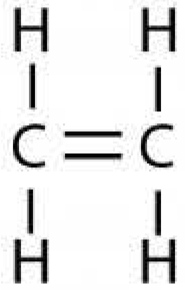
- Hydrogen (H_2)
- Halogens (X_2)
- Hydrogen halides (HX)
- Water (H_2O)

Recall from earlier that unsaturated hydrocarbons are defined as hydrocarbons with additional atoms that could possible bind to the compound; in other words, there must be **double bonds** (or triple bonds, etc.) present **between carbon atoms.** Let's quickly look at the following two compounds:

Ethane



Ethene



Of the two compounds (ethane and ethene), the first is saturated and the second is unsaturated. This is because the double bond could 'flip' out and bond to other atoms. Conversely, ethane would need to 'get rid of' a hydrogen atom in order to bond with another element.

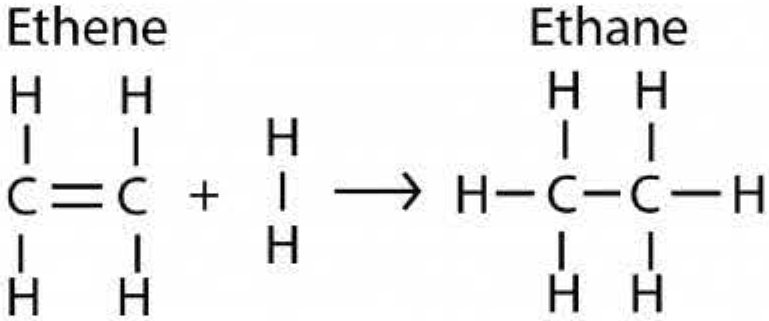
When a substance is reacted with an unsaturated hydrocarbon, resulting in the bonds 'flipping out' and additional elements being added, we call such a reaction an **addition reaction.**

3.1.1 Unsaturated hydrocarbons and hydrogen

When we add hydrogen gas $(H_{2(g)})$, unsaturated hydrocarbons will react to incorporate the additional hydrogen atoms into the molecule. Thus, it will become saturated. In chemical terms, this reaction will be:

$$C_2H_{4(g)} + H_{2(g)} \longrightarrow H_2C_{6(g)}$$

Structurally, this reaction would look like:



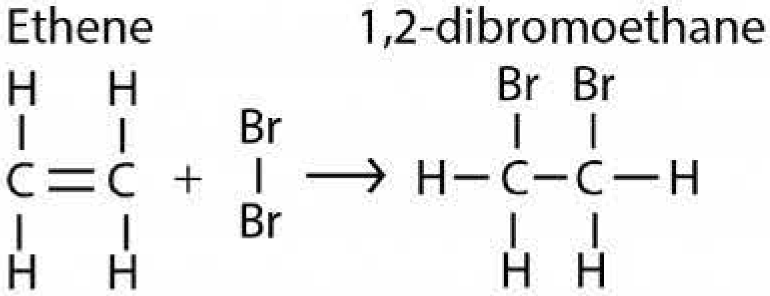
We call any reaction in which hydrogen gas is added to another compound **hydrogenation.**

3.1.2 Unsaturated hydrocarbons and halogens

When we add some sort of halogen $(X_{2(g)})$, unsaturated hydrocarbons will react to incorporate the additional halogen into the molecule. Since we're investigating the halogenation of hydrocarbons, we may as well look to a particularly interesting variety of halogenation: the addition of bromine.

Alkanes and alkenes tend to be colourless. However, when bromine atoms react with alkenes, the resulting molecule is likely to change colours. Thus, we can use bromine to see whether alkenes are present in a given region!

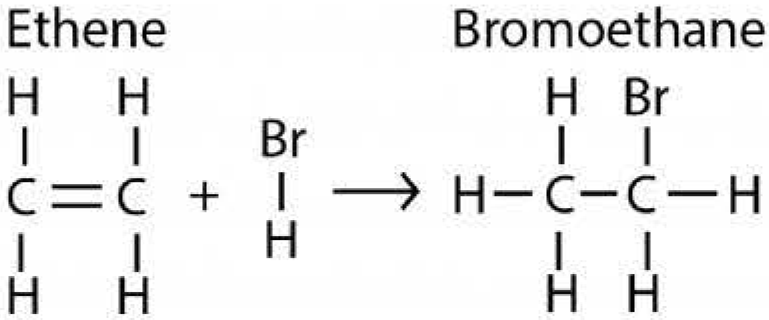
The addition of a halogen to an unsaturated hydrocarbon works in the same way as the previous example:



We can actually observe this reaction take place in a given vessel. The ethene will likely be gaseous, or colourless, whereas the bromine water (a brownish liquid) will become more and more clear as the reaction moves forward.

3.1.3 Unsaturated hydrocarbons and hydrogen halides

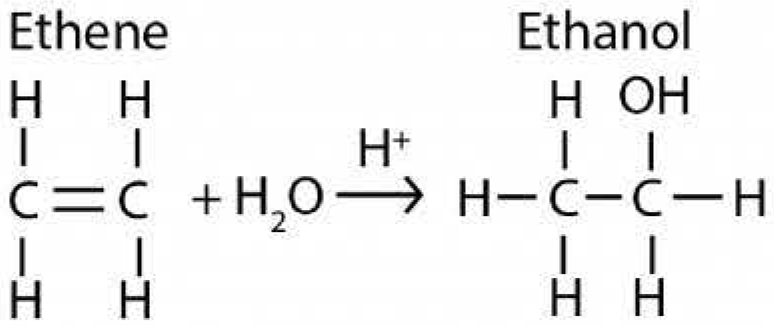
Hydrogen halides (HX) will also react with unsaturated hydrocarbons. Predictably, this reaction looks like the following:



Notice that, even if we flipped where the hydrogen or bromine atoms went, we would still have the same organic substance. Remember that a molecule, vertically or horizontally inverted, is still the same molecule!

3.1.4 Unsaturated hydrocarbons and water

Water will only really react with unsaturated carbons under acidic conditions. Under such conditions, the water and unsaturated hydrocarbons will undergo an addition reaction, producing an alcohol.

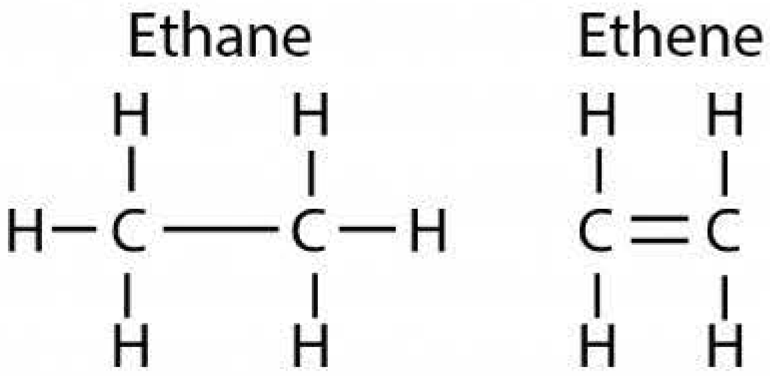


Key Point :
Unsaturated hydrocarbons undergo addition reactions by 'opening' double bonds and incorporating the additional molecule.

3.2 Saturated hydrocarbons

Syllabus :
Investigate, write equations, and construct models to represent the reactions of saturated hydrocarbons when substituted with halogens.

We know what an unsaturated hydrocarbon is, so it should be pretty clear what a saturated hydrocarbon is! Using the same example above:



...we can see that the first molecule (ethane) is saturated, as it does not contain any double bonds capable of accepting additional atoms. However, ethene is unsaturated, as its bonds could 'flip out' to incorporate any of the molecules looked at in the previous section.

Can saturated hydrocarbons undergo reactions? Clearly, they cannot undergo addition reactions, because part of the definition of addition reactions is that no extra product is left over, other than the now saturated hydrocarbon. In other words:

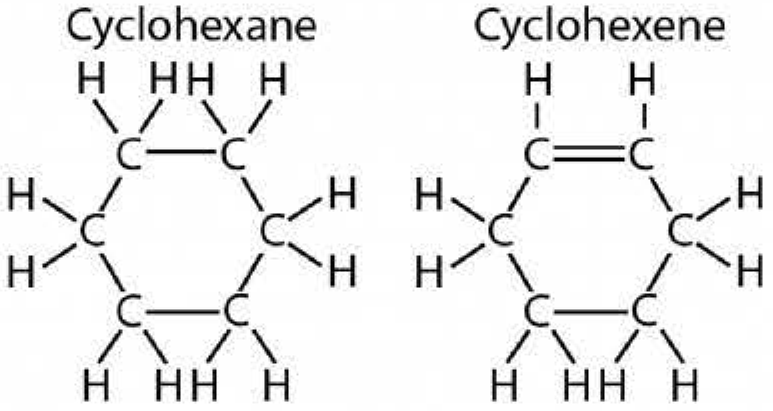
$$A + B \longrightarrow C$$

This can't work for saturated molecules as any reaction would have to **replace** hydrogen atoms and thus 'eliminate' them. Thus, a reaction would look more like:

$$A + B \longrightarrow C + D$$

We call these reactions **condensation reactions,** and they certainty do occur. However, they tend to require some sort of **catalyst,** such as UV light.

A good example is the comparison of cyclohexane and cyclohexene. You'll remember cyclohexane from a previous dot-point. Cyclohexene is the same molecule, but it contains a double bond somewhere in its structure. We look at cyclohexane and cyclohexene because they are liquid at room temperatures, unlike many of the other alkane/alkene pairs.



Cyclohexane is saturated (i.e. doesn't contain any double bonds), but cyclohexene *does* contain double bonds and is thus unsaturated. Let's perform a mock-experiment to see what happens when we add bromine water to these substances.

SAMPLE :

Aim: To compare the reactivities of saturated and unsaturated molecules.

Method:

1. 2 mL of cyclohexene were placed in two test tubes, and 2 mL of cyclohexane were placed in another two. 1 mL of bromine water was added to each test tube.
2. One of each type were covered with foil to show the effect of UV light.
3. Results (i.e. change of colour) were observed and recorded.

Results: The cyclohexane undergoes a substitution reaction forming two products (1-bromocyclohexane and hydrogen bromide), however this requires a UV catalyst (i.e. discolours in the test tube without foil after being placed in sunlight). The cyclohexene quickly decolourised the bromine water both with and without the UV catalyst. This is because cyclohexene underwent an addition reaction forming only one product (1,2-dibromocyclohexane).

Alkane: $C_6H_{12(aq)}+Br_{2(aq)} \xrightarrow[\text{strong UV}]{} C_6H_{11(aq)} + HBr_{(aq)}$

Alkene: $C_6H_{10(aq)}+Br_{2(aq)} \longrightarrow C_6H_{10(aq)}Br_{2(aq)}$

The experiment demonstrated the comparative reactivities of an alkene and its corresponding alkane. It demonstrated that the reactivity was higher in the alkene than the alkane when bromine water was added. This supports the theory that alkenes are more reactive because of their double bond. In particular, it shows that saturated molecules do undergo condensation reactions, although it is more difficult to make these move forward than for the equivalent addition reaction.

Note: the importance of using cyclohexene and cyclohexane is that they have the same chemical composition (except for the double bond) and they are both liquid at room temperature.

Topic 4

Alcohols

SYLLABUS :
Inquiry question: How can alcohols be produced and what are their properties?

4.1 Structure

SYLLABUS :
Investigate the structural formulae, properties and functional group including:

- Primary
- Secondary
- Tertiary alcohols

We've already looked at primary, secondary, and tertiary alcohols in a previous dot point. For a refresher, check page 48.

In terms of the properties of these alcohols, it is important to note that alcohols are **polar,** due to the presence of the hydroxide group (which is very negative when compared with the hydrogen functional groups). Thus, molecules will 'line up' in solution by dipole-dipole interactions, in which the negative hydroxide groups will be attracted to the positive hydrogen groups. The boiling point of alcohols is therefore very high, as the intermolecular forces are large.

Often, the boiling points of secondary and tertiary alcohols will be lower than for the equivalent primary alcohol. This is because the hydroxide group is more free to interact with other molecules, given that it isn't being covered by the many functional groups present in secondary and tertiary alcohols. As alcohols are polar, they are **very soluble** in water. You'll recall from earlier, and the Year 11 course, that **like-dissolves-like.** So, as water is polar, it is able to dissolve alcohols very well.

4.2 Properties

SYLLABUS :
Explain the properties within and between homologous series of alcohols with reference to the intermolecular and intramolecular bonding present.

The homologous series of alcohols is made up of the following:

Name	**Chemical formula**
Methanol	CH_3OH
Ethanol	C_2H_5OH
Propanol	C_3H_7OH
Butanol	C_4H_9OH
Pentanol	$C_5H_{11}OH$
Hexanol	$C_6H_{13}OH$
Heptanol	$C_7H_{15}OH$
Octanol	$C_8H_{17}OH$

Hopefully you've picked up the pattern – alcohols have the chemical formula $C_nH_{2n+1}OH$.

There are also a few properties worth mentioning. Alcohols are colourless liquids which are neutral in solution. Their reactions are similar no matter which alcohol you investigate, although their rate of reaction and reactivity may differ drastically. All alcohols are polar, and will have intermolecular hydrogen bonds present. They will thus have high boiling points.

Where the alcohols within a homologous series will differ is in the energy released upon burning. As the size of the molecule increases, the enthalpy of combustion will increase in magnitude per mole. So, when you burn one mole of octanol, you will get *much* more energy than burning one mole of methanol.

4.3 Equations and reactions

SYLLABUS :
Write equations, state conditions, and predict products to represent the reactions of alcohols, including but not limited to:

- Combustion
- Dehydration
- Substitution with HX
- Oxidation

4.3.1 Combustion

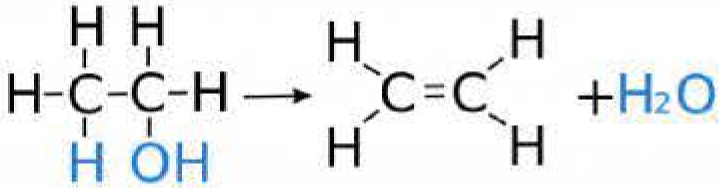
Combustion of alcohols is an extremely important reason for their production. Petrol often contains 10% ethanol, because it combusts easily and completely. **Complete combustion** occurs when *all* of the products are either water or carbon dioxide. **Incomplete combustion** occurs when some of the products include other substances, such as soot or carbon monoxide. Basically, incomplete combustion will occur when there isn't enough oxygen in the reaction space. As alcohols contain oxygen atoms within their chemical structure, they tend to combust completely.

$$C_2H_5OH_{(aq)} + 3O_{2(g)} \longrightarrow 2CO_{2(g)} + 3H_2O_{(l)}$$

To combust anything, you need to overcome some **initial activation energy,** typically by exposing the ethanol to a flame. Then, the reaction will go forward and produce the products above!

4.3.2 Dehydration

Dehydration is exactly what is sounds like – the 'removal' of a water molecule from the chemical structure of an alcohol. Remember that a water molecule is made up of two hydrogen atoms, and one oxygen atom. The easiest way to remove these substances from alcohols is to remove one random hydrogen atom, and the hydroxide functional group. However, to do this we will need a catalyst!



The catalyst we use is concentrated sulphuric acid, which lowers the activation energy of the dehydration of ethanol.

The forward reaction is exothermic and so the temperature of the system must not be too high (preferably 180°C) otherwise the yield of ethene is reduced. This catalyst also absorbs the water formed to stop the reaction from reversing (recall Le Chatelier's principle!).

$$CH_3CH_2OH_{(g)} \xrightarrow[\text{concentrated } H_2SO_4]{} C_2H_{4(g)} + H_2O_{(l)}$$

$$\text{Alcohol} + \text{Appropriate catalyst} \longrightarrow \text{Alkene} + \text{Water}$$

4.3.3 Substitution with HX

When alcohols react with hydrogen halides, the end product will no longer be an alcohol. Honestly, this area of chemistry is a bit complicated, and there are a number of steps that go into this reaction process. It seems very unlikely that you need to understand the structural changes which a molecule will undergo before reaching a final state, so here we will just give a brief overview of the reaction.

This reaction will go forward in acidic conditions. A substitution reaction will occur, producing water and a halogenated organic molecule.

$$C_2H_5OH_{(aq)} + HX_{(aq)} \xrightarrow[\text{acidic conditions}]{} C_2H_5X_{(aq)} + H_2O_{(l)}$$

$$\text{Alcohol} + \text{Hydrogen halide} + \text{Appropriate catalyst} \longrightarrow \text{Halogenated hydrocarbon} + \text{Water}$$

4.3.4 Oxidation

- **Primary alcohols:** when alcohols are oxidised, they can form one of two products. They may form aldehydes or, if there are sufficient oxidising agents, they may form carboxylic acids. Recall from the Year 11 course that we say that something has oxidised when it loses electrons. So, an oxidising agent is a chemical compound that causes another compound to oxidise; thus, the agent itself must be reduced.

  KEY POINT :
  Alcohols may be **partially oxidised** to form aldehydes (i.e. if the oxidising agent is the limiting reagent).
  Alcohols may be **fully oxidised** to form carboxylic acids (i.e. if the alcohol is the limiting reagent).

- **Secondary alcohols:** these will only oxidise to form ketones. The reaction conditions (i.e. the substance in excess) will not actually change the products!
- **Tertiary alcohols:** these won't oxidise at all! This makes for an easy pattern to remember!

KEY POINT :
When **primary** alcohols oxidise, they can form **aldehydes** (partial oxidisation) or **carboxylic acids** (total oxidisation).

When **secondary** alcohols oxidise, they form **ketones.**

Tertiary alcohols do **not** oxidise.

4.4 Alcohol production

SYLLABUS :
Investigate the production of alcohols, including:

- Substitution reactions of halogenated organic compounds
- Fermentation

We've already touched upon the substation reactions involved when alcohols come into contact with hydrogen halides. The halide will replace the hydroxide functional group, producing water.

Now, we will look specifically at the fermentation of glucose to form ethanol, since that is the most important fermentation reaction in modern chemistry.

Fermentation is the process by which yeast 'eats' and decomposes some sort of biomass, such as glucose. The reaction describing the formation of ethanol by fermentation is:

$$C_6H_{12}O_{6(aq)} \xrightarrow[\text{zymase (yeast)}]{} 2C_2H_5OH_{(aq)} + 2CO_{2(g)}$$

...where $C_6H_{12}O_6$ is glucose.

This fermentation process must be completed under certain conditions, including:

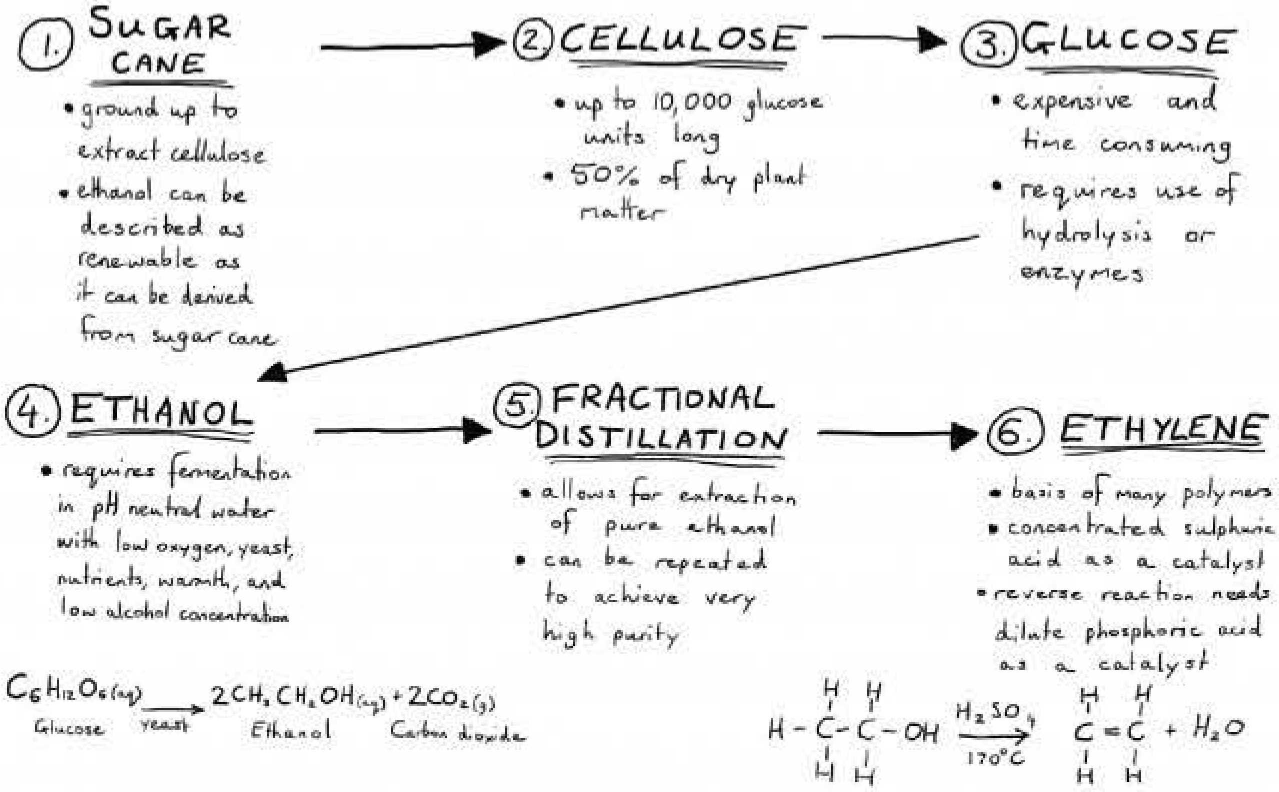
- pH neutral water, so that the enzymes are not denatured
- A micro-organism such as yeast which contains enzymes to catalyse the reaction
- Low O_2 concentration in an ambient environment
- Nutrients for yeast to use (e.g. glucose)
- Alcohol concentration of <14%
- A warm temperature (about 30°C)

Bubbles of carbon dioxide are evident and measurable by mass change over time. Yeast can produce ethanol concentration up to approximately 14–15% before the yeast itself is killed. To produce higher alcohol concentrations, the product must be distilled.

SYLLABUS :

Compare and contrast fuels from organic sources to biofuels, including ethanol.

I will just quickly summarise the steps involved in producing fuel (ethanol) from biomass, and the advantages and disadvantages of such fuel and production processes in the following flowchart and table:



Advantages of ethanol	**Disadvantages of ethanol**
It reduces the consumption of petrochemicals which can be put to alternative use.	As a means of production, it requires fuel (e.g. to operate tractors) so its 'carbon neutrality' is questionable
It can be dehydrated into ethene to produce plastics.	Arable land is required to grow sugar cane, potentially reducing crop production.
Ethanol undergoes complete combustion.	There is increased engine wear as ethanol absorbs moisture and accelerates corrosion.
It is fairly inexpensive to produce.	It only produces around 66% of the energy of petrochemicals.
Widespread ethanol use could reduce greenhouse gas emissions.	Widespread use would require the modification of fuel lines and engines for ethanol concentrations greater than 15%.
It is renewable.	Ethanol has a higher flash point.

These past few dot points were a major part of the old syllabus, so if you want more detail than the summary provided here, you can look up the countless notes available online for the 'Production of Materials' module. However, the information included here should be all you need to know!

Topic 5

Reactions of Organic Acids and Bases

SYLLABUS :
Inquiry question: What are the properties of organic acids and bases?

5.1 Structure

SYLLABUS :
Investigate the structural formulae, properties and functional group including:

- Primary, secondary, and tertiary alcohols
- Aldehydes and ketones
- Amines and amides
- Carboxylic acids

Each of the chemical compounds listed above have been investigated in some depth already, but, we'll take a quick look at some of the properties before moving on. **Aldehyde and ketones** tend to have high boiling points (particularly when compared to esters and alkanes with a similar molecular mass) due to the carbon-oxygen double bond. However, their boiling points tend to be lower than alcohols with corresponding molecular masses. Aldehydes oxidise quickly to form carboxylic acids, but ketones are much less likely to oxidise. **Amines** tend to have lower boiling points than alcohols with a similar molecular weight. However, this is dependant on the type of amine (i.e. primary, secondary, or tertiary). Tertiary amines do not have hydrogen bonds, and so their boiling points are similar to standard hydrocarbons of the same molecular weight. Amines form hydrogen bonds with water, and so will readily dissolve in it. **Amides** tend to have high melting and boiling points. They hydrogen bond in the presence of water, and thus dissolve easily. **Carboxylic acids** tend to be colourless, and strongly hydrogen bond with each other resulting in a high boiling point. They are very soluble, but solubility decreases as the size of the carboxylic acid increases.

5.2 Properties

SYLLABUS :
Explain the properties within and between the homologous series of carboxylic acids, amines, and amides with reference to the intermolecular and intramolecular bonding present.

5.2.1 Carboxylic acids

Property	Relationship with inter/intramolecular forces
Solubility	Carboxylic acids tend to be quite soluble in water. Intermolecular hydrogen bonds are very strong dipole-dipole interactions, and so water molecules will 'break' the intermolecular bonds between carboxylic acids molecules. Carboxylic acids of up to four carbons in length (formic, acetic, propionic, and butyric acid) are completely soluble in water. Above that, the solubility decreases with length.
Boiling point	Because of the presence of hydrogen bonds between carboxylic acid molecules, it is quite difficult to 'rip' these intermolecular bonds apart. As such, carboxylic acids have quite high boiling points. Boiling points increase as the length of the carbon chain gets longer, as there are greater numbers of potential intermolecular hydrogen bonds present.

5.2.2 Amines

Property	Relationship with inter/intramolecular forces
Solubility	Amines tend to be quite soluble in water. Intermolecular hydrogen bonds are very strong dipole-dipole interactions, and so water molecules will 'break' the intermolecular bonds between amine molecules.
Boiling point	Because of the presence of hydrogen bonds between amine molecules, it is quite difficult to 'rip' these intermolecular bonds apart. As such, amines have quite high boiling points. Boiling points increase as the length of the carbon chain gets longer, as there are greater numbers of potential intermolecular hydrogen bonds present.
pH	Amines are basic substances. The pH will tend to increase as the size of the amine molecule increases (i.e. as the carbon chain gets longer).

5.2.3 Amides

Property	Relationship with inter/intramolecular forces
Solubility	Smaller amides tend to be soluble in water, due to intermolecular forces including hydrogen bonding.
Melting point	Amides tend to be solid at room temperature, as they have fairly high melting points. Hydrogen bonds between molecules (based on the lining up of hydrogen atoms with oxygen atoms) are very strong, and require a lot of energy to break.

5.3 Simple esters

SYLLABUS :
Investigate the production, in a school laboratory, of simple esters.

Whilst this is obviously a practical dot point to be covered in class, it's worth just briefly summarising some information about esters.

We already know about two basic homologous series: alkanols and alkanoic acids. Below is a brief summary:

- **Alkanols:** contain hydroxyl groups (–OH) and can be described by the formula $C_nH_{2n+1}OH$. They have a low level of hydrogen bonding, and a moderate boiling/melting point.

Number of C atoms	Name of corresponding alkanol
1	Methanol
2	Ethanol
3	Propanol
4	Butanol
5	Pentanol
6	Hexanol
7	Heptanol
8	Octanol

- **Alkanoic acids:** contain carboxylic acid groups (–COOH) and can be described by the formula $C_nH_{2n+1}OOH$ of $C_{n-1}H_{2n-1}COOH$. They have high levels of hydrogen bonding, and a high boiling/melting point.

Number of C atoms	Name of corresponding alkanol
1	Methanoic acid (or formic acid)
2	Ethanoic acid (or acetic acid)
3	Propanoic acid
4	Butanoic acid
5	Pentanoic acid
6	Hexanoic acid
7	Heptanoic acid
8	Octanoic acid

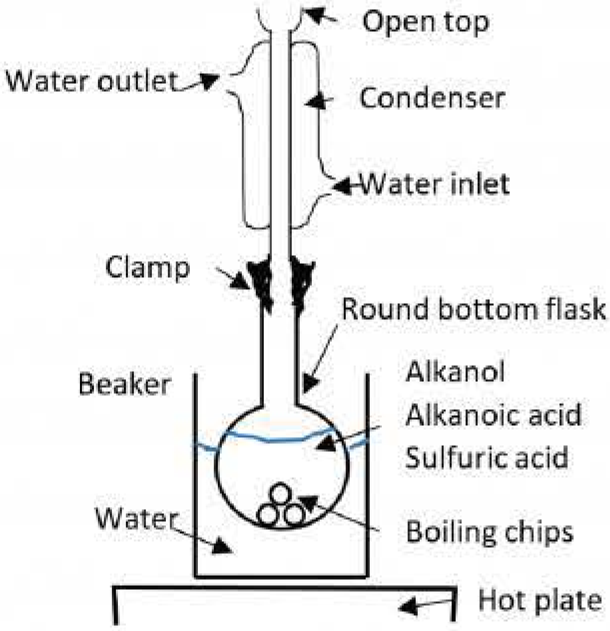
Now, we introduce a new chemical compound called an ester. Esters are the product of a reaction between alkanols and alkanoic acids. They are named based on the two reactants used, and takes the form **alkyl alkanoate.**

Alkanol	Alkanoic acid	Corresponding ester
Methanol	Hexanoic acid	Methyl hexanoate
Ethanol	Propanoic acid	Ethyl propanoate
Hexanol	Octanoic acid	Hexyl octanoate
Butanol	Propanoic acid	Butyl propanoate

Esters are created by the reaction between alkanols and alkanoic acids, catalysed by a strong acid such as concentration sulphuric acid. Water is also a product of the reaction.

Esterification typically takes a very long time, and a complete reaction is unachievable. However, using an acid catalyst, as described above, can speed up the process. It acts as a dehydrating agent, which will push the equilibrium to the right (i.e. towards the side of the products) by Le Chatelier's principle to make up for the change. This will increase the overall yield of the ester.

Esterification is a complex process, and requires a set up that allows for heating without bubbles being formed, condensation of vapour and pressure to be built without a loss of volatile liquid, all of which reflux achieves. As such the apparatus on the right would be set up.



Esters make up many smells and tastes associate with naturally occurring fruits and other foods. They are not also used for similar, synthesised purposes. Esters are produced to artificially change the flavour or smell of foods. They also operate as solvents and thinners in industrial settings, and as lubricants for engines.

KEY POINT :
Esters are the product of an **alkanol** and an **alkanoic acid.**

Concentrated acids are used as a catalyst.

5.4 Organic acids and bases

SYLLABUS :
Investigate the differences between an organic acid and an organic base.

Organic acids agree with the Brønsted-Lowry definition of acids that we looked at earlier; namely, they donate protons. Similarly, organic bases agree with the Brønsted-Lowry definition of bases; they accept protons.

Organic acids and bases have quite varying pHs, depending on their ability to ionise in solution (i.e. their strength) and their intramolecular bond strength.

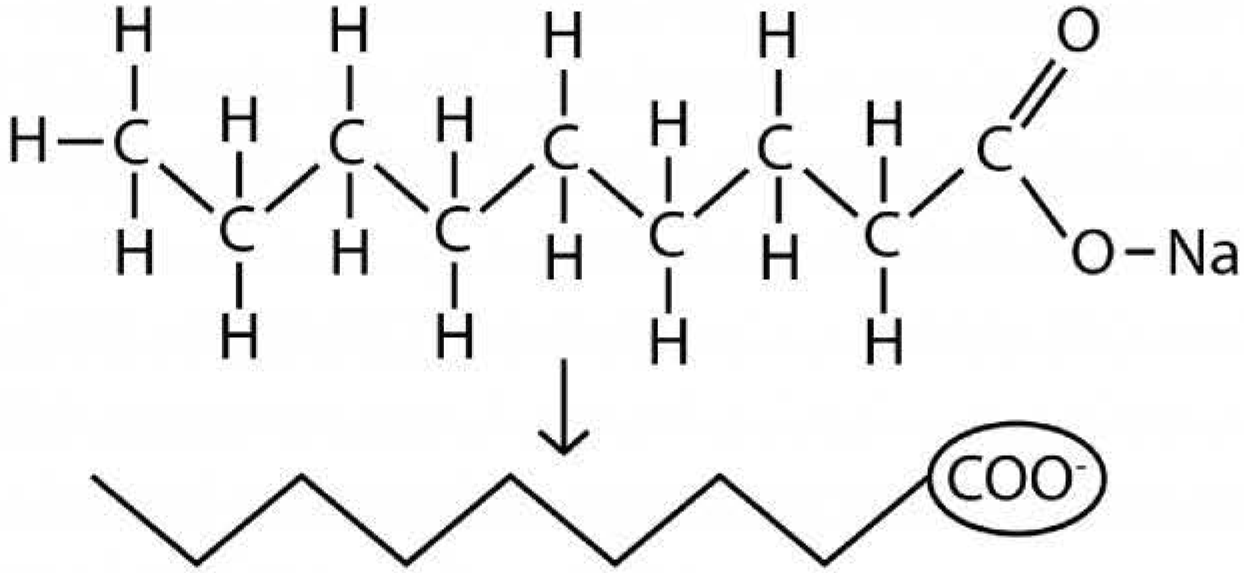
5.5 Soaps and detergents

SYLLABUS :
Investigate the structure and action of soaps and detergents.

You'll be well aware of the cleaning ability of soaps and detergents. But how does this work on a molecular level? Why are some substances good at cleaning, and others are not?

First, it's worth understanding what problem soaps and detergents are trying to solve; specifically, the removal of dirt from surfaces. Dirt itself is largely oily, and so the removal of oils is necessary to clean things. However, oil is non-polar, so as water is polar and 'like dissolves like,' we can't clean dirt using just water.

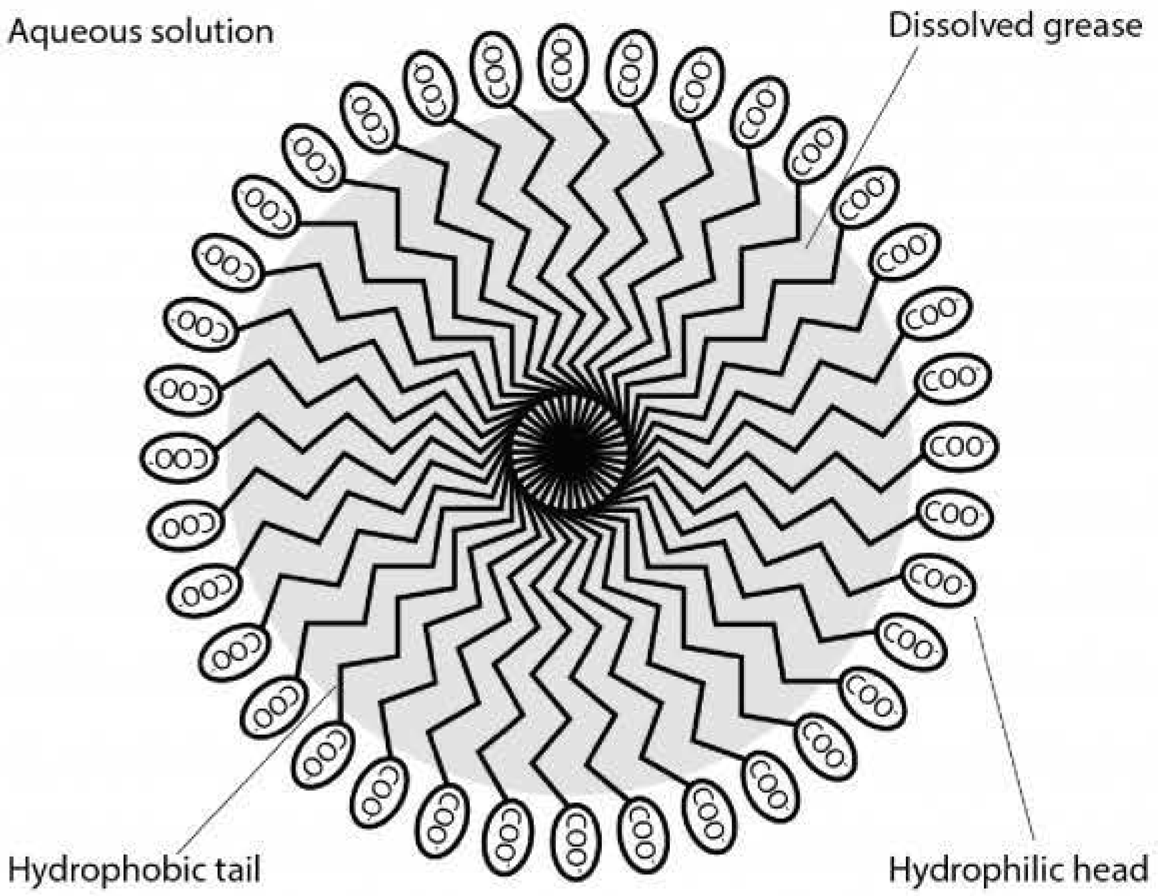
Soap itself is made up of long carboxylic acid chains, which typically include sodium or potassium salts. Let's investigate the structure on the basis of polarity.



The top structural diagram is the full form, outlining where each atom is placed and where the COO– side is. The bottom structural diagram is a simplified form, and is the standard way of writing out soaps and detergents. The left-hand side is known as the tail, and the right hand side is known as the head.

It's pretty clear that the tail is non-polar, and the head is polar. Like dissolves like, and so the polar head will dissolve in water, and the non-polar tail will not. We call the head **hydrophilic** because it 'likes' water and is attracted to water molecules due to its polarity. We call the tail **hydrophobic** because it 'dislikes' water and is repelled by it, meaning that it avoids interacting with water. Because the hydrophobic end is polar, that end will dissolve the oils we were talking about earlier in the dot point.

Now we get to the interesting part. Because the head attracts/is attracted to water (note that this isn't strictly accurate terminology, but it explains what's happening conceptually!), and the tail attracts/is attracted to the oils (and is repelled by water), the soap molecules form a microscopic ball known as a **miscelle.** In the middle of the miscelle will be some oily dirt, and the aqueous solution will be outside the miscelle. These miscelles thus remove oily grease from the surface of an item, cleaning it over time!

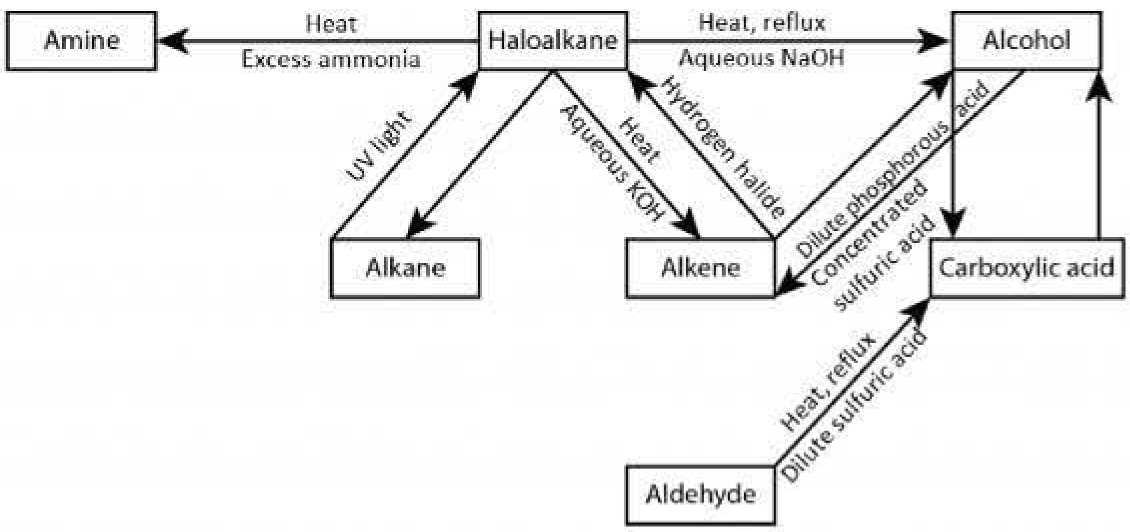


Looks pretty cool, right? These miscelles form, extract grease and dirt from the items being cleaned, and get washed away!

5.6 Chemical synthesis

SYLLABUS :
Draft and construct flow charts to show reaction pathways for chemical synthesis, including those that involve more than one step.

Chemical synthesis is the utilisation of known chemical properties to form particular reactants by particular chemical processes. So, for example, we already know that we can synthesise alkenes from alcohols, or alkanes from alkenes. All of these processes are easier to conceptualise when you construct flow charts showing these reaction pathways! There are heaps and heaps of potential reaction pathways, so what follows are a collection of important ones.



Topic 6

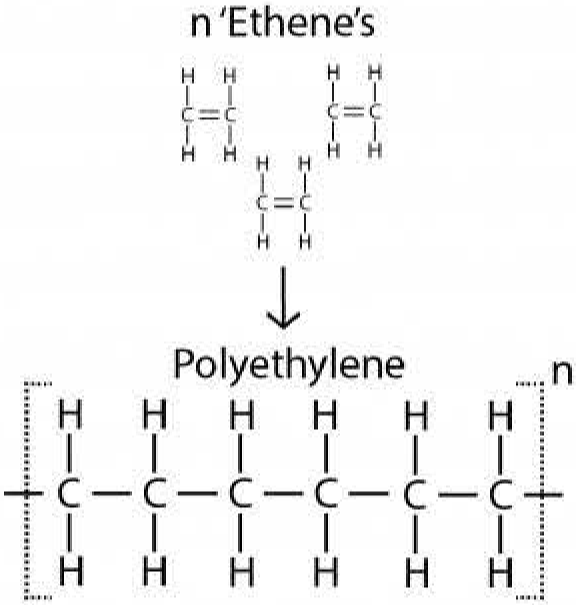
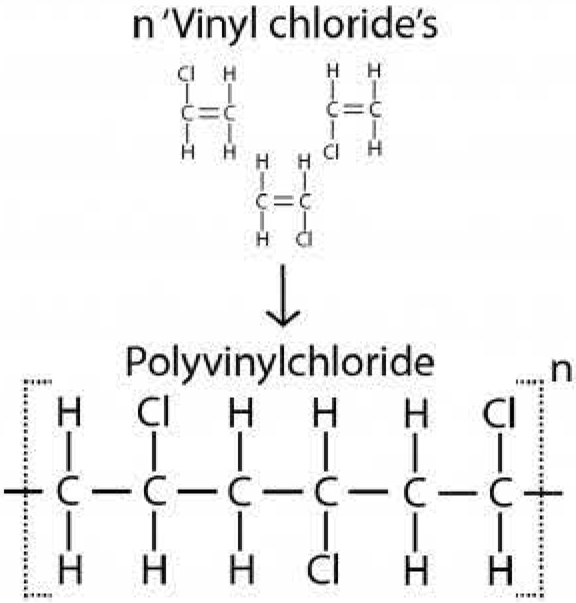
Polymers

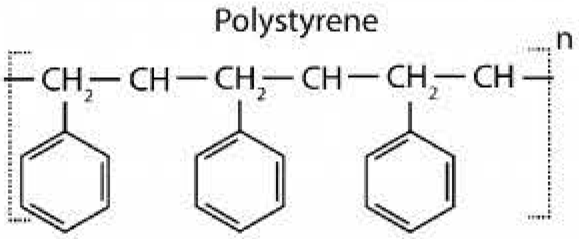
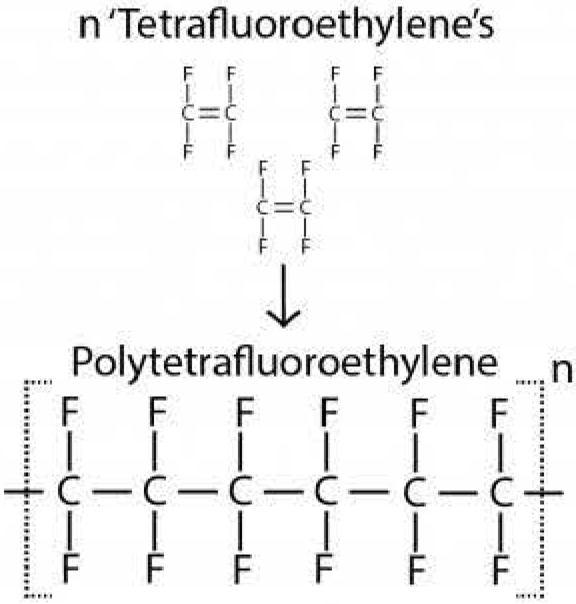
SYLLABUS :
Inquiry question: What are the properties and uses of polymers?

SYLLABUS :
Model and compare the structure, properties and uses of addition polymers of ethylene and related monomers, for example:

- Polyethylene (PE)
- Polyvinyl chloride (PVC)
- Polystyrene (PS)
- Polytetrafluoroethylene (PTFE)

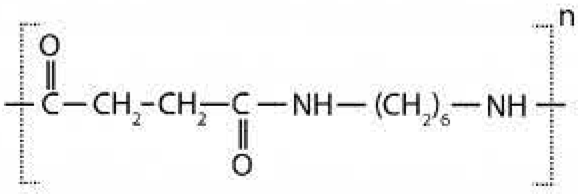
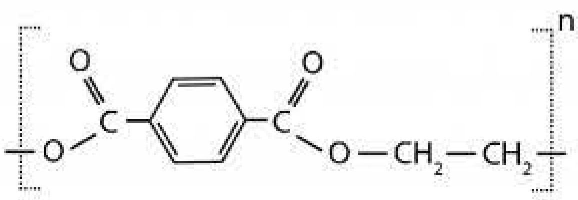
Whenever you draw a polymer (as all of the below are), make sure to draw at least three monomers!

Name	Structure	Properties	Uses
Polyethylene		There are two types of polyethylene: high density polyethylene (HDPE), and low density polyethylene (LDPE). **LDPE:** Boiling point ~115 degrees, very flexible, not very strong, quite tough and pliable, lower density. **HDPE:** Boiling point ~135 degrees, rigid, strong, higher density.	**LDPE:** Glad wrap, plastic bottles, plastic bags. **HDPE:** Piping, buckets, insulated for wires.
Polyvinyl chloride		Very chemically stable, fire-resistant, durable, mechanical stability, mouldable, rigid, low flexibility, water resistance, low melting point.	Pipes, fittings, squeeze bottles, electrical wire coating, building materials.

Polystyrene		Rigid due to the benzene group, stiff, thermoplastic, lightweight, brittle, hard.	CD cases, insulation, toys and utensils, thermal cups, tool handles.
Polytetrafluoroethylene		More commonly known as Teflon, PTFE is flexible, chemical resistant, thermal resistance, stick-proof and electrically resistant. Very high boiling point, chemically inert.	Coatings of utensils, pans and building materials. Useful in medical equipment.

SYLLABUS :
Model and compare the structure, properties, and uses of condensation polymers, for example:

- Nylon
- Polyester

Name	Structure	Properties	Uses
Nylon	 $\left[C(=O)-CH_2-CH_2-C(=O)-NH-(CH_2)_6-NH \right]_n$	Nylon is strong and elastic and is somewhat water resistant. It is strong and resilient, and heat resistant.	Nylon is used in clothing items such as stockings, elastics, socks, active wear etc.
Polyester		Polyester is a hydrophobic polymer that is very durable, chemical resistant, elastic and strong	Polyester is used in clothing items, such as polyester suits, shirts, pants etc. Polyester is also commonly used to make bottles.

KEY POINT :
Addition polymers form when a number of the same monomers bond together.

Condensation polymers form when different monomers bond together in regular patterns.

Part IV

Module 8: Applying Chemical Ideas

Topic 1

Analysis of Inorganic Substances

SYLLABUS :
Inquiry question: How are ions present in the environment identified and measured?

1.1 Monitoring the environment

SYLLABUS :
Analyse the need for monitoring the environment.

Monitoring the environment is absolutely crucial for the continued survival and growth of humanity. Here, I'll just quickly summarise a few possible aspects of the environment that require monitoring, as well as a brief explanation as to how we monitor that region of the environment.

1.1.1 Waterways

We rely on waterways to support crop growth, maintain flora and fauna, and feed into our water treatment plants and dams for our domestic and industrial use. The composition of waterways need to be monitored quickly and efficiently, as there are often contaminants present within them.

Factories near waterways often leach poisonous heavy metals, such as lead, into waterways. This will enter the soil, kill animals that drink from the waterway, and can cause major damage to human life that is exposed to even tiny levels of heavy metals. We monitor heavy metals in waterways using **Atomic Absorption Spectroscopy,** or AAS for short. AAS vaporises water, and uses light absorption to measure the quantity of heavy metal present.

The acidity of waterways also affects crop growth and aquatic life. Debris and run off can affect the acidity of waterways quite rapidly. Simple universal indicators can be used to detect changes in pH.

1.1.2 Atmosphere

The composition of our atmosphere is naturally vital to the continued wellbeing of humanity. In particular, atmospheric pollutants such as carbon monoxide can cause debilitating lung problems, which are often experienced most severely by children and the elderly. Natural environments can change rapidly based on the pollutants they are exposed to.

High levels of sulphuric oxides and nitrous oxides can result in the synthesis of acid rain. These oxides are often produced on the combustion of fossil fuels. They form acid rain through the following chemical reactions:

$$SO_{3(aq)} + H_2O \longrightarrow H_2SO_{4(aq)}$$

$$2NO_{2(g)} + H_2O_{(l)} \longrightarrow HNO_{3(aq)} + NHO_{2(aq)}$$

Acid rain has devastating effects on society. Firstly, it increases erosion of land, whether that be in the form of landscape erosion, building erosion or monument erosion. It also destroys plants and croplands, increasing the acidity of the soil. Finally, it acidifies waterways, which can poison human and marine life alike.

1.1.3 Ozone crisis

The ozone crisis was the last major ecological disaster before climate change. Ozone is a molecule which is produced naturally from the breakdown of nitrogen oxides by sunlight. Ozone molecules protect us for UV rays, and their destruction can cause:

- Greater rates of skin cancer
- Eye irritation which leads to retina damage
- Breaking covalent bonds which leads to gene mutations
- Decreases in plant growth
- Negative effects in aquatic organisms

The production of CFCs, or chlorofluorocarbons, destroyed ozone molecules by the following chemical reactions.

First, CFCs dissociated in UV radiation to form a **chlorine free radical** (Cl.):

$$CClF_{3(g)} + UV \longrightarrow Cl._{(g)} + CF_{3(g)}$$

Secondly, these extremely reactive free radicals interact with ozone to produce a chlorine oxide radical and an oxygen molecule.

$$Cl._{(g)} + O_{3(g)} \longrightarrow ClO._{(g)} + O_{2(g)}$$

Finally, this reacts with an oxygen free radical to form a chlorine free radical and an oxygen molecule.

$$ClO._{(g)} + O._{(g)} \longrightarrow Cl._{(g)} + O_{2(g)}$$

Thus, the overall reaction is:

$$O_{3(g)} + O._{(g)} \longrightarrow 2O_{2(g)}$$

You can see that the CFC acts as a catalyst, as it does not take part in the overall reaction. You can also see that CFCs destroy ozone, and thus are damaging to our natural environment. Because of CFC use, the ozone layer in our atmosphere decreased by 50% in Antarctica.

It took the 1974 Nobel prize winner to demonstrate the above set of reactions; they proved that the need to monitor environmental change was crucial.

1.2 Qualitative investigations

SYLLABUS :
Conduct qualitative investigations – using flame tests, precipitation, and complexation reactions as appropriate – to test for the presence in aqueous solution of the following ions:

- **Cations:** barium (Ba^{2+}), calcium (Ca^{2+}), magnesium (Mg^{2+}), lead(II) (Pb^{2+}), silver ion (Ag^{+}), copper(II) (Cu^{2+}), iron(II) (Fe^{2+}), iron(III) (Fe^{3+})
- **Anions:** chloride (Cl^{-}), bromide (Br^{-}), iodide (I^{-}), hydroxide (OH^{-}), acetate (CH_3COO^{-}), carbonate ($CO^{2-}{}_3$), sulphate ($SO_4{}^{2-}$), phosphate ($PO_4{}^{3-}$

1.2.1 Flame test

Flame tests can be used to identify certain cations, based on a characteristic colour produced when the particular ion is placed into a flame. The following is a summary of the colour you would expect when placing a particular ion into a Bunsen burner flame.

Salt	Cation	Flame colour	How to remember
Barium nitrate	Ba^{2+}	Apple green	An apple a day keeps the doctor away, so you won't have to barium!
Calcium nitrate	Ca^{2+}	Brick red	Calcium keeps the bones as strong as bricks!
Copper nitrate	Cu^{2+}	Blue/green	A copper's (policeman's) uniform is usually blue!
Iron nitrate	Fe^{2+}/Fe^{3+}	Sparkly yellow	Iron Man has a yellow helmet!
Lead nitrate	Pb^{2+}	White	If you use lead pencil, you won't need whiteout!

It is worth noting that you should not perform Iron nitrate or lead nitrate flame tests, as the oxides of lead and iron are toxic.

1.2.2 Precipitation reactions

We've already discussed these in-depth, so check page 25 for a refresher. First, figure out what the general rule (NAGSAG) for a given substance is. Then, check if any exceptions (PMS, CASTROBeAR) apply!

1.2.3 Complexation reactions

Complexation reactions are new to the HSC syllabus, and you likely won't need to know these ones off-by-heart. Complexation reactions occur when two substances combine to form a complex of some kind, which typically holds a charge, and results in an obvious colour change. Such a colour change indicates the type of complex that has been formed.

These complexes generally require a very high concentration of at least one of the reagents before any product is formed. You'll be given any information needed to carry out a complexation reaction; all that will be left to you is the standard molar calculations we've used in many other sections!

1.3 Testing for ions

SYLLABUS :
Conduct investigations and/or process data to determine the concentration of coloured species and/or metal ions in aqueous solution, including by not limited to, the use of:

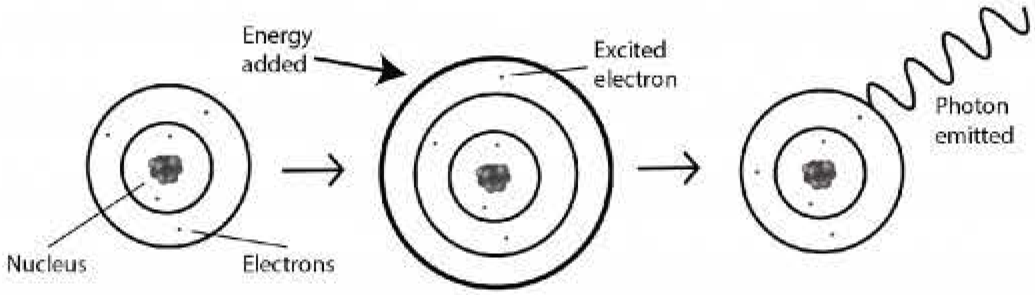
- Colourimetry
- Ultraviolet-visible spectrophotometry
- Atomic absorption spectroscopy

Here, I'll try to summarise the theory behind the three methods above. You may get to actually perform some of these in class!

1.3.1 Colourimetry

Colourimetry is used to test the concentration of different coloured liquids. Many chemical compounds form particular colours in solution (e.g. copper ions are blue). To test of the concentration of these substances, its important to remember what colour actually is!

Why do different elements emit different wavelengths of light? It all has to do with the discrete energy levels associated with an atom.



When heat is added to the atom, some of the electrons move to an 'excited' state. This excited state is in an energy level higher than its original state. For instance, an electron might move from n=1 to n=2. Then, the electron will want to move back to its initial state (i.e. from n=2 to n=1). To do that, it has to release energy. It does this in the form of a photon, which is a particle of light. The energy (and therefore the wavelength/colour) of the light is dependant on the distance between energy levels of the particular atom. Different atoms have different distances, and so emit different colours!

KEY POINT :
Light is emitted when excited electrons move back to their ground state.

Each element will release a 'characteristic' wavelength of light. Colourimetry exploits this principle, by shining light through chemical substances and measuring the absorbance and release of particular wavelengths of light. From this information, the concentration of particular substances can be determined!

1.3.2 Ultraviolet-visible spectrophotometry

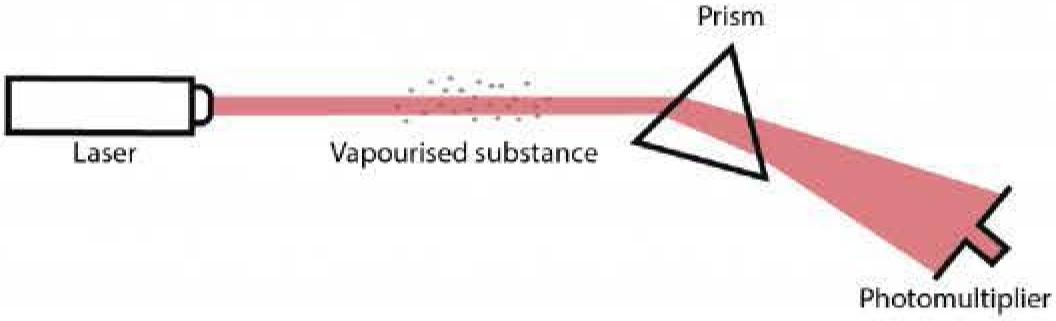
This method of quantitative analysis is similar to that above. By measuring the absorbance of particular wavelengths of light, and correlating these wavelengths with the bond energies of particular compounds, it is possible to gain information (to quite a high level of accuracy) about the concentration of particular ions. This method uses light in the visible spectrum, and in slightly adjacent sections.

1.3.3 Atomic absorption spectroscopy (AAS)

AAS uses light to discriminate between very small quantities of known substances. It operates in the following way:

1. A sample is aspirated and sprayed into a flame.
2. The vaporised substance has light shone through it.
3. A detector behind the substance measures the intensity of light absorbed, compared with the intensity of light let through.
4. A computer records the absorbance in units comparing the light absorbed against the calibration curves of the standard (i.e without the aspirated substance).

This process allows concentration of heavy metals to be detected down to ppm (parts per million).



This is particularly useful as our body requires small amounts of **trace elements,** such as zinc and magnesium, for efficient action of enzymes. AAS can therefore be used to understand these trace elements and their concentration in the human body, which can aid in medical treatment.

Trace element	Function/effect in living things
Zinc	Needed for metabolism of amino acids and in energy production
Manganese	Required to reduce risk of blood clotting and involved in carbohydrate and fat metabolism
Copper	Needed for the production of enzymes involved in oxidation reactions
Cobalt	Enables immune functions to operate efficiently
Magnesium	Needed to maintain good muscle condition

Topic 2

Analysis of Organic Substances

SYLLABUS :
Inquiry question: How is information about the reactivity and structure of organic compounds obtained?

2.1 Qualitative investigations

SYLLABUS :
Conduct qualitative investigations to test for the presence in organic molecules of the following functional groups:

- Carbon–carbon double bonds
- Hydroxyl groups
- Carboxylic acids

This dot point is largely a practical one, however we can briefly look at some tests which would determine (qualitatively) whether the above functional groups are present in organic molecules.

2.1.1 Carbon–carbon double bonds

We've discussed compounds with carbon-carbon double bonds quite extensively above. Recall that these compounds are called unsaturated, due to their double bond. The easiest way to test whether unsaturated molecules are present is to determine whether an addition reaction moves forward.

The addition of halogens, halogen halides, or even just hydrogen gas, can determine whether a substance is saturated or not. In particular, see the cyclohexane/cyclohexene experiment in the above dot points.

The most common tests for unsaturated hydrocarbons are:

- **Bromine test:** the addition of bromine water causing a change in colour
- **Baeyer's test:** the addition of alkaline $KMnO_4$ causing a change in colour

2.1.2 Hydroxyl groups

The most common tests for whether a substance contains a hydroxyl group, and therefore whether it is an alcohol, are:

- **Active metal test:** the addition of metals such as sodium, producing hydrogen gas
- **Ester test:** whether the substance forms an ester, as per the above discussion of esters
- **Ceric ammonium nitrate test:** the addition of ceric ammonium nitrate causing a change in colour
- **Acetyl chloride test:** the addition of acetyl chloride to form an ester and hydrogen chloride gas

2.1.3 Carboxylic acids

The most common tests for carboxylic acids are:

- **Litmus test:** turns blue litmus paper red
- **Sodium hydrogen carbonate test:** the addition of sodium hydrogen carbonate produces carbon dioxide
- **Ester test:** reflux with an alcohol to form an ester

2.2 Analysis tools

SYLLABUS :
Investigate the processes used to analyse the structure of simple organic compounds addressed in the course, including but not limited to:

- Proton and carbon-13 NMR
- Mass spectroscopy
- Infra-red spectroscopy

All of the below techniques are quite technical, and mostly based on physical chemistry which is beyond this course. In addition, to have a property understanding of the principles you would need to go into a fair bit of detail with respect to mathematical and physical equations, including some really tough electrodynamics. I'll try to keep each description brief, giving you a rough outline of how the processes work and what they are used for.

2.2.1 Proton and carbon-13 NMR

NMR, or nuclear magnetic resonance, is a hugely important chemical tool in identifying particular elements or compounds within a sample. The physics behind NMR is quite complicated, and I won't go into too much depth here. I'll just give a brief overview of NMR generally, and the difference between proton NMR and carbon-13 NMR.

Atoms, when subjected to a magnetic field, will act differently depending on their magnetic and electric properties. This is due to the 'magnetic moment' and 'angular momentum' of the nucleus of elements. Since each nucleus is unique, each element will exhibit unique properties!

Each atom will have a unique 'resonance frequency', however this resonance frequency is dependent on the strength of the magnetic field induced by the NMR machine. This means that a machine can use a varying magnetic field, and thus locate where particular substances are within a sample based on their response frequency. When a magnetic field is applied to a substance, the proton in the nucleus will tend to 'line up' with that field. However, there is a small chance that they will exactly 'oppose' the field by spinning in the opposite direction; this is known as the 'anti-parallel' state. By determining the rate at which the nucleus 'switches' its spin state, and how likely it is to do so, we can get a very good understanding of what kind of nucleus it is.

The most commonly used nucleus that is investigated is the proton, or the Hydrogen ion (recall that they are the same thing!). This nucleus is very sensitive to NMR signals, and provide a clear, characteristic 'peak' in NMR outputs. This is the atom used in MRI machines, which work by very similar principles to NMR. These peaks can be determined very quickly, and can provide a lot of important information.

Carbon-13 is also widely used as a detective atom, even though it only occupies about 1% of atoms on earth. It is a very stable molecule, but the measurements take an extremely long time to take. Because of its stability, the amount of carbon-13 in a substance can be used to date the substance over long periods of time.

2.2.2 Mass spectroscopy

Mass spectroscopy is used to determine the components of a substance based on the mass of the elements present within it. There are a number of steps relevant to the mass spectroscopy process.

Firstly, the molecule we are interested in is ionised to form a positive ion. Since we are really only looking at the mass of the element in this method, and the mass of an electron is so, so much smaller than the mass of the nucleus, this ends up making no difference to the final result.

Secondly, not that we have a positively charged particle, we can accelerated it using a magnetic field. Magnetic fields can make ions move very quickly in large or small circles, depending on the strength of the field and the forward velocity of the ion. By getting each ion to the same kinetic energy, we know that the 'radius of curvature' will be dependant on mass only.

Next, we use this notion of 'radius of curvature' to deflect the ions based on their mass. Lighter particles will be deflected more easily, whilst heavier particles take a stronger field to be deflected. This separates out the ions based on their mass!

Finally, we detect 'how many' ions pass through particular regions of the detector, which will be based on how much the ions have been deflected (i.e. how heavy the ions are).

The mass spectroscopy machine will output a mass/charge ratio, and 'count' how many ions have that particular mass/charge ratio. We can therefore determine how many different isotopes of a particular element were present after the ionisation stage above!

2.2.3 Infrared spectroscopy

Infrared spectroscopy is used very widely in chemistry to determine the structures and compounds present within a particular substance.

When an atom absorbs energy, it will 'increase' in quantised energy levels. However, it can do this in three distinct ways. Firstly, as we know from above, electrons can absorb the energy and either be 'knocked out' of their orbit, or be raised to a higher energy level before releasing a photon of light and coming back down to their original level. Secondly, the particles can 'spin' more quickly, adding rotational energy to the system. Thirdly, the particles can 'vibrate' more quickly. We are interested in the third energy change here.

When infrared energy is incident on a molecule, it can cause the molecule to transition between vibrational energy states. This changes what is known as the 'dipole moment' of the molecule, which can be thought of as a 'spin' of a particle in relation to its two 'poles' (i.e. positive and negative end). By measuring the absorbance of IR light, and the re-emission once the vibrational energy goes back to its ground state, we can get a 'characteristic' curve of the molecule in question.

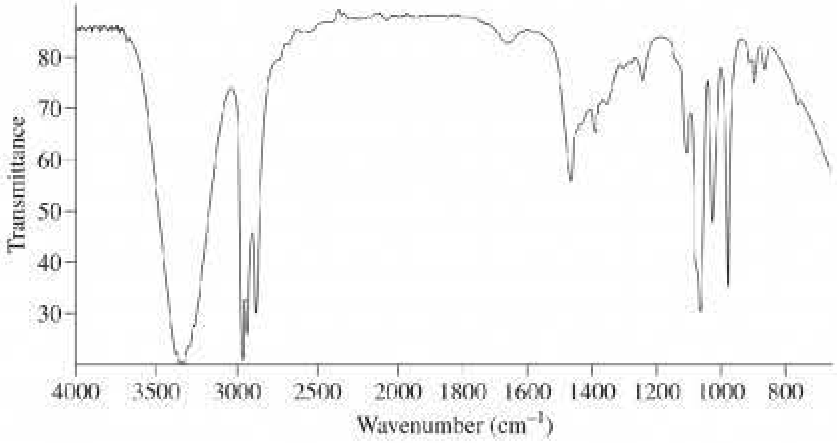
2.2.4 Answering analysis tool questions

Now that you have a basic understanding of the mechanism outlined in the course, it's important to be able to put this understanding into practice, because you will almost definitely get NMR/infrared/UV absorption questions on the exam!

If you take a look at the second page of the formula sheet, you will find a number of tables with common bond-types or chemical structures, and associated numerical values. You need to be able to read these tables, and interpret the data that an exam question provides.

Infrared

The sample questions posted on the NESA website asks you to identify the compound analysed which resulted in the following infrared spectrum:



So how can we determine a compound using this information? The options we are given are:

- a) Butane
- b) Propanol
- c) Propanal
- d) Butanoic acid

Now, let's turn to the formula sheet. Specifically, we only really care about the **Infrared absorption data** table, which looks like this:

Bond	Wavenumber/cm^{-1}
N—H (amines)	3300–3500
O—H (alcohols)	3230–3550 (broad)
C—H	2850–3300
O—H (acids)	2500–3000 (very broad)

What does the table tell us? Remember that infrared spectroscopy involves shining infrared light at molecules, which will either be reflected or absorbed. When the energy is absorbed, the receiver will see a 'dip' at that particular characteristic wavelength of infrared light. So, particular bonds will have particular characteristic wavelengths, which can be used to determine the substance. We know from the weird-looking plot thing that there is a big dip around 3,300–3,500 cm^{-1}. Looking at our table, we now know that the dip could be due to an amine subgroup or an alcohol subgroup. It's probably an **alcohol** subgroup, because we are told that these dips are **broad,** but we can't say that for sure.

Let's return to our multiple choice options. How many of them have either amines or alcohols? Well, thankfully, none of them have amines in their structure, so we can rule that out. However, there are two substances of O–H subgroups: propanol and butanoic acid. We also have a dip around the 2800–3000 cm^{-1} point, which is probably due to the many C–H bonds. However, this doesn't help us distinguish between compounds either, as both of the remaining options contain plenty of C–H bonds.

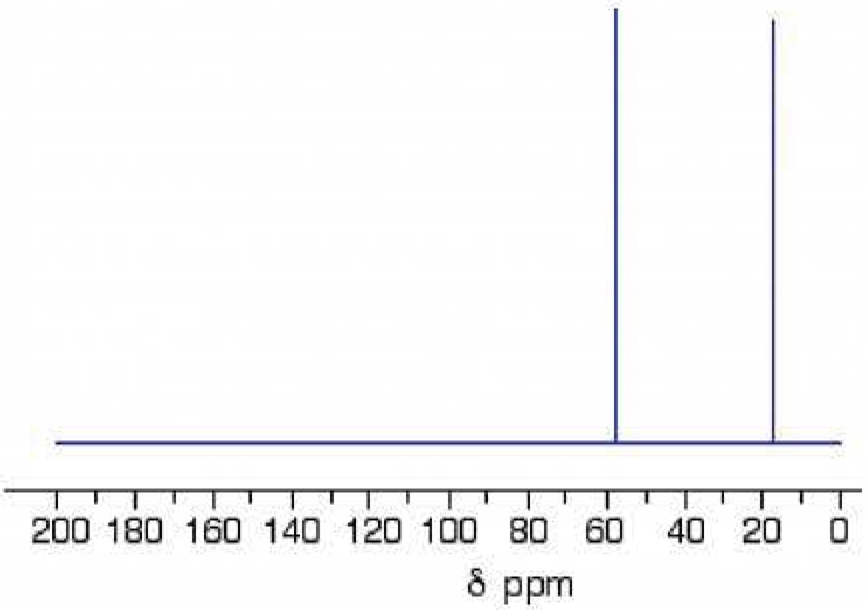
To distinguish these two compounds, we ask ourselves: *what is the main difference, in terms of bonds, between propanol and butanoic acid?* Well, it's probably that butanoic acid has a double-bonded oxygen atom attached to the backbone structure! From the table:

C=O	1680–1750

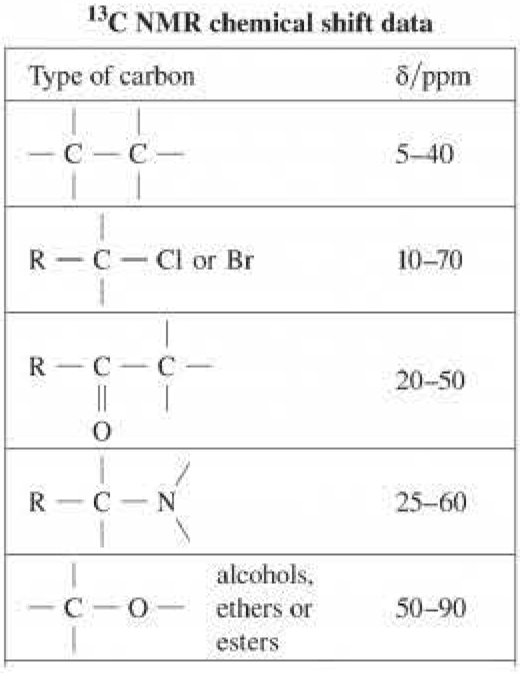
So, if the substance is butanoic acid, we expect a dip around 1680–1750 cm^{-1}. But, there is no such dip! So we can rule out any alkanoic acid, and determine that the substance is in fact propanol (answer b).

NMR

Let's take a look at characteristic NMR plots which show peak spikes as a function of ppm. The plot below on the left shows peaks at around 59 ppm and 19 ppm. Usually the plot will be more 'noisy' in that the data will be more spread out and a bit less clear, but let's use this one as an example. On the right is the table provided by NESA.



^{13}C NMR chemical shift data



Type of carbon	δ/ppm
$-\overset{\vert}{\underset{\vert}{C}}-\overset{\vert}{\underset{\vert}{C}}-$	5–40
$R-\overset{\vert}{\underset{\vert}{C}}-Cl$ or Br	10–70
$R-\underset{\underset{O}{\Vert}}{C}-\overset{\vert}{\underset{\vert}{C}}-$	20–50
$R-\overset{\vert}{\underset{\vert}{C}}-N\lt$	25–60
$-\overset{\vert}{\underset{\vert}{C}}-O-$ alcohols, ethers or esters	50–90

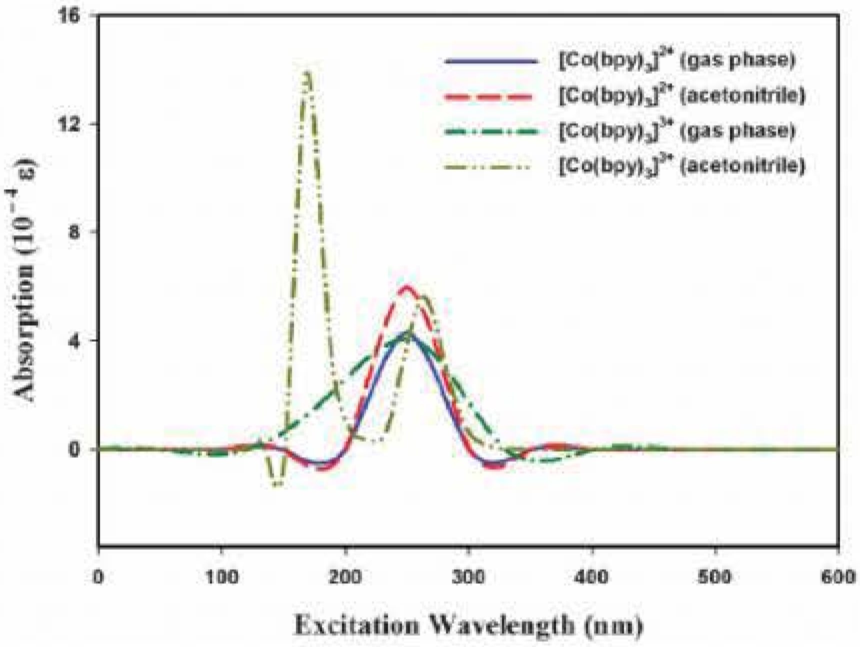
So, we can be pretty confident that the 19(ish) ppm peak is either a C–C bond, or a C–Cl/C–Br bond. It's also likely that the 59(ish) ppm peak is either a C–N bond or a C–O bond.

It's difficult to actually work out the answer from here. Because there's only one peak of each, it suggests that the relevant bonds exist only *once* in the chemical compound. It just so happens that the above plot is the NMR spectrum of ethanol: one carbon–carbon bond, and one carbon–oxygen bond (where the carbon chain is attached to the hydroxyl functional group). You would only really be able to work that out by a process of elimination in multiple-choice questions!

UV

UV plots look pretty complicated, and it can be hard to get good data out of them without spending quite a lot of time on it. However, it is an extremely commonly used tool (particularly in Chemistry and Biology to determine the concentration of a liquid) because it is cheap, quick, and reasonably accurate.

The plot is usually one of absorbance against wavelength. Maximum absorbance will occur at the peak of the plot. Below is an example, though you are not expected to be able to read anything from it!



Topic 3

Chemical Synthesis and Design

SYLLABUS :
Inquiry question: What are the implications for society of chemical synthesis and design?

3.1 Evaluating chemical synthesis

SYLLABUS :
Evaluate the factors that need to be considered when designing a chemical synthesis process, including but not limited to:

- Availability of reagents
- Reaction conditions
- Yield and purity
- Industrial uses (e.g. pharmaceutical, cosmetics, cleaning products, fuels)
- Environmental, social, and economic issues

This dot point appears to be entirely self-explanatory; all of the listed factors must be considered when designing and implementing a chemical synthesis process, as they are all relevant either to the chemical process or the impact of the chemical process. Rather than discussing each of the listed items in depth and in isolation, I think it's easier to look at a particular process and link the above factors to that process. I've decided to use the Haber process for this.

3.1.1 Availability of reagents

Prior to WWI, ammonia for use in fertilisers was sourced from Chile. However, during WWI the British blockaded Germany by stationing their navy around the coastal ports of the country. This stopped any import of ammonia, which was vital for crop growth. As a result, crop yield decreased dramatically, and hundreds of thousands of German civilians starved to death. Additionally, ammonia was required to develop many munition resources vital for the German war effort.

Haber, a German scientist, developed a process to synthesis Ammonia for fertilisers and explosives. By extracting the reagent from the air, and carefully monitoring the conditions of synthesis, Ammonia could be synthetically produced.

The Haber process abides by the following chemical formula:

$$N_{2(g)} + 3H_{2(g)} \leftrightharpoons 2NH_{3(g)} \qquad \Delta H = -92 \text{ kJ mol}^{-1}$$

3.1.2 Reaction conditions

When it comes to the reaction conditions of a particular chemical synthesis process, there are typically four things we consider. Three of them have to do with Le Chatelier's principle, as discussed earlier on the course. The fourth has to do with the activation energy of a particular synthesis.

Concentration

First, we investigate what happens when we change the concentration of either the reactants or the products. If we increase the concentration of the reactants, then by Le Chatelier's principle the equilibrium will shift to minimise that change. In other words, the equilibrium will want to decrease the amount of reactants present, and thus shift away from the reactants. We call this 'shifting to the right', as more products are produced.

In the Haber process an increase in the concentration of nitrogen and hydrogen will cause a shift towards the ammonia gas. If we increase the concentration of the products, then by Le Chatelier's principle the equilibrium will shift to minimise that change. In other words, the equilibrium will want to decrease the amount of products present, and this shift away from the products. We call this 'shifting to the left', as more reactants are produced.

Thus, the reaction conditions we want to maintain the Haber process is a 1:3 ratio of reactants, and we want to continually add reactants and remove products.

Pressure

Next, we investigate what happens when we increase or decrease the pressure of a system. If we increase the pressure of a system, then by Le Chatelier's principle the equilibrium will shift so as to minimise that change. In other words, the equilibrium will want to lower the pressure. Remember that each mole of gas occupies the same volume. So, to lower the pressure, the equilibrium will shift towards the side with fewer moles of gas! In the Haber process, an increase in pressure will cause the equilibrium to shift to the right (as the products contain 2 mole of gas, but the reactants contain 4 moles of gas).

If we decrease the pressure of a system, then by Le Chatelier's principle the equilibrium will shift so as to minimise that change. In other words, the equilibrium will want to raise the pressure. Remember that each mole of gas occupies the same volume. So, to raise the pressure, the equilibrium will shift towards the side with more moles of gas! In the Haber process, we want to keep up quite a high pressure. This is to ensure that the equilibrium shifts to the right. However, we must keep economic and safety factors in mind as well. It can be very expensive, and very dangerous to maintain very high pressures (particularly when dealing with flammable hydrogen gas!). So, the pressure typically used is about 200 atm.

Temperature

Next, we investigate what happens when we increase or decrease the temperature of a system, If we increase the temperature of a system, then by Le Chatelier's principle the equilibrium will shift so as to minimise that change. In other words, the equilibrium will want to lower the temperature. Remember that exothermic reactions release energy when moving to the right, and endothermic reactions absorb energy when moving to the right. So, to lower the temperature, exothermic reactions will move to the left (towards the reactants), and endothermic reactions will move to the right (towards the products). In the Haber, we know that the system is exothermic (as it releases heat in order to move forward). As such, to lower the temperature the equilibrium will shift to the left (towards the reactants).

If we decrease the temperature of a system, then by Le Chatelier's principle the equilibrium will shift so as to minimise that change. In other words, the equilibrium will want to raise the temperature. Remember that exothermic reactions release energy when moving to the right, and endothermic reactions absorb energy when moving to the right. So, to raise the temperature, exothermic reactions will move to the right (towards the products), and endothermic reactions will move to the left (towards the reactants). In the Haber process, to maximise yield we want the equilibrium to shift to the left. As the forward reaction is exothermic, we would technically want to lower the temperature of reaction. However, we need to also consider kinetic factors such as collision theory. As temperatures increase, the speed of gas molecules increase. This makes it much more likely for a reaction to occur on a given collision (as well as makes it more likely for collisions to occur). As such, the process is usually performed at about 450°C.

Catalysts

Catalysts do not affect the reaction itself; instead, they lower the activation energy. In the case of the Haber process, we use a magnetite (Fe_3O_4) catalyst fused with K_2O, Al_2O_3, and CaO in a fine powder. This promotes the reaction and allows for decreased pressure (and is thus a more cost-effective process).

3.1.3 Yield and purity

All of the previously mentioned conditions have been determined based on the best way to increase the yield of ammonia (i.e. to force the reaction towards the products). This yield, according to online sources, is typically about 10–20%. Thus, it is important to purify the resulting gaseous mixture, or the majority of the extracted gas would actually be the left over reactants.

3.1.4 Industrial uses

Ammonia is used for many, many purposes in industry. Like in WWI, ammonia is still used in fertilisers, and to develop explosives. In addition, ammonia is a very common chemical used for many other synthesis reactions. Ammonia is used in refrigeration systems, pharmaceuticals, paper production, and cleaning liquids.

3.1.5 Environmental, social, and economic issues

Generally speaking, the Haber process has had a very positive impact on environmental, social and economics components of today's society. By allowing for very productive fertilisers, the space required for crops had diminished dramatically. The ease with which crops may grow decreases the cost of food to the consumers, and allows a more efficient farming economy. By minimising the space needed for crop growth, the environmental impact is minimised.

However, by facilitating this ease of food production, the planet's human population has been capable of growing at an enormous rate since the first World War. Energy consumption has grown exponentially, as has the demand for consumable goods, land, and luxuries. Such a population growth, whilst obviously not exclusively attributable to the chemical process discussed above, is necessarily damaging to the environment. Many people have been left behind due to their socioeconomic status, and the distribution of wealth has never been so uneven.

Synthesis of chemicals can have a huge impact on the environment, on society, and on the economy. It's often hard to predict the impact of a new development in Chemistry; many scientists argue that they shouldn't even have to consider this impact when discovering new principles. This is a decision you'll have to make for yourself; should scientists always be conscious of the society around them, and the impact of their work on that society? Or is their only job to discover new information, to push science further and further, without regard to the potential consequences?

Part V

Exam and Revision Tips

Section 1

How to Prepare for the Chemistry Exam

1.1 Chemistry study tips

Few people have a 'natural' Chemistry ability. Most people, in fact, start on an absolutely level playing field when it comes to Year 12 Chemistry. Therefore, Chemistry is anyone's subject. What will decide how well you do, and whether or not you ace the subject, is how successfully you employ various study techniques throughout your year. These are my top tips; give them a go, and I promise you that you'll see better results.

Summarise Chemistry at the end of every week

It's super easy to fall behind in the Chemistry curriculum. There are just so many (often random and totally unrelated) dot points that you will cover in class; even missing one or two can be fatal. There is no way to tell when one of the tiny dot points, or a single prac, will be worth 8 marks in your HSC exam.

Thankfully, you've already taken the first step to success; you've got these complete course notes! However, it's also important to stay on top of your class work by taking your own notes.

It's hard to stay motivated throughout the year. But, if you regularly set aside a time to summarise Chemistry once a week, I promise that you won't regret it. Making sure that keep up to date with your learning, keeping a collated document with everything you need to know for your HSC, will make your journey through this subject so, so much easier. Even just writing down difficult chemical terms makes it so much quicker to memorise them.

Use colours

Make sure to allocate certain colours to have certain meanings (important chemical equations, or extra information that you don't necessarily need to memorise). Just taking that little bit of extra time assessing how you are going to structure your notes will make them much, much more effective. Set your notes out in the smartest possible way. If you can use a table, use a table. If you can use dot points, use dot points. If you can include a diagram, include a diagram. Use these notes as inspiration if you need to!

This is particularly important, as the same goes for answering actual assessment questions. If you can answer in an advantages/disadvantages table, you should! By setting your notes out such that you can literally transplant them into answering difficult HSC-style questions, you are far more likely to retain the information. Think of it as a neurological trick: if you've looked over an answer a few times, it would be far easier to remember the information if you write it down in the order, and the style, that you read it!

Keep a separate list of vital equations

This is something that I only did at the end of my HSC. I would write a list of the information I absolutely *had* to memorise, and just hadn't yet. Then, I would do my best to learn the information, see what I could cross off my list, and then do it again. By the time it got to my actual Chemistry HSC examination, this is the only document I was studying from.

The document only needs to make sense to you. The idea is that you compile absolutely all the information you need to memorise, so that you can study effectively from that document. You should really be doing this for every single subject.

Use worksheets

Now that you have the list of difficult information, how do you actually memorise it? In my opinion, there are two ways. Firstly, just answer an unbelievable amount of actual questions. Complete questions applying the information you are trying to learn, first with the aid of the one-page cheat sheet, and later without it.

Secondly, write yourself worksheets applying the information. Create worksheets with blank spaces where you need to fill in the important information (whether that be statistics, chemical equations, or anything else!). Photocopy the worksheets a hundred times, fill it in five times a day. If you start doing that soon, you'll know absolutely *everything* by your exams, and be able to recite it in your sleep by the HSC. This was one of the most useful tools I implemented in my HSC; give it a go, and thank me later!

Work through past papers

Past papers. Do them. Do them again. Do them again and again and again.

Scroll through sample or past exams and *any* question related to the topics that you will be assessed on. By doing practice questions, you get a headstart on the actual exam that you will sit, and it's also the best way of figuring out which areas you need to build your knowledge of.

Use online resources

If you don't understand a topic, there is an incredible amount of information online. Try YouTube, or just Google whatever it is that you're struggling with.

The best place to look for HSC-specific answers is atarnotes.com where where you can ask any question you like (about Chemistry, or any other subject). We'll answer you asap, for free!

1.2 Common exam problems

Mathematical mistakes

Many student struggle to answer Chemistry questions that involve Mathematics. Whilst I try to explain how to answer particular kinds of questions in these notes, there are some more general considerations worth keeping in mind.

Units

It's seriously important that you have a good understanding of units in the HSC Chemistry curriculum. Obviously, it is imperative that you include units on any answer you give. If you are asked to find the heat of combustion, and you answer '1360', what does that actually mean? 1360 mol? 1360 camels? No, 1360 kJ or 1360 kJ/mol, depending on the question. Always keep units in mind when answering questions.

However, you can also use units as a powerful tool when attacking maths problems. For instance, students often struggle to convert a measure of kJ/mol to kJ/g. However, if you understand units, it's easy! Recall that

$$1 \text{ mol} = \text{M g}$$

where M is the molar mass of the substance. Therefore, we can rewrite our original units:

$$\frac{\text{kj}}{\text{mol}} = \frac{1 \text{ kJ}}{1 \text{ mol}} = \frac{1 \text{ kJ}}{\text{M g}}$$

So, $X\frac{\text{kJ}}{\text{mol}} = \frac{X}{M}\frac{\text{kJ}}{\text{g}}$! Easy – when you have a good understanding of units.

Significant figures

It is important to use the correct number of significant figures. Recall that an answer should have the same number of significant figures as the lowest s.f. value in the question. So, if a question gives you

m = 120 g   C=3.4 mol/L g   V=24.79 mL

You need to quickly be able to recognise that the mass value has three significant figures, the concentration value has two significant figures, and the volume value has four significant figures. Thus, you should be using two significant in your answer! Issues like this won't always make you lose a mark, but they may sometimes cause you to miss out on full marks.

Chemical details

Sometimes, you'll find that it is possible to answer a question without referring specifically to chemical principles. For instance, in answering an 8 mark Production of Ethanol question, you could conceivably describe the process without actually using chemical equations. **This would be a mistake.**

Wherever possible, it is important that you include chemical equations and diagrams. It's required to get high marks, and will also ensure that you get the result you deserve. If you could include chemical detail, you should include detail.

Structuring your answers

Remember that, when answering an extended response question, there are more structures than just paragraph after paragraph after paragraph. In fact, there are many preferably techniques you can implement to simplify the marking process for the marker. The easier your answer is to read and process, the more likely a marker is to give you a good result!

Firstly, try to use tables when you're comparing ideas/properties. Advantages/disadvantages tables are very common, and should be used in almost every exam. When reading these notes, think about where I've used tables and think about whether you might consider doing the same.

Secondly, if you do have to write paragraphs, make sure to use subheadings. It tells the marker that you've hit all the important points, and clarifies your argument or discussion. Before answering an extended response question, take some time to think about the important information you will include and compartmentalise that information into particular subheadings.

Finally, don't be afraid to use flow-charts (as long as you box in all of your objects). Flow-charts are the best way to display any process. I would write out my flow chart, with specific information next to each box. Give it a go! See if it works for you.

Length of your answers

Use the number of line given to answer a question as a guide to the length of your answer. If you are writing on the back of a page, or into the margins, it is very likely that you are writing too much. Sometimes (very rarely) this is necessary. However, it is presumed that the space given is sufficient.

So, particularly for extended responses, spend some time thinking about how you may best use your space and your time. Don't just write; consider the question, and the best way to answer it.

1.3 Scaling of HSC Chemistry

Often, students spend a lot of the year fretting over scaling. How are my marks aligned? What is the Band 6 cut-off? Will Chemistry pull up/pull down my overall ATAR? How does scaling even work? So, I want to quickly try to cover how Scaling effects HSC Chemistry, to hopefully clarify any concerns you may have. For a complete picture of scaling, check out http://atarnotes.com/how-does-scaling-work/.

First things first; at the end of the day, there isn't any way to 'play' the system when it comes to scaling. If you choose subjects you enjoy or are good at, you'll get the best marks possible. If you choose a subject purely for the 'scaling benefits,' you'll likely do worse in it overall, because you won't feel as committed to the topic area. So, the only way to get the highest possible ATAR is to work as hard you can.

The alignment of your exam marks is fairly tricky to have real data for. The HSC Raw Marks Database is an online resource where past students have uploaded their raw exam marks, and what was reported on the HSC. Students only know this information if they've paid to see their raw exam marks. There is no way to know if all of this information is 100% accurate, but it is an indicator nonetheless. According to that database, here are some examples of what alignment did to some Chemistry students' raw marks in 2015.

Before alignment	After alignment
91/100	96/100
87/100	93/100
68/100	82/100

So, if your exam is particularly tough, your marks will be increased! This means you don't need to worry if you leave the exam thinking you totally messed it up; likely, it was just a really difficult paper, and the whole cohort will receive a nice little scaling boost.

The other thing that happens, beyond moderation of internal marks and things like that, is **scaling.** This is something we do have reliable data for. I highly recommend a read of the **Report on the Scaling of the NSW Higher School Certificate,** which is published every year by UAC. According to that report, here are some examples of what your HSC Marks would have scaled to in 2015. I've put them on top of English and Mathematics for some comparison. Notice that pretty much everything scales down; that's nothing to worry about and it is compensated for when your ATAR is calculated from your aggregate!

Chemistry	**HSC Mark /50**	49.5	47.5	45.0	42.5	38.5	34.5
	Scaled Mark /50	50	47.1	43.1	39.2	32.9	25.2

English Advanced	**HSC Mark /50**	49.5	47.5	45.5	43.5	40.5	38.0
	Scaled Mark /50	50	46.7	42.3	38.2	32.8	27.0

Mathematics	**HSC Mark /50**	50	39.0	46.5	33.0	40.0	36.5
	Scaled Mark /50	50	47.6	42.9	38.6	32.8	25.5

Remember, whilst an understanding of scaling is pretty important, it really won't make much of a difference to your mark. Just try as hard as you can, and your ATAR will reflect that work.

Best of luck with your Chemistry studies!