

The Student's Guide to
HSC Chemistry

Core Content
& Industrial Chemistry

2009 Edition

The next best thing to cheating

Alan Wong

Illustrations by Michael Zhou

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About the Guide

The Student's Guide to HSC Chemistry is a brand new form of study guide, acknowledging the deficiencies of the way existing resources are presented to students while simultaneously accentuating their strengths. All of this is done in a way which closely mirrors the way many students already organise their own notes a method adopted for several reasons. Whilst textbooks and various other resources may have been compiled according to their own structure for time immemorial, it has been noticed that time and time again students turn to a syllabus dotpoint format when constructing their own notes. It is my belief that this is an extremely effective way, organising the students thoughts in a comprehensive approach which includes all necessary details while omitting yet other details which are really quite superfluous to scoring full marks in an exam.

This guide offers a means of revision, and in doing so must be distinguished from a textbook. A syllabus is provided by the Board of Studies for every course, detailing precisely which areas are examinable and what is expected of students. As stated above, the unique structure of this guide allows for a definitive treatment of each dotpoint, detailing exactly what must be learnt to achieve the highest possible marks, while offering the benefit of streamlining all information such that the student is far less likely to feel overwhelmed with information. At the very least, it provides a useful overview for introductory and/or review purposes so as to make Chemistry that much simpler.

By the very same token, it must be stressed that this guide is in itself simply a means of revision. While I have no doubt it is sufficient to gain a student a respectable mark by itself, if you truly want to achieve your highest potential mark, I strongly urge you to turn to the myriad of resources around you. Where possible, dotpoints are expanded upon even at the sake of conciseness simply so that the point of a dotpoint appears that much more logical. However, how a student learns a concept is highly dependant upon how they personally view the concept, and as such, it may be that one student needs to read up about Concept A in Textbook X yet another student may need to read up about Concept B in Textbook Y. These resources should not be overlooked in the blind belief that simply accepting a fact to be true is sufficient. Textbooks, teachers, internet sites, and most importantly your fellow peers will form the crux of these resources.

At the end of the day, if you get one thing from this guide, let it be the fact that most (I am reluctant to say every, but most) results can be reached through a relatively logical process. If you can reason a result out step-by-step, then you will have no difficulty in expanding upon it come exam time as you not only know it, you *understand* the concept. In utilising this guide alongside other resources, I have no doubt that this learning process will be greatly simplified.

All that then remains to be said is best of luck!

Alan Wong

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Chapter 1

Production of Materials

1.1 Energy and Raw Materials from Fossil Fuels

1.1.1 Identify the industrial source of ethylene from the cracking of some of the fractions from the refining of petroleum

Several forms of cracking are possible. However, simply learning catalytic cracking is adequate. Be prepared to write at least one equation to demonstrate this.

Ethylene is certainly one of the most useful products derived from the refining of petroleum. As such, a process known as catalytic cracking is often used to break down the higher molecular weight hydrocarbons into more useful, lower molecular weight hydrocarbons such as ethylene.

In the process of cracking, special catalysts called zeolites made of inorganic compounds are used. These zeolites are porous, such that many cavities exist within the structures, thereby increasing their surface area and thus effectiveness. Zeolites are typically made of compounds of aluminium, oxygen and silicon.

Catalytic cracking requires atmospheric pressure, an absence of air, and temperatures of approximately 500°C. Long hydrocarbon chains are repeatedly broken down into smaller chains, typically one alkane and one alkene, until the desired product such as ethylene is created.

Breaking down a hydrocarbon chain into the smaller products of decane and ethylene:



Remember- The purpose of catalytic cracking is to produce ethylene, which holds unlimited potential in the petrochemical (plastics) industry.

1.1.2 Identify that ethylene, because of the high reactivity of its double bond, is readily transformed into many useful products

The crux of this dotpoint is the presence of the double bond within the ethylene molecule. Make sure you understand what the terms 'saturated' and 'unsaturated' mean, and try to understand why the double bonds are so reactive, not simply that they are. An explanation of electronegativity is provided below to help with this.

A hydrocarbon can be either saturated or unsaturated. A saturated hydrocarbon contains only single bonds, and no more atoms can be added to it. In contrast, an unsaturated hydrocarbon can have double, or even triple bonds, giving the possibility of further atoms or molecules joining the existing hydrocarbon chain.

Unlike alkanes, alkenes are unsaturated, as they contain a double bond which readily allows them to undergo addition reactions. The double bond of alkenes such as ethylene is inherently more unstable than a single bond, and thus breaks relatively easily to bond with other atoms and/or molecules.

In addition, the double bond present in ethylene is also a site of high electron density. Thus electronegative species such as halogens readily react with ethylene

Electronegativity is a property which describes the strength of an atom's pull on electrons. Halogens are highly electronegative, and thus an area of high electron density would naturally attract a halogen.

Remember- The high electron density of the double bond present within ethylene that the molecule becomes such a useful building block in the petrochemical industry.

1.1.3 Identify data, plan and perform a first-hand investigation to compare the reactivities of appropriate alkenes with the corresponding alkanes in bromine water

The bromine water experiment is relatively simple, demonstrating the differences between the reactivities of alkanes and alkenes through observations. The trick is to identify which solutions decolourise the brown-coloured bromine water.

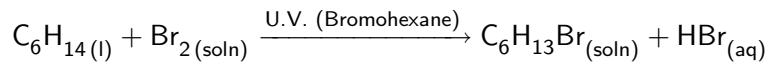
Procedure:

1. Pour 10mL of bromine water each into two small beakers.
2. Using a dropper bottle, place a few drops of hexane into one of the beakers, noting its effect. Stir gently with a glass stirring rod.
3. Repeat with a dropper bottle of ethylene.
4. Repeat the procedure with a variety of alkanes and alkenes. With solutions that alkanes have been added to, place them near a bright window and observe the effect over time.

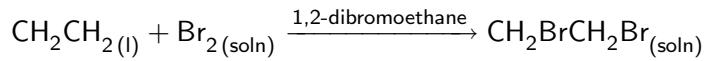
Expected results:

Alkanes do not react with bromine water, meaning that the brown solution does not decolourise. However, the presence of ultraviolet light will result in a reaction, thus decolourising will occur. Although this reaction may be hard to see, as the reaction can be somewhat slow, it does still occur.

Below are examples of reactions between an alkane and bromine water.



Due to the reactivity of the double bonds in alkenes, alkenes will decolourise the brown-coloured bromine water.



Remember- The presence of UV light will decolourise the bromine water solution if an alkane is added. If an alkene is used, UV light is not required.

1.1.4 Identify that ethylene serves as a monomer from which polymers are made

Ethylene may undergo addition or substitution reactions with other monomer units to form polymers such as polyethylene, polyvinyl chloride (PVC), polystyrene, polypropylene, as well as many others. As such, ethylene is a monomer from which polymers are made.

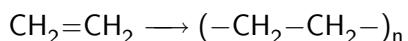
One example is the formation of polyethylene. This is achieved through the breaking of the double bond so that each monomer unit can now attach to one another to form a chain of monomer units - a polymer.

Remember- Due to the high reactivity of the double bonds, ethylene serves as an extremely versatile monomer in the polymerisation of products such as polyethylene, PVC, polystyrene, and many other polymers.

1.1.5 Identify polyethylene as an addition polymer and explain the meaning of this term

Although these dotpoints introduce polymerisation with the process of addition polymerisation, it is important to remember that other processes such as condensation polymerisation also exist. A brief explanation of both is provided below.

An addition polymer is a polymer which is formed by the joining of individual monomers without the loss of any atoms. This differs from a condensation polymer, where a small molecule is usually removed from the chain for every monomer unit present. Polyethylene is one such example of an addition polymer as the double bonds present in the monomer unit ethylene allow for the addition of many ethylene molecules to form polyethylene.



This may be hard to visualise, so take full advantage of any molecular modelling kits your school may have in order to understand how the polymers are formed.

Remember- Polyethylene is an addition polymer, formed as monomer units join together. This is not to be confused with condensation polymerisation, as there is no loss of any atoms or molecules.

1.1.6 Outline the steps in the production of polyethylene as an example of a commercially and industrially important polymer

There are two main processes in the production of polyethylene, each producing a distinct product which reflects the process used. Learn each process well, noting that the density of each product is resultant upon the structure of the polymer determining by the process.

In addition, be prepared to relate the use of each product to their characteristics, such as low or high density. Examples of LDPE products include trays and containers, as well as flexible and pliable components. Examples of HDPE products include plastic bags and pipes as well as plastic bottles.

Low-density polyethylene (LDPE)

For the production of LDPE, pressures of approximately 2000 atmospheres and temperatures of 300°C are used along with an organic peroxide called an 'initiator'. This process, known as the older gas phase as well as free radical polymerisation relies upon the initiator to open the double bonds in the monomer units, which then combine. This process results in alkyl groups periodically branching out where hydrogen atoms are usually present, producing a polymer that is low in density.

High-density polyethylene (HDPE)

The Ziegler-Natta process is used for the production of HDPE, where pressures about 20atm and temperatures of approximately 60°C are used along with a catalyst usually made from titanium chlorides and trialkyl aluminium compounds. This allows for surface polymerisation to occur, producing closely-packed polytethylene molecules without the presence of alkyl branches. As such, HDPE is considerably denser than LDPE.

1.1.7 Analyse information from secondary sources such as computer simulations, molecular model kits or multimedia resources to model the polymerisation process

This dotpoint is useful for students unable to see the reactions taking place simply through the chemical equations. More importantly, it is useful to note the benefits and limitations of models, as this question does appear from time to time in various papers. For this reason, a list of the benefits and limitations of modelling is provided below.

Benefits:

- Provides a physical representation of the type and quantity of atoms involved in a molecule
- Demonstrates the difference between the various bonds in a molecule
- Provides a simple representation to aid understanding

Limitations:

- Relative sizes of, and distances between the atoms are unrealistic
- The dynamic nature of various molecules and their bonds is not shown
- Oversimplifies the model

Remember- On top of noting these benefits and limitations, also refer to dotpoint 1.1.5 on page 4 for further information regarding the addition and condensation polymerisation processes.

1.1.8 Identify the following as commercially significant monomers: vinyl chloride and styrene- by both their systematic and common names

Common Name	Systematic Name
Vinyl Chloride	Chloroethene
Styrene	Phenylethene/ethenylbenzene

The polymerisation of the monomers vinyl chloride and styrene yield polyvinylchloride and polystyrene respectively. Each polymer is of great commercial significance due to the properties they display. This point is touched upon further by the next dotpoint 1.1.9 on page 6.

Remember- Vinyl Chloride is also known by its systematic name of chloroethene and polyvinyl chloride is referred to as polychoroethene. Styrene can be known by both phenylethylene and ethenylbenzene. By the same token, polystyrene may be referred to as polyphenylethene as well as polyethnylbenzene.

1.1.9 Describe the uses of the polymers made from the above monomers in terms of their properties

Always relate the uses of the polymer with its properties. State one property, such as the water resistant nature of PVC, and then relate this to its use as raincoats and shower curtains. Simply listing properties followed by a list of users will not get you the best possible marks

Polyvinylchloride is both water and flame resistant, as well as relatively durable as it does not readily react with many chemicals. It is also rigid, strong, and does not conduct heat or electricity. As a result of these properties, PVC is commonly used for insulation and drain pipes, as well as raincoats and shower curtains. This range of PVC products, from thin films to rigid items, demonstrates the versatility of PVC.

Polystyrene is an effective heat, cold, and electrical insulator. When these characteristics are taken into consideration along with the ability for gas to be blown into the polymer (Producing a rigid and low-density material), uses such as insulators and packaging are self-evident. Polystyrene is also not chemically reactive, allowing for safe use in plates and foam cups. Because polystyrene has few crystals, it can also be made transparent, thus enabling uses such as CD cases and various containers.

1.2 Materials from Biomass

1.2.1 Discuss the need for alternative sources of the compounds presently obtained from the petrochemical industry

This dotpoint can be argued many ways, all of which are valid if supported by logical reasoning. However most arguments can be categorised under the headings of 'scarcity' and 'environmental impact', so it may do well simply to remember these headings and to argue them whatever way you choose.

The need for alternative sources of petrochemical product derivatives comes down to two main points: scarcity and environmental impact.

- Petrochemical products are derived from non-renewable sources of crude oil. With some experts placing the lifespan of current petroleum sources well under 50 years, and natural gas sources within 100 years, alternative sources are required simply because current production trends are unsustainable.
- As roughly 95% of crude oil is used up as fuel, the consumption of fuel products has an enormous impact upon the environment. In comparison to other potential fuels such as ethanol, the current petrol products consumed by most of the world burns relatively uncleanly, leading to environmental problems such as the greenhouse effect and acid rain. The biodegradability of many products also places considerable strain on our landfills. Alternative sources of compounds obtained from the petrochemical industry may at the least alleviate such problems.

Remember- Alternative sources of the petrochemical products must be discovered for both practical reasons of scarcity and for fear of doing irreparable harm to our environment.

1.2.2 Use available evidence to gather and present data from secondary sources and analyse progress in the recent development and use of a named biopolymer. This analysis should name the specific enzyme(s) used or organism used to synthesise the material and an evaluation of the use or potential use of the polymer produced related to its properties

There are many possible biopolymers that you may choose to satisfy this dotpoint. For the purposes of this guide, a biopolymer commonly used by students will be examined.

The headings below have been used as a general template. When answering any questions, be sure that you have answered exactly what the question is asking for, as they may sometimes attempt to focus on a specific point. Be prepared to also focus on something as simple as the advantages and limitations of using the chosen biopolymer.

First developed in 1925 by Maurice Limoigne, Polyhydroxybutanoate, or PHB, is a naturally occurring biopolymer which has been produced in laboratories for its many special uses.

Production

PHB is produced by feeding bacteria on a nutrient-rich diet until large colonies of the bacteria begin to form. At this point, glucose is withdrawn from the diet, and the bacteria will automatically secrete PHB as an energy store, similar to the body fat of humans.

One bacterium which produces PHB is *Alcaligenes eutrophus*. In recent times, genetic engineering techniques have enabled scientists to locate the specific gene responsible for the secretion of PHB, and then to transfer this to bacterium such as *Escherichia coli*, more commonly known as *E. Coli*. Advantages to using *E. Coli* generally centre around the fact that scientists are more familiar with the bacterium, and as such, find it much easier to work with, providing a faster production rate.

Properties and Uses

- Naturally occurring- As a biopolymer, PHB is both non-toxic and renewable, offering an eco-friendly alternative to most plastics.
- Biodegradable- Decomposing into carbon dioxide and water, PHB has a large environmental advantage over other polymers which take up large amounts of space in landfills.
- Physically similar to polypropylene- Although PHB has a chemical structure markedly different from the polymer polypropylene, its physical properties are quite similar. As such, it can be readily used as a substitute in many of its applications.
- Biocompatible- Compatible with biological systems, PHB has a highly practical application in the medical industry with items such as medical sutures.

Development & Impact

Although the development of genetic engineering techniques has been remarkable, the costs of production of biopolymers such as PHB are still too high to make the process economically viable. This was most noticeable by the initiation and subsequent termination of the production of PHB shampoo bottles and razor handles.

Despite this, applications within the medical industry have been comparatively successful, as the non-toxic and biodegradable nature of PHB removes the need for follow-up surgeries to remove medical sutures, which will now decompose over time.

However, in general, such success is constrained by its low usage. With increasing petroleum prices and decreasing PHB production costs, it may be possible for PHB to one day to emerge on the market after further research and innovation. Certainly, its impact upon the environment will be significant.

1.2.3 Explain what is meant by a condensation polymer

Not all molecules condensed out during condensation polymerisation are water molecules, as in the polymerisation of cellulose.

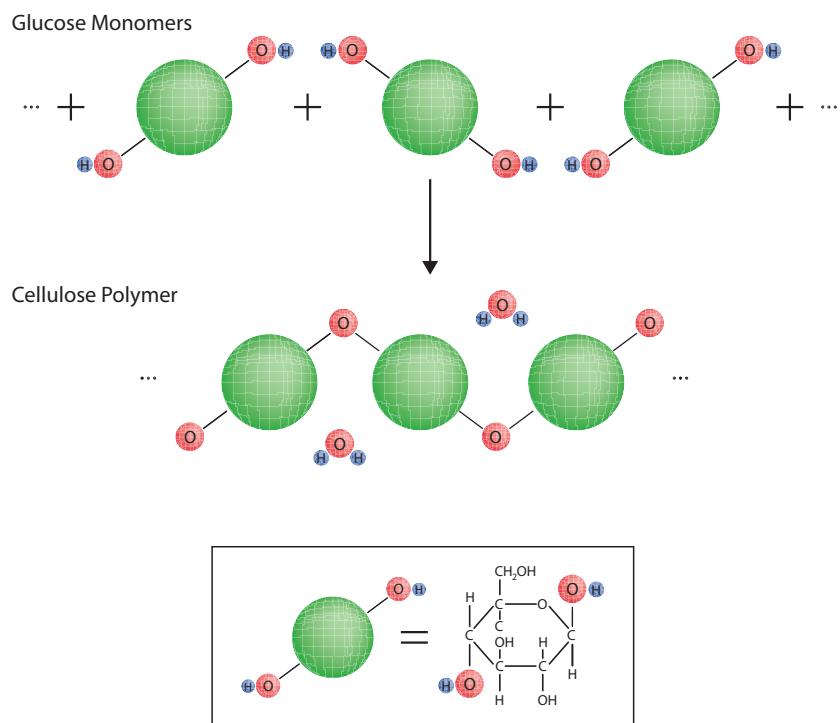
A condensation polymer is a polymer chain formed by the joining of monomer units which condense out small molecules as the polymer forms (Parts of the actual monomer detach in order to 'unlock' the monomer and enable polymerisation). One example of a condensation polymer is cellulose, which is formed from glucose monomers condensing out water molecules as they join.

Remember- A condensation polymer forms by joining together monomer units which have been 'unlocked' by simultaneously releasing small molecules.

1.2.4 Describe the reaction involved when a condensation polymer is formed

The easiest way to gain an understanding of a condensation polymer is by going through an example of how a common condensation polymer is formed. I will lead on from dotpoint 1.2.3 and use cellulose as an example.

In the formation of cellulose, glucose molecules join together in a chain, where in between every consecutive glucose molecule a hydroxyl group (-OH) from each glucose molecule condenses out as a water molecule, leaving a single oxygen linking the two monomers. This process occurs between every monomer unit which is added to the chain, forming the condensation polymer cellulose. Alternating monomers are inverted in the chain.



Don't forget to include the water molecule which is condensed out!

Remember- In the formation of cellulose, glucose molecules join by condensing out a water molecule between ever two glucose molecules, and inverting every alternate glucose molecule.

1.2.5 Describe the structure of cellulose and identify it as an example of a condensation polymer found as a major component of biomass

This dotpoint requires you to be able to differentiate between intermolecular bonding (Between molecules) and intramolecular bonding (Within a molecule). The structure of cellulose, in particular its insolubility, relates to the strength of the intermolecular bonds.

Due to the hydrogen bonding present within the molecule, cellulose is insoluble because the intermolecular forces cannot be easily broken. It is also important to note that when glucose monomers combine to form cellulose, every second glucose unit is effectively flipped upside down. This produces a reasonably linear molecule, increasing the density and strength of the molecule.

For the purposes of this dotpoint, it would also be useful to note that cellulose is a major component of biomass, where biomass can be defined as any material produced by living organisms. This most frequently refers to plant material and animal excreta.

Remember- Cellulose is a dense, insoluble condensation polymer commonly found in various forms of biomass such as grass, trees and trees. It is a condensation polymer and a major component of biomass.

1.2.6 Identify that cellulose contains the basic carbon-chain structures needed to build petrochemicals and discuss its potential as a raw material

This dotpoint is really leading up to the potential of cellulose as an alternative source of petrochemical products. The main point is to note the verb 'discuss', and argue the For and Againsts of using cellulose.

Cellulose contains a basic carbon-chain structure common to many of the compounds used within the petrochemical industry. Readily abundant and renewable, cellulose presents a tempting alternative to the non-renewable resources currently used such as petroleum.

However, although there are two primary methods of converting cellulose to its glucose components, acid digestion and enzyme digestion, both processes require an immense amount of energy in order to overcome the strong hydrogen bonds present within the cellulose structure. As a result of the energy input required, these processes simply aren't economically viable, and as such there are currently no means of converting cellulose into the traditional monomers used in the polymerisation of materials such as PVC and Teflon.

Producing cellulose is no longer a significant barrier, as modern bacterial production methods involving strains such as *E. Coli* present many opportunities in the mass production of cellulose-based substances. However, it is only when cellulose can be effectively broken down that it can be used effectively as an alternative to crude oil.

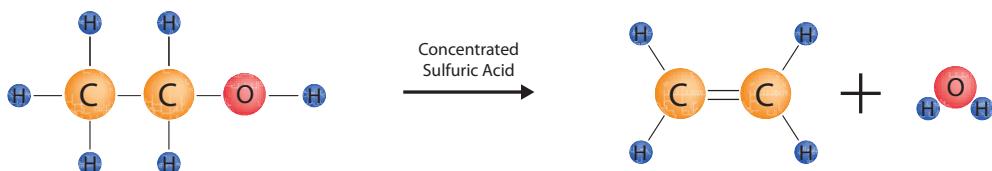
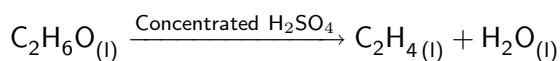
Although the inability to convert cellulose into its glucose components in any economically viable manner does prevent its use as an alternative source of petrochemical products, should this barrier be overcome cellulose has virtually unlimited potential given its abundance and chemical make-up.

1.3 Ethanol - Use and Manufacture

1.3.1 Describe the dehydration of ethanol to ethylene and identify the need for a catalyst in this process and the catalyst used

The ability to write the chemical equation is important, so take the time to remember what molecules are present in the reactants and in the products, and then balance as a last step.

Although ethanol can be readily dehydrated to form ethylene, a strong catalyst is required as the hydroxyl functional group (-OH) is bonded relatively strongly to the CH_3CH_2 chain. As such, a catalyst lowers the activation energy required for a chemical reaction to take place by providing an alternate pathway for the reaction to occur. Most commonly, concentrated sulfuric acid or phosphoric acid is used as the catalyst.



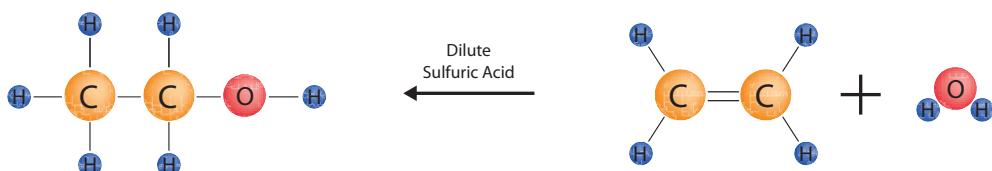
Water is a product as well, so don't forget to include it.

Remember- The dehydration of ethanol to ethylene involves the conversion of ethanol into ethylene and water using a catalyst of concentrated sulfuric acid.

1.3.2 Describe the addition of water to ethylene resulting in the production of ethanol and identify the need for a catalyst in this process and the catalyst used

Understanding why this reaction only requires a weak catalyst to occur will save you the pain of memorising the specific catalysts required. If you find yourself confused as to which catalyst is required (concentrated or dilute), simply ask yourself which bonds are easier to break: The reactive double bond in ethylene, or the strong hydrogen bonding in ethanol? Clearly a dilute acid is appropriate for the double bonds, whereas concentrated sulfuric acid would be appropriate for breaking the bonds in ethanol.

Just as ethanol may be dehydrated to form ethylene, the addition of water to ethylene will produce ethanol. However, unlike the dehydration process, the presence of the double bonds within ethylene considerably decreases the energy required in order for the reaction to occur. As such, a weaker catalyst such as dilute sulfuric acid may be used.

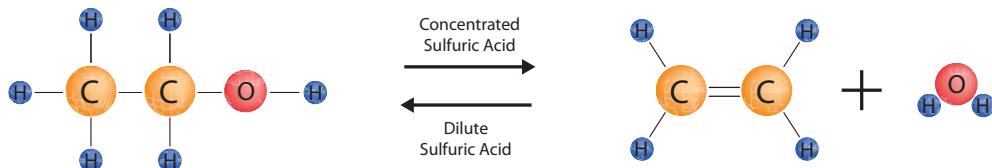
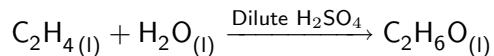
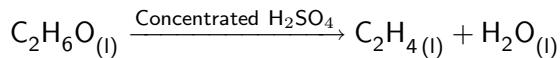


Once again, do not forget the presence of the water molecule, which is of course on the reactants side for this equation.

Remember- The hydration of ethylene to ethanol involves the conversion of ethylene and water into ethanol using a catalyst of dilute sulfuric acid.

1.3.3 Process information from secondary sources such as molecular model kits, digital technologies or computer simulations to model the addition of water to ethylene and the dehydration of ethanol

If you are able to write down these reactions from scratch, remembering the presence of catalysts and understanding why they match the given reaction, this dotpoint will prove little trouble. Below are the equations once again for your reference.



1.3.4 Process information from secondary sources to summarise the processes involved in the industrial production of ethanol from sugar cane

Nothing in this dotpoint exceeds what will be established in dotpoints 1.3.13 (page 22) and 1.3.14 (page 22). Hence the following steps are provided as a quick summarisation. The last step is common with the industrial production of ethanol.

1. Suitable crops such as sugar cane are cultivated and glucose is extracted.
2. A sugary mixture with yeast added is heated at room temperature in anaerobic conditions.
3. A concentration of 15% ethanol can be reached through this process, after which the yeast will no longer produce ethanol.
4. The solution may be distilled to obtain higher concentrations of ethanol.

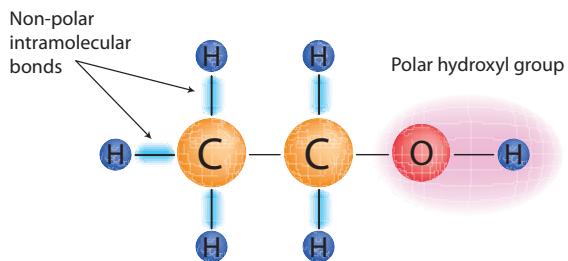
1.3.5 Describe and account for the many uses of ethanol as a solvent for polar and non-polar substances

This dotpoint will largely call upon fundamentals taught in the preliminary course such as the polarity of molecules, as well as the meaning of dipole-dipole bonds, or dispersion forces. Simply keep in mind the adage 'Like dissolves like' and there should be no problem. What this means is that a polar solution is likely to dissolve a polar substance, and the same with a non-polar solution.

The structure of ethanol, C_2H_5OH , accounts for its many uses as a solvent for both polar and non-polar substances.

The slightly polar hydroxyl functional group (-OH) allows ethanol to act as a solvent for polar substances, as electronegative species are able to dissolve via dipole-dipole interactions, or through hydrogen bonding. In addition, the dispersion forces (temporary induced dipole forces) present within the CH bonds combined with the hydrogen bonding present within the molecule serve to dissolve non-polar substances.

It is for this reason that ethanol is seen as the second-most important solvent after water.



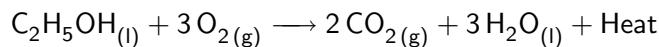
Remember- It is due to the presence of a polar, OH end as well as a non-polar CH end that ethanol is effective as a solvent.

1.3.6 Outline the use of ethanol as a fuel and explain why it can be called a renewable resource

Once again, be prepared to provide the formula for the combustion of ethanol. This will prove invaluable in demonstrating the low concentration of oxygen required for complete combustion (Of which the products are always carbon dioxide and water). Keep in mind many petrol stations in Australia already sell fuel with an ethanol content of approximately 10%, so this is a question that is very much relevant to current affairs.

Characteristics of ethanol which favour its use as a fuel include:

- The presence of an oxygen atom within each molecule, allowing for ethanol to burn relatively cleanly. The cleaner a fuel burns, the less harmful by-products are produced, such as carbon monoxide. Shown below, 3 moles of oxygen gas allow for the complete combustion of 1 mole of ethanol. In contrast, octane fuel requires 12.5 moles of oxygen gas. As such, the use of ethanol increases the chances of complete combustion, and carbon monoxide emissions can be reduced up to 20% if ethanol is substituted for octane.



- Ethanol is easily transportable.
- Ethanol has a heat of combustion of 1360 kJ/mol, giving a fairly high energy-per-mole output.
- Current motors can already accept a 10-20% mixture of ethanol mixed with current petrol. However, ethanol attracts water molecules, and without the implementation of engine modifications, the use of ethanol is restricted to this relatively low-ratio mix.

Ethanol is fermented from glucose, which consists of carbon dioxide and water. As such, sugar cane farming, which has high levels of glucose, is an effective primary source of the components used to produce ethanol. It is for this reason that ethanol is called a renewable fuel.

Remember- Ethanol's primary benefit comes from its nature as a renewable fuel, a fuel which can be reproduced without fear of diminishing resources. When combined with its high energy-per-mole output, cleaner burning nature, and ease of transport, ethanol has great potential as an alternative fuel.

1.3.7 Identify the IUPAC nomenclature for straight-chained alkanols from C1 to C8

Prefix	Number of carbon atoms
Meth-	1
Eth-	2
Prop-	3
But-	4
Pent-	5
Hex-	6
Hept-	7
Oct-	8

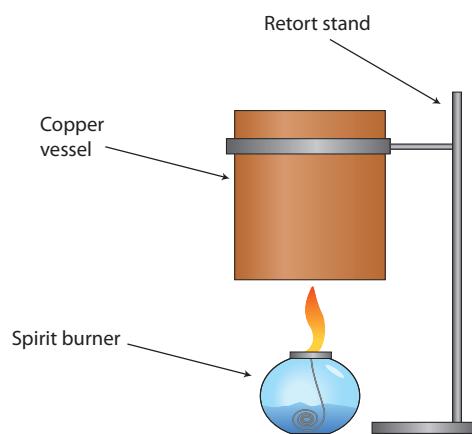
Naming a straight-chained alkanol is fairly simple if you focus on the number of carbon atoms present in the molecule. Above is a list of the prefixes for the first eight straight-chained alkanols.

1.3.8 Identify data sources, choose resources and perform a first-hand investigation to determine and compare heats of combustion of at least three liquid alkanols per gram and per mole

Understanding that the heat released by the fuel is theoretically equal to the heat absorbed by the water in the copper can is crucial to this experiment. However, you must also note that this experiment will often yield large errors, as heat will be lost to the environment. It is extremely likely that examination questions will raise this, so be prepared with methods of reducing the errors obtained. At the end of this dotpoint is a list of measures that you can go through, many applied in the procedure outlined above.

Materials

- Copper can
- Retort stand and ring clamp
- Glass rod
- 3 spirit burners filled with methanol, ethanol and 1-propanol
- 100mL measuring cylinder
- Thermometer
- Balance

Diagram:**Procedure:**

1. Light the first spirit burner.
2. Adjust the height of the can so that the tip of the flame just touches the can.
3. Replace the cap on the spirit burner to extinguish the flame. Do not blow out the flame in order to extinguish it.
4. Weigh the burner with its liquid contents and record.
5. Add 200 mL of cold water to the copper can using a measuring cylinder. Place the thermometer in the water and record its initial temperature.
6. Light the wick and stir the water gently with the stirring rod to ensure uniform heat.
7. Monitor the temperature and extinguish the flame by replacing the cap when the temperature has risen by 10°C. The thermometer should be kept halfway in the water.
8. Reweigh the burner.
9. Remove soot from the bottom of the can and replace the water in the copper can before testing the next alcohol.

Expected results:

Now,

$$\begin{aligned}\Delta H_1 &= \text{Heat absorbed by water (Joules)} \\ &= mC\Delta T\end{aligned}$$

where

$$\begin{aligned}m &= \text{Mass of water} \\ C &= 4.18 \text{ (Specific heat capacity of water)} \\ \Delta T &= \text{Change in temperature}\end{aligned}$$

Of course, the heat released by the fuel is equal to the heat absorbed by the water, ignoring loss of heat to the environment.

Therefore

$$\begin{aligned}\Delta H_2 &= \text{Heat released by fuel (Joules)} \\ &= -\Delta H_1\end{aligned}$$

per x grams of fuel used (i.e. A multiple of the heat released per gram). Thus,

$$\frac{\Delta H_2}{\left(\frac{\text{mass of fuel}}{\text{molecular mass of fuel}}\right)} = \text{Molar heat of combustion of the fuel}$$

Methods of reducing error when determining heats of combustion

- Ensure the tip of the flame touches the bottom of the can so as to minimise heat lost directly.
- Use a copper can in order to contain as much of the heat as possible.
- Enclose the experiment within a covering of aluminium foil in order to minimise the loss of heat into the surroundings.

Remember- The amount of energy absorbed by the water is equal to the amount of energy released by the fuel combusting (Ignoring loss of heat to the environment). However, this loss of energy can be considerably large, and thus large inaccuracies will occur in your findings. Note the above improvements to the experiment, as the reduction of such errors is integral to this dotpoint.

1.3.9 Define the molar heat of combustion of a compound and calculate the value for ethanol from first-hand data

The molar heat of combustion is the amount of energy released in the form of heat when one mole of a substance is combusted to form products in their standard states (solid, liquid or gas) at 100 kPa and 25°C (298K).

Ethanol's molar heat of combustion is 1360 kJ/mol.

Remember- The molar heat of combustion is simply the amount of energy released as heat from one mole of a substance. This is not indicative of the energy released by one unit of a substance (E.g., one gram, or one millilitre).

1.3.10 Process information from secondary sources to summarise the use of ethanol as an alternative car fuel, evaluating the success of current usage

This dotpoint serves largely to bring attention to current applications of ethanol as a fuel. Brazil is a great example for this, so be sure to mention it if it is relevant to the question.

As supplies of non-renewable sources of fuel are slowly dwindling, governments are gradually turning towards renewable sources such as ethanol. However, the lack of significant research in the field has led to considerable costs, costs which have deterred many from pursuing ethanol as an alternative fuel within the near future. Despite this, it is undeniable that there has been an increasing trend in the turn towards ethanol as an alternative car fuel.

The Brazilian government subsidised the production of ethanol during the 1970s in order to reduce oil imports and stimulate employment growth. By using sugar cane waste to produce ethanol, approximately a third of the motor vehicles in Brazil (over four million) were able to use pure ethanol as a sole source of fuel. Today, approximately 50% of Brazilian cars are able to use 100% ethanol as fuel. The majority of the remaining vehicles use a mixture containing at least 20% ethanol.

Currently, many countries (Including Australia) make use of fuel with 10-15% ethanol. Although further engine modifications are required before this concentration can increase, research and innovation coupled with rising petrol prices offer much incentive to see this change undergone as soon as possible.

Remember- Despite such potential, the viability of using ethanol as a fuel is largely limited by the costs of producing large quantities through industrial fermentation of glucose-high crops such as sugar cane.

1.3.11 Assess the potential of ethanol as an alternative fuel and discuss the advantages and disadvantages of its use

The advantages and disadvantages of ethanol as an alternative fuel will be listed out for the purposes of this dotpoint. Depending on the question asked, feel free to elaborate on these points.

If the question requires it, include a stance. Do this by acknowledging the potential of ethanol, but the current costs of producing ethanol, as well as the practical issue of redesigning engines to cope with higher concentrations of ethanol. It may be of use to note, and perhaps mention that vehicles in countries such as Brazil have achieved compatibility with 100% ethanol fuels.

Advantages:

- Ethanol is renewable, reducing the consumption of non-renewable resources.
- Greenhouse gas emissions could be reduced due to the cleaner-burning nature of ethanol, which could potentially reduce carbon monoxide emissions by up to 20%.

Disadvantages:

- Large areas of arable land would need to be devoted to the cultivation of suitable plants such as sugar cane. This would lead to erosion, deforestation, salinity, and many other environmental problems.
- Much of the biomass produced in the production of glucose is not used in the actual fermentation process. Unless other uses are found for such products, this can present major environmental problems.
- The process requires a lot of energy, energy which is currently provided by fossil fuels. In some cases, the energy used to produce ethanol via the combustion of fossil fuels can be higher than the energy produced by the resulting ethanol.
- As a result of requiring larger inputs of energy, ethanol is currently more expensive to obtain than petrol.
- Although ethanol burns more cleanly, whether or not it reduces the greenhouse effect is debatable, as many greenhouse gases are released during the harvesting of the crops used to produce ethanol as well as the distillation of the product.
- Ethanol reduces less energy per mole than octane.
- Engine modifications are required if ethanol is ever to be used in compositions larger than 20%.

1.3.12 Solve problems, plan and perform a first-hand investigation to carry out the fermentation of glucose and monitor mass changes

This experiment is fairly simple, only requiring some time for it to carry out to completion. As such, try to attempt this experiment early on in the day so that hourly observations can be made. Always keep in mind in any experiment methods of improving the experiment in terms of reliability and accuracy, and these inevitably appear in exams. As such, keep in mind possible errors which may arise in simple tasks such as measuring weight, and formulate possible solutions such as using an electronic scale, or perhaps increasing the sample size.

Procedure

1. Place warm water, glucose, and yeast in a flask and stopper the top with either a cork, or cotton wool.
2. Weigh the flask and its contents and record the results.
3. Place the flask in a warm area such as the window sill, and allow to sit. Ideally, set this experiment up in the morning, and take measurements regularly (In hourly intervals) over four hours, recording the weight at each observation.
4. Reweigh the flask and its contents, and record the results along with the difference in weight.

Optional: Connect a cork and rubber tubing to the flask, running the end of the tubing into a flask of limewater. The limewater turning cloudy indicates the presence of carbon dioxide.

Expected results:

The loss in mass should be equal to the carbon dioxide produced. The end solution is likely to be an extremely dilute solution of ethanol.

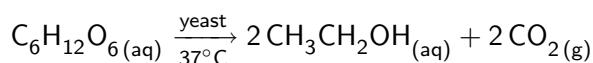
If you monitored the mass changes at regular intervals, you will notice that the rate of change decreases over time, gradually coming to a complete stop after a few hours.

1.3.13 Describe conditions under which fermentation of sugars is promoted

Several other factors are often included in various publications. However, the four points below are really the only crucial points necessary to gain full marks. As a point of interest, the yeast itself is not required, but rather an enzyme secreted by yeast known as zymase is what actually brings about the fermentation process. However, simply stating yeast as a requirement is fine.

Fermentation is simply the process by which glucose is broken down to form ethanol and carbon dioxide. Several conditions must be satisfied in order for fermentation to successfully take place:

- A glucose-containing solution must be present. If a solid containing glucose is used, water must be added to form a solution.
- Yeast must be present.
- The reaction must be conducted in anaerobic conditions, i.e. no oxygen.
- A temperature roughly at body temperature is best, i.e. 37°C.

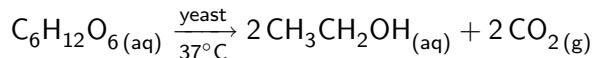


Although a temperature of 37°C is not necessary for the reaction, which will occur at lower temperatures, a higher temperature will cause the reaction to occur faster. However, the temperature should not be raised above 37°C, as the yeast cannot survive.

Remember- The conditions necessary for fermentation to occur: A glucose solution, yeast, anaerobic conditions, and a temperature of approximately 37°C.

1.3.14 Summarise the chemistry of the fermentation process

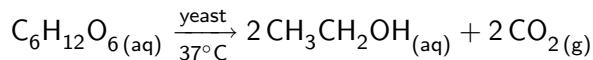
The yeast included in the reaction secretes an enzyme complex known as zymase which catalyses the process of glucolysis, which refers to the conversion of glucose into ethanol and carbon dioxide. An ethanol percentage of 15% can be successfully produced, at which point the yeast will begin to die due to the concentration of ethanol, and the reaction will halt.



Remember- Fermentation is the process whereby glucose is converted into ethanol and carbon dioxide.

1.3.15 Present information from secondary sources by writing a balanced equation for the fermentation of glucose to ethanol

This dotpoint will have been covered quite sufficiently by this point, but as the equation is quite important, it is written again below. Do not forget to include the necessary condition of 'yeast' when writing the equation.



1.4 Energy from Redox Reactions

1.4.1 Perform a first-hand investigation to identify the conditions under which a galvanic cell is produced

This experiment is relatively easy, with simple results that are relatively easy to obtain. The only problem that may occur is incorrectly setting up the electrodes, so take care to ensure that the more reactive metal is being used as the anode. If not, the galvanic cell will not generate a current.

Procedure:

Set up the galvanic cell as shown in the diagram in dotpoint 1.4.2 on page 24.

Expected results:

There are a number of conditions which must be met in order for a galvanic cell to function.

- There must be an anode and cathode, each within their respective electrolytic solutions and physically separated.
- A conducting circuit must connect the two half-cells to provide the only path by which electrons can flow.
- A salt bridge must exist. The purpose of the salt bridge is to maintain electrical neutrality, as ions migrate from half-cell to half-cell in order to maintain a balance in the charges. If a charge imbalance built up (as would occur if a salt bridge was not present), the cell would stop functioning, as ions will accumulate in both electrolytes until the potential difference due to the ions is exactly opposite to the potential difference from the reaction. Thus the flow of electrons would be negated, and there would be no current.

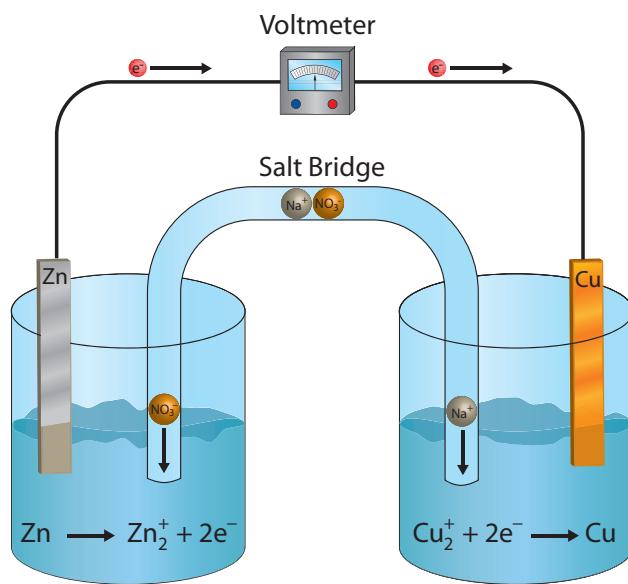
Remember- The salt bridge exists to maintain electrical neutrality. As such, given that ions will be passing through the bridge and into the relevant electrolytes, ensure that a precipitate will not form. For this reason, NaNO_3 is often used to soak the salt bridge, as both Na^+ and NO_3^- will not form any precipitates regardless of which ions are present in the electrolyte.

1.4.2 Perform a first-hand investigation and gather first-hand information to measure the difference in potential of different combinations of metals in an electrolyte solution

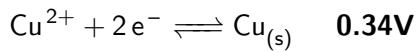
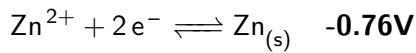
This experiment is fairly useful for gaining an idea of how a galvanic cell works. Simply arrange the cell as shown in the diagram above, and connect a voltmeter to the external circuit between the two electrodes. Although this dotpoint does suggest a practical approach to determining the difference in potential of the different combinations of metal, for the purpose of clarity, a theoretical approach will be adopted here. You will be required to write both half equations and full equations, so take the time to learn how. Fortunately, the HSC data sheet lists pretty much all combinations you will be required to know, so it is simply a matter of choosing the right half equations.

With any given reaction that occurs in a galvanic cell, there will be two half-equations. Each half-equation has a standard potential E° (pronounced 'e-naught') value, a list of which will usually be provided at the back of the exam paper where necessary.

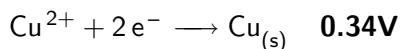
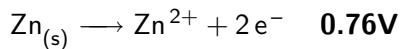
Although not essential, it may be useful to know that all E° values are calculated as the potential differences using a standard hydrogen electrode. This allows for a more convenient calculation of figures rather than recalculating the difference in potential between a pair of electrodes every time. You can see this quite easily by referring to the HSC data sheet, where the hydrogen electrode half equation has a potential of 0V.



Using a galvanic cell of zinc and copper as an example, the following half-equations apply:

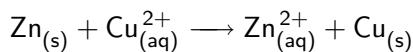


Flipping the oxidation half-equation around, i.e. at the zinc electrode, and keeping the positions of the reduction reaction as is:



Why did we flip the oxidation half-equation around? Because the more reactive metal displace the less reactive metal and releases electrons. Hence the electrons are on the right side of the equation in the oxidation equation, while on the left side of the equation in the reduction equation.

Adding together the two half equations, the difference in potential of the two metals in this case is 1.10V. This should correspond to the value obtained when conducting the experiment. The full equation is then



Remember- The more reactive metal must thus be the anode, and is found above the cathode on the list of standard potentials in the data sheet, as the data sheet is sorted in descending order of reactivity.

In addition, remember to balance the half equations by their number of electrons. This is important as for some pairs such as sodium and zinc, sodium will have one electron in the products, yet zinc will have two electrons in the reactants. As such, double all species in the sodium half equation before writing the full equation.

1.4.3 Explain the displacement of metals from solution in terms of transfer of electrons

If you have difficulty understanding what is happening at this point, especially with the different relative activities, don't worry too much as the next few dotpoints do make it clearer. For now, simply understand that a displacement reaction involves a transfer of electrons.

A displacement reaction, where metals are displaced from a solution, is a reaction in which there is a transfer of electrons between a metal and a metal ion. This occurs because each metal has a different relative activity.

Remember- Ionisation simply means that an atom is gaining or losing electrons. As such, if a metal ion becomes an atom, and a metal atom becomes an ion, then clearly all that has occurred is a transfer of electrons.

1.4.4 Identify the relationship between displacement of metal ions in solution by other metals to the relative activity of metals

Exactly which metal displaces what is often the main question that students will find themselves asking. This is conveniently summarised at the end of this dotpoint.

Of two different metals, the more reactive metal, X, will displace the other metal, Y, when it is an ionic solution. How reactive a metal is can be measured by how easily it oxidises. Potassium and sodium are highly reactive as they oxidise very easily.

Electrons will transfer from the metal X to the solution Y, resulting in X becoming a positive ion in solution, and Y turning back into a metal atom.

Remember- The more reactive metal displaces the less reactive metal.

1.4.5 Account for changes in the oxidation state of species in terms of their loss or gain of electrons

This dotpoint is really just introducing the concept behind redox reactions. Note that when one species loses electrons, another gains the same amount. The mnemonic at the end of this dotpoint is a simple way of remembering which process is what.

A change in the oxidation state of a species simply refers to its loss or gain in electrons. When a species is oxidised, it loses electrons (And thus its charge becomes more positive). When a species is reduced, it gains electrons (And thus its charge becomes more negative).

Reduction and oxidation reactions, or 'redox' reactions, occur simultaneously rather than independently of one another. As such, the electrons lost by one species are gained by another.

Remember- A simple mnemonic commonly used to remember this is 'OILRIG', or 'Oxidation is Loss, Reduction Is Gain'.

1.4.6 Describe and explain galvanic cells in terms of oxidation/reduction reactions

Creating a galvanic cell is a great way of understanding what is happening here. Barring that, the diagram in dotpoint 1.4.2 on page 24 should provide a useful point of reference.

A galvanic cell converts chemical energy into electrical energy. Several things must be noted regarding galvanic cells:

- Two electrodes (An oxidant and a reductant) , The anode releases electrons through oxidation reactions, which flow into the cathode through reduction reactions. These two electrodes are kept physically separated, but are joined by an external circuit so as to allow a charge to flow.
A common mix-up that tends to confuse some students is what exactly an oxidant or reductant is. Simply keep in mind that an oxidant is a substance that is reduced, and a reductant is a substance that is oxidised. Oxidants and reductants are also known as oxidising agents and reducing agents respectively.
- Appropriate electrolytes to allow the flow of electrons within the two half-cells.
- A salt bridge, used to allow the migration of ions to maintain a balance of negative and positive charges in each half-cell. Anions migrate to the anode, and cations migrate to the cathode. This occurs in order to offset the flow of electrons to and from the cathode and anode respectively, thereby maintaining electrical neutrality.

Remember- A galvanic cell is powered by the flow of electrons, i.e. redox reactions. So long as the two metal electrodes are present (physically separated by joined by an external circuit), appropriate electrolytes are present, and a salt bridge spans between the two electrolytes, then a current will flow.

1.4.7 Outline the construction of galvanic cells and trace the direction of electron flow

Although this dotpoint seeks only for a brief explanation of the structure of a galvanic cell, I believe the image in dotpoint 1.4.2 on page 24 will be infinitely more useful in explaining the operation of the galvanic cell. Learn the significance of each part, as well as how to draw it, and the construction of the galvanic cell will be simple to describe.

If you have difficulty remembering what reaction occurs at which electrode, use the term 'AnOx' to remember that oxidation occurs at the anode. Once you remember this, it follows that at the cathode, the reduction reaction occurs. Otherwise, 'Red Cat' can be used to remember that reduction occurs at the cathode.

1.4.8 Define the terms anode, cathode, electrode and electrolyte to describe galvanic cells

Again, learn how to draw the galvanic cell, and you will find recounting each part of the galvanic cell a relatively simple task.

- The anode is the electrode at which the oxidation reaction occurs, releasing electrons into the external circuit.
- The cathode is the electrode at which the reduction reaction occurs, consuming electrons which have been released into the external circuit.
- An electrode is a conductor connected to the external circuit through which electrons may pass.
- An electrolyte is a substance, either a solution or a solid in a molten state, which accommodates the flow of electrons and hence conducts electricity.

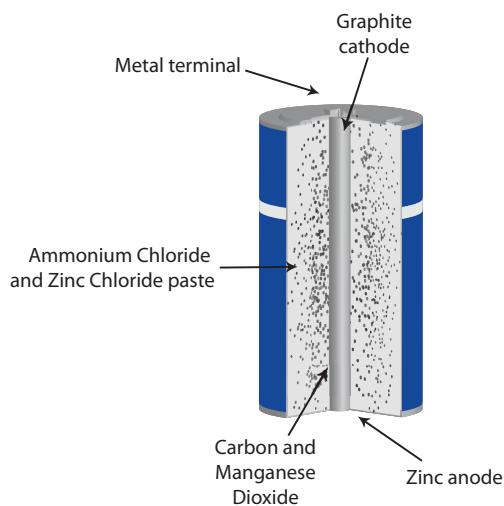
1.4.9 Solve problems and analyse information to calculate the potential requirement of named electrochemical processes using tables of standard potentials and half-equations

Calculating the E° value, pronounced 'E-naught', is simply a matter of finding the difference in potential already covered in dotpoint 1.4.2 on page 24.

1.4.10 Gather and present information on the structure and chemistry of a dry cell or lead-acid cell and evaluate it in comparison to one of the following: button cell, fuel cell, vanadium redox cell, lithium cell, liquid junction photovoltaic device (eg. the Gratzel cell) in terms of: chemistry, cost and practicality, impact on society, environmental impact

Unfortunately this dotpoint, even if you understand the concepts behind the galvanic cell, is ultimately an exercise in memorisation. Choosing any of the above cells is fine, but for the purposes of this guide, the dry cell and the silver oxide button cell will be examined.

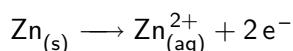
The Dry Cell (Also known as the Leclanché Cell)



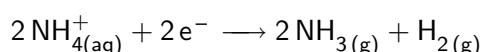
Chemistry

Zinc casing and a graphite rod serve as the anode and cathode respectively, and the electrolyte consists of a paste of ammonium chloride and zinc chloride is the electrolyte. Manganese dioxide is also present within the cell.

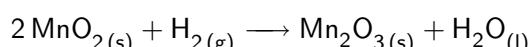
The zinc anode oxidises according to:



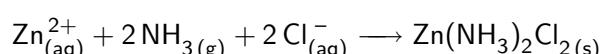
The electrons produced reduce the ammonium ion at the cathode.



Manganese dioxide converts the hydrogen gas to water as it is an oxidising agent.



The ammonia gas reacts with the Zn^{2+} ion.



These reactions have a net potential of 1.5 volts.

Cost and practicality

The dry cell is one of the most common and cheapest of the commercially available cells and is widely used in small or portable devices. Disadvantages of the cell are that it does not give a very large output given its size (both voltage and currents) and can leak when the zinc casing anode is gradually corroded. Relative to other cells, the dry cell also has a short shelf life, as the acidic ammonium ions reacts with the zinc casing. As this occurs, the voltage drops rapidly, after which the battery is said to go 'flat'.

Impact on society

The dry cell was the first commercially available battery, and as such most of the earliest devices were modelled around it. Offering a small, portable power source, the dry cell had an enormous impact on society, and still offers a useful power source for low-current devices such as torches.

Environmental impact

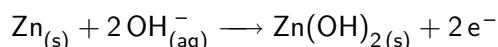
The dry cell is relatively safe for the environment. Manganese (III) oxide oxidises to the harmless manganese (IV) oxide, ammonium salts are harmless, and carbon is also relatively harmless. Although large quantities of zinc may pose environmental problems, particularly when it leaches into the soil, small quantities generally do pose a problem.

The Silver Oxide Button Cell

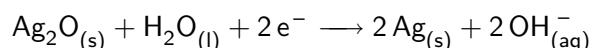
The silver oxide button cell has essentially the same structure as the dry cell, but uses different chemicals inside a smaller case.

Chemistry

In this cell, zinc and silver oxide serve as the anode and cathode respectively. An alkaline substance, usually either sodium hydroxide or potassium hydroxide, serves as the electrolyte. The following oxidation half-equation takes place:



Silver oxide then reduces to silver:



These reactions have a net potential of 1.5 volts.

Cost and practicality

The silver oxide button cell has found its ideal use within cameras and watches, due to its small size but relatively high (1.5V) output. Its ability to produce a very steady output is also ideal for medical equipment such as pacemakers and hearing aids. In such instruments, its size and reliability take precedence over cost.

Impact on society

Considerably smaller than the dry cell while matching its output, the silver oxide button cell has had a large impact on society as smaller devices are able to make use of the batteries. The high reliability of the cell has also had a large impact upon the medical industry, as the life spans of various medical devices such as pacemakers have been extended. The size of the cell has also drastic implications, most noticeably through everyday items such as watches.

Environmental impact

As with the dry cell, small amounts of zinc do not pose an environmental problem. Similarly, silver oxide cells do not produce any highly toxic wastes. This is in contrast with the mercury button cell, which is more effective through its control of zinc corrosion, yet has a much more harmful effect upon the environment due to the use of mercury.

All batteries still have an anode, cathode, and various electrolytes even if they look markedly different from other batteries. You will be required to write down all relevant half equations, so take care to learn them thoroughly.

1.5 Nuclear Chemistry

1.5.1 Distinguish between stable and radioactive isotopes and describe the conditions under which a nucleus is unstable

The definition of an isotope should have been encountered in the preliminary course, but is stated again below for convenience. Note that an unstable nucleus is really one that is 'falling apart', giving off radiation as it does so. This occurs for many reasons, but largely because the atom is growing too large for the nucleus to hold together. Keep this in mind, and the exact definition as well as the criteria listed below should prove easier to understand.

An isotope of an element is an atom with the same number of protons, but a different number of neutrons. For example, one isotope of hydrogen may have 1 proton and 1 neutron, but another isotope may have 1 proton and 2 neutrons.

Radioactive isotopes (commonly referred to as radioisotopes) have an unstable nucleus due to the particular number of neutrons they have, and emit radiation as they spontaneously disintegrate due to their unstable nuclei.

Instability will generally occur if

- The atomic number of the element is greater than 82, where lead is the 82nd element.
- The n:p ratio (neutron to proton ratio) lies outside the zone of stability, which is 1:1 for elements with atomic numbers less than 20, and increasingly greater than 1 for higher atomic numbers.

Radiation	Symbol	Identity	Relative Charge	Relative mass	Penetrative Power	Formula
Alpha	α	Helium nucleus	+2	4	Low	${}^4_2\text{He}$
Beta	β	Electron	-1	1/2000	Moderate	${}^0_{-1}\text{e}$
Gamma	γ	Electromagnetic radiation	0	0	High	

Remember- A radioisotope is simply an isotope of an element with an unstable nucleus. This means that its neutron to proton ratio is not within the zone of stability.

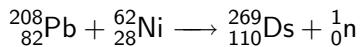
1.5.2 Process information from secondary sources to describe recent discoveries of elements

Be prepared to name at least one recent discovery such as Darmstadtium for the purposes of this dotpoint

Of the 25 transuranic elements to be created, only the first three were produced within nuclear reactors (Those with the atomic numbers 93, 94, and 95). The remaining transuranic elements were created by accelerating a small nucleus within a particle accelerator to collide with a heavy nucleus. New discoveries are hard to verify, as some have life spans significantly less than one second.

One transuranic element which has been created is Americium, which is produced by the bombardment of Pu-239 with neutrons. Americium is often used in smoke alarms.

One transuranic element which has been discovered far more recently is darmstadtium, an element discovered in Darmstadt, Germany. Previously known as ununnilium, darmstadtium has an atomic number of 110 and is produced by bombarding Lead-208 with Nickel-64. This radioisotope decays within microseconds as it is highly unstable, with its more stable isotopes such as Darmstadtium-281 having a half-life of around 11 seconds.



Although many other radioisotopes are commonly used in society, this dotpoint requires a 'recent' discovery. As such, darmstadtium is a safe option, as it was only verified by IUPAC within the last decade.

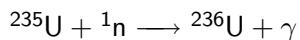
1.5.3 Describe how transuranic elements are produced

Transuranic elements are all elements with an atomic number of 93 or higher, and all have unstable nuclei. Remembering how one transuranic element is produced, perhaps neptunium, is highly recommended.

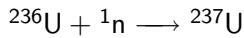
Transuranic elements are elements with an atomic number over 92, i.e. past that of uranium. As uranium is the heaviest natural element, all transuranic elements are artificially produced. These are produced by bombarding certain nuclei with neutrons. Some isotopes will ‘split’ when hit by the neutrons in a process known as fission, while others will ‘absorb’ the neutron, resulting in a larger atomic weight.

More recently, transuranic elements have been produced through the use of machines known as cyclotrons, or linear accelerators. In these cases, a high speed, positively charged particle such as a helium or carbon nuclei is bombarded against a larger nuclei.

For example, in the production of neptunium: Uranium-235 is first bombarded with neutrons to form Uranium-236



Further neutron capture creates Uranium-237, which then decays to form Neptunium-237 through beta decay:



The mass of the reactants should equal to the mass of the products. The same holds true for the atomic numbers of both sides.

Remember- Transuranic elements are produced by either fission through neutron bombardment or cyclotrons.

1.5.4 Describe how commercial radioisotopes are produced

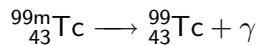
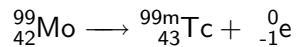
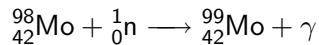
This dotpoint closely resembles the previous dotpoint 1.5.3, but focuses rather on the production of commercial radioisotopes. As such, be prepared to list at least one example. Although this may seem demanding now, you will find that you will need to recall at least two radioisotopes and their relevant methods of production by the end of this topic. As such, this dotpoint should pose little trouble if the later dotpoints are learnt satisfactorily.

On a commercial level, nuclear reactors are often used. Suitable target nuclei are placed in the reactor core and bombarded with neutrons to produce the desired isotope. Sometimes this isotope may decay further into other isotopes.

As mentioned in the previous dotpoint, cyclotrons may also be used to create radioisotopes. However, such machines are significantly more costly.

In the commercial production of radioisotopes, consideration must be given to the life span of the radioisotope. As such, reactors such as Lucas Heights, presently run by ANSTO, the Australian Nuclear Science and Technology Organisation, may supply Sydney destinations with radioisotopes, but destinations further away may not be able to receive such isotopes due to their short half-lives.

One example of a commercial radioisotope is Technetium-99, which is produced in Lucas Heights. It is produced by bombarding Molybdenum-98 with a neutron to form Molybdenum-99. Molybdenum-99 then decays via beta decay to form Technetium-99.



Particle accelerators are also used to produce commercial radioisotopes. Technetium-99 can also be produced this way: Heavy hydrogen (An isotope of hydrogen, ${}^2\text{H}$, also known as Deuterium) is accelerated and bombarded onto target Molybdenum-98, producing Technetium-99 and an excess neutron. This is done in the Royal Prince Alfred Hospital in the National Medical Cyclotron.

The main difference between nuclear reactors and particle accelerators is that reactors produce neutron rich isotopes, and accelerators produce neutron deficient isotopes.

Many radioisotopes decay extremely quickly. As such, commercial production facilities must be located relatively close to where they will be used, and importing or exporting radioisotopes is altogether impossible for many.

Remember- Commercially, most radioisotopes are produced in nuclear reactions or cyclotrons.

1.5.5 Identify instruments and processes that can be used to detect radiation

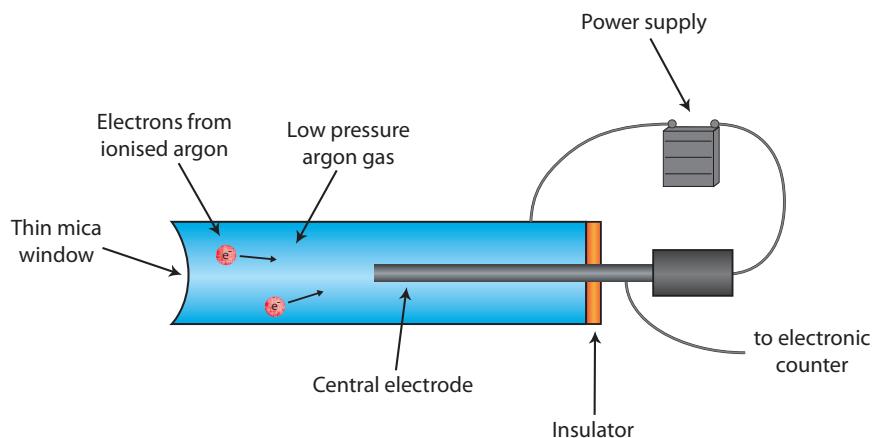
You will often be expected to be able to list at least three methods of detecting radiation, giving one or two in detail. Be sure you understand what the question requires from you, as some may simply require a list, whereas others may require a detailed explanation.

'Photographic film' offers a simple method of detecting radiation, as it darkens when in contact with radiation. This is due to a reaction within the silver halide crystals within the film. It is common for handlers of radioisotopes to wear radiation badges made of photographic film as a security precaution.

The 'scintillation counter' is a method of detecting non-ionising radiation (low energy radiation unable to ionise atoms). The radiation transfers energy to a solvent molecule, and then to a fluorescent molecule in order to give off light. Light is given off as excited electrons jump up to higher-energy shells when energy is absorbed, and give off light as they return to their ground state. This light is then passed through a photomultiplier, which in turn emits an electrical pulse, which is recorded by a counter.

The 'cloud chamber' is another means of detecting radiation. It consists of an air space with supersaturated water or alcohol vapour. When radiation passes through the device, the air is ionised, with the ions serving as nucleation points upon which the vapour may form droplets. These droplets appear as 'clouds' inside the chamber. Different forms of radiation which form different paths as they travel through the chamber. Alpha particles will leave straight lines, beta particles will leave a fainter zig-zagging path, and gamma rays will leave an even fainter path.

The 'Geiger-Muller' tube is an effective way of detecting ionising radiation. The radiation enters through a thin mica window at the end of the tube, and ionises a gas molecule within. The electron knocked out then accelerates towards the central electrode, ionising more gas as it proceeds. The molecules shed further electrons and become positive ions, moving towards the negative outer casing. The flow of electrons forms an electrical pulse when they come into contact with the central electrode. This pulse is amplified and used to generate clicks in an audio amplifier, or measured using an electronic digital counter.



1.5.6/1.5.7 Identify one use of a named radioisotope in industry and in medicine (including 'Describe the way in which the above named industrial and medical radioisotopes are used and explain their use in terms of their chemical properties')

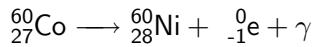
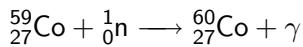
The above two dotpoints, 1.5.6 and 1.5.7 have been answered together for convenience. While you are free to choose other radioisotopes, the two examined here are great examples to use.

Industry- Cobalt 60

Cobalt-60 is used for several specialised purposes in an industrial context. One example relates to its use in gauging metal thickness, and another is to find faults within objects such as metal pipes. Both of these uses are relatively similar, as both rely upon the detection of the gamma rays emitted by the radioisotope. A source of Cobalt-60 is placed on one side of the object within a sealed container, and photographic film is placed on the other side. Both variances in thickness and the presence of defects in the metal are then identified by the changes in the level of radiation exposure indicated on the photographic film. This is caused by a darkening of the silver halide crystals as the gamma rays strike the surface of the film.

With a half-life of 5.3 years, Cobalt-60 can be left within a container inside metal pipes and other parts so that routine maintenance checks can be carried out without a frequent requirement to replace the radioisotope. Relatively low emission of radiation also limits the potential damage to anyone working with the radioisotope.

The production and decay of Cobalt-60 is shown in the following equations:

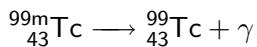
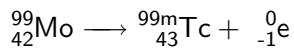
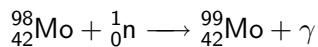


Medicine- Technetium 99m

Technetium 99m is used for more than 80% of all medical tracer diagnosis. Emitted gamma radiation can be picked up by equipment and converted into an image on a monitor, effectively ‘tagging’ the desired area.

With a half-life of 6 hours, Technetium 99m is ideal for such use, as the patient’s exposure to the radiation is minimised. The energy released is also low energy gamma radiation, minimising any tissue damage. However, most useful of all is Technetium 99m’s ability to be combined with other compounds to study different areas of the body. For example, it can be combined with tin to attach itself to red blood cells to examine the heart and blood vessels. Technetium 99m is not used by any part of the body, and is therefore not absorbed, further minimising radiation exposure. This makes it suitable to study highly radiation-sensitive organs such as the brain, kidney, bones, liver, and spleen.

The production and decay of Technetium-99m is shown in the following equations:



Note that $^0_{-1}\text{e}$ is beta ‘radiation’.

Also note that radioisotopes are chosen after considering a variety of factors. As a general rule, the half-life (And thus length of exposure), penetrating power, and chemical properties of the substance are the three main questions which must be covered. Relate the use of the radioisotope to each category.

1.5.8 Use available evidence to analyse benefits and problems associated with the use of radioactive isotopes in identified industries and medicine

You will find that, as with most of this course, the following points place a heavy reliance upon common sense. As such, learn the previous dotpoints well, particularly taking time to note a radioisotope used in medicine and industrially, and you will find that this dotpoint is relatively simple to answer.

Benefits:

- Medical applications of Tc-99m have reduced costs and provided a convenient, non-invasive method of diagnosis.
- A greater understanding of diseases and infections are possible through the examination of images which can be obtained by sensors monitoring the radioisotope travelling through a patient.
- Within an industrial context, tracer have provided a simply, cheap, and effective method of gauging metal thickness and identifying structural defects.
- Isotopes can be chosen with half-lives appropriate to their purpose. In medicine, short half-lives are desired so that the radiation exposure is minimised, whereas longer half-lives may be desired in industrial settings to avoid the need to frequently replace the radioisotope.

Problems:

- Radiation can cause damage to organic tissue, disrupting normal cellular processes, DNA, and proteins, potentially leading to abnormalities such as tumours, genetic mutations, or cancer.
- Some radioisotopes are chemically similar to elements within the human body. For example, Strontium-90 is chemically similar to calcium, and may replace the calcium within bones, potentially causing leukemia.
- Further problems may exist. Without further research, the long-term effects of certain types of radiation, such as the irradiation of food, are unknown.
- Many radioisotopes are costly to produce, and entirely unrealistic to transport as their use is highly limited by their half-lives.

Remember- Much of the allure of using radioisotopes industrially and in medicine is that the benefits are proven, yet the problems have not been fully identified. Be prepared to note this point when conducting any analysis of the use of any current or future radioisotopes

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Chapter 2

The Acidic Environment

2.1 Indicators

2.1.1 Perform a first-hand investigation to prepare and test a natural indicator

Various fruits, vegetables, and other items can be used to prepare indicators which are quite effective. This study guide will use red cabbage as an example, outlining the procedure you would take to prepare the indicator, and results you should expect.

Procedure:

1. Cut or grate a portion of a red cabbage, placing the small pieces into a large beaker and add enough boiling water to the beaker to cover the red cabbage (or use room-temperature water and heat with a Bunsen burner).
2. Allow the red cabbage pieces to boil for several minutes, not stopping until the water has become a dark, reddish or purplish colour.
3. Decant or filter the mixture into another beaker. Record the colour of the indicator obtained.
4. Test various substances with the indicator and record any changes in colour. Some substances you may wish to test include baking soda, vinegar and house salt (NaCl).

Expected results:

The colour of the starting indicator will vary according to several factors such as water pH, but it will generally be a dark reddish or purplish colour.

When an acidic substance is added to the indicator, the resultant colour should be pink. When a basic substance is added to the indicator, the resultant colour should be green. Between these two colours, the colour should be purple.

An accurate range will not be expected of you. Simply know the basic colours at both ends (Pink-Green).

Remember- Boiling red cabbage yields a purple indicator which turns pink in acid, and green in base.

2.1.2 Identify data and choose resources to gather information about the colour changes of a range of indicators

The four indicators litmus, phenolphthalein, methyl orange, and bromothymol blue will generally be enough for testing you may require, but to satisfy this dotpoint are a few more indicators and their respective colours and ranges.

Indicator Name	pH Range	Lower-range colour	Upper-range colour
Alizarine Yellow	10.2-12.0	Yellow	Red
Bromocresol Green	3.8-5.4	Yellow	Cyan
Methyl Red	4.4-6.2	Red	Yellow
Phenol Red	6.8-8.4	Yellow	Red
Thymol Blue	1.2-2.8	Red	Yellow

With multiple indicators, you can narrow the possible pH range quite easily. For instance, if a substance turns blue when bromothymol blue is added, and is colourless when phenolphthalein is added, then the pH must be between 7.6 and 8.3.

2.1.3 Identify that indicators such as litmus, phenolphthalein, methyl orange and bromothymol blue can be used to determine the acidic or basic nature of a material over a range, and that the range is identified by change in indicator colour

I don't recommend remembering all the indicators listed below. Remembering phenolphthalein and bromothymol blue should suffice.

An indicator is a substance which, in solution, changes colour depending on the pH of a substance. However, indicators do not quantitatively measure a substance's pH by itself, as a change in colour only gives a loose guide as to the nature of the substance. Generally, an indicator will only show you if a substance is an acid or a base, but will not tell you how strongly acidic and basic the substance is. An exception to this is universal indicator, which can be used to give a relatively accurate indication of a substance's nature.

Below is a table of the indicators litmus, phenolphthalein, methyl orange, and bromothymol blue, showing their ranges and the different colour transitions each one is subject to.

Indicator Name	Range	Lower-range colour	Upper-range colour
Litmus	5.0- 8.0	Red	Blue
Phenolphthalein	8.3- 10.0	Colourless	Red
Methyl Orange	3.1- 4.4	Red	Yellow
Bromothymol Blue	6.0- 7.6	Yellow	Blue

An effective indicator is one that allows you to easily distinguish between an acid and a base. One that gradually changes (Large range) is not as useful as one with a noticeable point at which it changes (Small range).

Remember- Indicators are substances which change colour depending on whether they are in an acidic or basic environment.

2.1.4 Classify common substances as acidic, basic or neutral

Broadly, acids are substances which generally taste sour, and will corrode metal. Stronger concentrations of acid will also give a burning feeling upon contact. Acids can conduct electricity.

Examples of common substances which are acidic in nature include: Lemon juice, vinegar, formic acid from ant and bee stings, as well as common acids within the classroom such as carbonic, hydrochloric, and sulfuric acid.

Bases are often bitter tasting, and slippery to the touch. Like acid, stronger concentrations will often give a burning sensation if it comes into contact with skin. Bases also conduct electricity.

Examples of common substances which are basic in nature include: Ammonia, baking soda, soap, toothpaste, antacids, and various household detergents and cleaning solutions, as well as common bases within the classroom such as sodium hydroxide, potassium hydroxide, and calcium carbonate.

Examples of neutral substances include common table salt (NaCl) and pure water.

Remember- Acidic substances include lemon juice and vinegar. Neutral substances include table salt and water. Basic substances include baking soda and detergents.

2.1.5 Solve problems by applying information about the colour changes of indicators to classify some household substances as acidic, neutral or basic

This is simply a practical version of dotpoint 2.2.3 on page 46. Take the time to understand how indicators work- particularly when multiple indicators are used to narrow the possible pH range of a substance down- and this dotpoint should prove no problem.

Possible results

Adding Phenol Red to lemon juice turned the indicator yellow, indicating those substances were acidic. When Phenol Red was added to baking soda, it turned red, indicating it was basic.

Similarly, when Alizarine Yellow was added to vinegar, it turned yellow. However, since the lower range of Alizarine is a pH of 10.2, this only indicated vinegar was not strongly basic. However, since vinegar also turned red when Methyl Red was added, it must therefore be acidic.

2.1.6 Identify and describe some everyday uses of indicators including the testing of soil acidity/basicity

Focusing on the soil point below should be enough to answer any dotpoints. However, keeping fish tanks and swimming pools in your head may prove useful just to pull it out if necessary. Another possibility which may or may not be considered everyday use is managing certain chemical wastes from labs, workshops (e.g. disposing of acids in the school labs, or disposing of photographic waste from studios).

Indicators find a wide variety of uses everyday. Below is a list of some uses:

- Soil- Different plants, flowers, fruits and vegetables require different levels of pH in order to thrive. In addition environmentalists often monitor and record soil pH in different areas for various purposes. For example, hydrangeas grow purple flowers in acidic conditions and pink flowers in basic conditions. Monitoring pH is thus important if a specific colour is desired.
- Fish tanks- Different fish require different levels of pH (Although this range is quite narrow) in order to thrive.
- Swimming pools- Swimming pools must be maintained at a pH a little higher than 7 for safety reasons.

2.2 Acidic Oxides and the Atmosphere

2.2.1 Identify oxides of non-metals which act as acids and describe the conditions under which they act as acids

Most oxides of non-metals, excluding the neutral oxides of carbon monoxide, nitrogen monoxide, and nitrous oxide (CO, NO and N₂O respectively), are acidic when in solution. Examples of such acidic oxides include carbon dioxide (CO₂) and nitrogen dioxide (NO₂).

Remember- Non-metal oxides are acidic in solution.

2.2.2 Analyse the position of these non-metals in the Periodic Table and outline the relationship between position of elements in the Periodic Table and acidity/basicity of oxides

The easiest way to remember the information below is to think of substances you know that are basic or acidic. We know that nitrogen dioxide is a component of acid rain, so must be acidic. Nitrogen is a non-metal, so generally non-metal oxides are acidic. We also know that lime, calcium oxide, is basic. Calcium is a metal, so generally metal oxides are basic. This will save you learning by rote.

Metal oxides on the left side of the Periodic Table are generally basic. Non-metal oxides on the right side of the Periodic Table are generally acidic. Oxides of the five lower metalloids, those elements bordering metals and non-metals, can be amphoteric depending on their oxidation states. An amphoteric substance shows both acidic and basic properties.

Remember- The left side of the periodic table generally consists of acidic oxides, while the right side is generally made of basic oxides. Thus metals form acidic oxides, and non-metals form basic oxides.

2.2.3 Define Le Chatelier's principle

The symbol for equilibrium is ⇌ and simply means that, in a closed system, the rate of the forwards reaction is equal to the backwards reaction. This simple means that the reactants are converting to products at the same rate that the products are converting back into the reactants. Whilst there appears to be no change on a macroscopic level, the system is continually changing on a microscopic level. This process, known as dynamic equilibrium, results in the concentration of the substances in the system remaining constant.

According to Le Chatelier's principle, if a system at equilibrium is disturbed, then the system will adjust itself in order to minimise the disturbance. However, note that the effects of the disturbance are never fully removed. They are only minimised, or lessened to a degree.

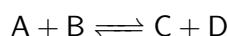
2.2.4 Identify factors which can affect the equilibrium in a reversible reaction

Le Chatelier's principle is one which plays a crucial role in the HSC Chemistry course. Thus, a sound understanding of it is important, and it may appear again in this subject depending upon what Option you do. For this reason, a treatment sounder than required for this dotpoint will be provided.

Several factors can affect the equilibrium in a reversible reaction. These disturbances to the system can be in the form of changes in concentration, pressure, volume, or temperature.

Concentration

Imagine a system in equilibrium of four compounds, A, B, C, and D.



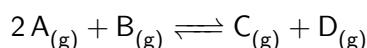
The simplest way of visualising changes in concentration is simply seeing Le Chatelier's principle as working to minimise any changes made to the equilibrium. As more of A or B is added, then the system will try to minimise the change by converting more A and B into C and D. As such, the equilibrium shifts to the right.

Conversely, if more of C or D is added, increasing the concentration of the products, then the system will convert more C and D into A and B, shifting the equilibrium to the left.

Note that a system can only minimise a disturbance. It cannot completely undo it.

Pressure

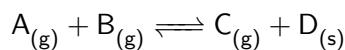
Imagine a system in equilibrium of four compounds, A, B, C, and D. Unlike the example used to illustrate changes in concentration, the four compounds in this example are gases, and the number of moles of A is two rather than one.



Determining the affect of changes in the pressure of a system is simply an exercise in counting moles of gases. In the equilibrium above, there are three moles of gas on the left side, and 2 moles of gas on the right. Any increase in pressure will result in the system trying to relieve the pressure by 'leveling' the moles of gas within the system. As such, in the above system, an increase in pressure will lead to a shift in the equilibrium to the right. This occurs simply because the system is essentially counteracting the fact that three moles of gas are becoming two moles of gas.

Conversely, a decrease in pressure will shift the above equilibrium to the left in an attempt to increase pressure once again.

Changes in pressure affect only gases. Increasing the pressure in the following system will lead to equilibrium shifting to the right, as there are two moles of gas on the left side and only one on the right.



Volume

Any change in volume in a gaseous equilibrium is simply a change in pressure. As such, treat increases in volume as decreases in pressure, as there are more moles of gas in the fixed space, and treat decreases in volume as increases in pressure.

Temperature

The effect of Le Chatelier's principle with changes in temperature can often be confusing. However, simply thinking of heat as either a product or reactant greatly simplifies any problems, as shown in the equilibrium below, where the reaction is endothermic (Absorbs heat in order for the reaction to occur) rather than exothermic (Releases heat).



In the above endothermic equilibrium, an increase in temperature will result in the system working to reduce the temperature by shifting the equilibrium to the right, converting A and B into C and D in order to reduce temperature.

Conversely, a decrease in temperature will shift the equilibrium to the left, converting C and D into A and B in order to produce more heat.

In the case of an exothermic reaction, the equation will be of the form



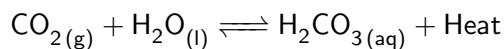
As shown above, treating heat energy as an actual item in the equilibrium is a much simpler method of thinking of a problem. Simply determine whether a reaction is exothermic forwards, i.e. the heat is placed on the right, or endothermic forwards, i.e. the heat is placed on the left.

Remember- Changes in concentration, pressure, volume and temperature will all disturb a system in equilibrium.

2.2.5 Describe the solubility of carbon dioxide in water under various conditions as an equilibrium process and explain in terms of Le Chatelier's principle

This dotpoint is really just an application of Le Chatelier's principle using the solubility of carbon dioxide as an example. As such, set yourself in the habit for such questions by starting with the equation, and then working through changes in concentration, pressure, volume and temperature.

Note that when you have soft drink in a glass or open bottle, you can see bubbles rising in it. This is because the carbon dioxide gas is constantly escaping, thereby constantly favouring the backwards reaction in an attempt to minimise the disturbance to the system. In comparison, a closed bottle of soft drink has no bubbles unless you shake it, because it is in equilibrium.



Using the above equilibrium as a practical example of Le Chatelier's principle:

- An increase in the concentration of $\text{CO}_{2(\text{g})}$ will shift the equilibrium to the right, converting carbon dioxide and water into carbonic acid in order to reduce the concentration of carbon dioxide.
- An increase in pressure will shift the equilibrium to the right, converting carbon dioxide and water into carbonic acid in order to reduce the pressure.
- An increase in the volume of $\text{CO}_{2(\text{g})}$ will shift the equilibrium to the right, converting carbon dioxide and water into carbonic acid in order to reduce the volume of carbon dioxide. Thus the system will attempt to counteract this change by favouring the backwards reaction.
- An increase in temperature will shift the equilibrium to the left, converting carbonic acid into carbon dioxide and water in order to reduce the temperature.
- An increase in temperature will shift the equilibrium to the left, converting carbonic acid into carbon dioxide and water in order to reduce the temperature.

Remember- Le Chatelier's principle will ensure that equilibrium is reached once again. However, this new point of equilibrium will not be same as the original point of equilibrium, as the impact was only minimised, not completely reversed. This is the reason why opened soft drinks will go 'flat' irreversibly.

2.2.6 Calculate volumes of gases given masses of some substances in reactions, and calculate masses of substances given gaseous volumes, in reactions involving gases at 0°C and 100kPa or 25°C and 100kPa

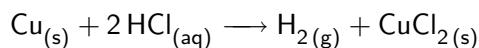
These calculations are made simpler by the fact the HSC data sheet will provide you with the relevant values. The only thing remaining for you to do is to calculate the number of moles and multiply this by the relevant constant. As a note, be wary of earlier papers which used different constants.

According to the hypothesis developed by Avogadro, equal volumes of gases contain equal moles when at the same temperature and pressure.

At 0°C and 100kPa, 1 mole of gas is equal to 22.71L. At 25°C and 100kPa, 1 mole of gas is equal to 24.79L.

Example: Excess hydrochloric acid is added to 5g Cu. What is the volume of gas emitted at 0°C and 100kPa?

Step 1: Write out the equation. In this case:



Step 2: Find the required moles. 5g of copper = $\frac{5}{63.55}$ moles = Number of moles of hydrogen gas produced

Step 3: Use the constants provided. At 0°C and 100kPa, 1 mole of gas is equal to 22.71L. $\frac{5}{63.55} \times 22.71 = 1.79\text{L}$ of H₂ gas emitted.

In making these calculations, always remember that $n = \frac{g}{M}$, (where n is equal to the number of moles, g is equal to weight, and M is equal to the molecular weight of respective element). This equation is crucial to virtually all stoichiometric calculations.

2.2.7 Identify data, plan and perform a first-hand investigation to decarbonate soft drink and gather data to measure the mass changes involved and calculate the volume of gas released at 25°C and 100kPa

If you are unsure how the volume of gas is calculated, review the method outlined in dotpoint 2.2.6 on page 50. Also don't forget that the volume of gas released is given in litres, not millilitres.

Procedure:

1. Record the weight of a bottle of soft drink.
2. Shake the bottle slightly, opening it carefully to release the gas without spilling the contents.
Repeat this step until shaking ceases to have a noticeable effect.
3. Remove the cap, and immerse the bottle in a warm bath for several minutes.
4. Remove the bottle from the bath and leave it overnight in a place away from sources of evaporation.
5. Replace the cap and reweigh the bottle and record its weight.
6. Calculate the volume of carbon dioxide gas released.

Expected Results:

This experiment is simply a practical application of the dotpoint 2.2.6 on page 50. The formula $n = \frac{g}{M}$ will thus be used. Divide the change in weight (in grams) by the molecular mass of carbon dioxide ($M = 44.01$), and multiply by the constant (At 25°C and 100kPa, the relevant constant is 24.79)

For example, if the change in mass was 10g, the number of moles was thus 0.23, and the volume of carbon dioxide gas released must have been 5.63L.

You must be able to work backwards with this formula as well as forwards. Should you be given the volume of gas released, you should know the change in gas.

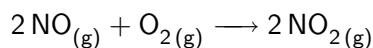
2.2.8 Identify natural and industrial sources of sulfur dioxide and oxides of nitrogen

You will notice this dotpoint turning towards acid rain by first introducing the oxides from which it is formed. Be sure to include all the oxides of nitrogen. These include: Nitrogen monoxide (Also known as nitric oxide), nitrogen dioxide, and nitrous oxide. If you have trouble remembering these, focus on the 'monoxide' and 'dioxide', as these make it clear there are one and two oxygen atoms in the respective molecules.

Natural Sources

Two-thirds of all sulfur dioxide (SO_2) are produced naturally by geothermal hot springs and volcanoes.

Nitrogen monoxide (NO) is produced by lightning, as the high localised temperatures created by lightning are sufficient to convert oxygen and nitrogen in the atmosphere to form nitrogen monoxide. Nitrogen monoxide then reacts slowly with oxygen to form nitrogen dioxide (NO_2):



Nitrous oxide (N_2O) is created naturally by certain bacteria in nitrogen-rich soils.

Industrial Sources

Industrial sources of sulfur dioxide include the combustion of fossil fuels and extraction and refinement of metals from sulfide ores, where SO_2 is often released during the smelting of the ores in order to remove sulfur impurities from the metal. This is because there are traces of sulfur within the ores which are released upon extraction and smelting.

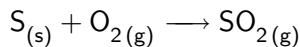
Combustion within power stations and automobiles is an industrial source of nitrogen monoxide and nitrogen dioxide, referred to as NO_x . Just as with lightning, the high temperatures allow the conversion of oxygen and nitrogen into nitrogen monoxide, which combines with oxygen to form nitrogen dioxide.

Remember- Sulfur dioxide occurs naturally due to volcanoes and industrially during the smelting of sulfur-contaminated metal ores. Nitrogen monoxide is formed from atmospheric reactions due to lightning, and nitrogen dioxide forms as nitrogen monoxide in turn reacts with the atmosphere. They are also formed industrially through combustion in automobiles. Nitrous oxide forms from bacteria.

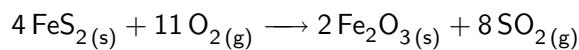
2.2.9 Describe, using equations, examples of chemical reactions which release sulfur dioxide and chemical reactions which release oxides of nitrogen

This dotpoint naturally follows from the previous one, providing the equations for the processes described.

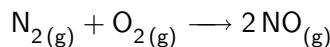
When various fossil fuels are combusted, sulfur within the compound can combine with oxygen to produce sulfur dioxide.



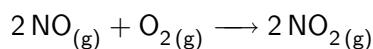
The smelting of iron pyrite produces sulfur dioxide.



High localised temperatures produced in conditions such as lightning or combustion chambers result in the reaction of nitrogen and oxygen to form nitrogen monoxide.



Nitrogen monoxide slowly combines with oxygen to form nitrogen dioxide.



2.2.10 Explain the formation and effects of acid rain

Approach a question like this by splitting it up into formation (atmospheric pollutants, solubility of the acidic oxides) and the effects (environmental and human). If you can mentally organise it this way, you should not only be able to structure a response, but recall the information quite easily as well.

As described before, both sulfur dioxide and nitrogen dioxide are soluble. When the concentrations of the emissions of these gases reach high levels, then rain can become fairly acidic, depending on the amount of gas dissolved. When hydrogen ion concentrations reach around 10^{-5} mol/L, i.e. pH of 5, the rain is described as acid rain.

Acid rain has had a large impact on the environment, increasing the acidity in some lakes to the point where marine life can no longer inhabit the waters for extended periods of time. In addition, forests on an international level have been ravaged by the effects of acid rain, as the surface of leaves upon which the leaves are dependant upon drawing in the water necessary for photosynthesis are destroyed.

The effects of acid rain are not limited to the environment, as man-made structures have been similarly affected. In particular, limestone and marble statues and buildings have had considerable damage dealt to them, as acid corrodes carbonates with considerable ease.

Remember- The reason acid rain is formed is because acidic oxides such as sulfur dioxide and nitrogen dioxide are soluble, reducing the pH of rain considerably at large concentrations. Acid rain impacts the environment and man-made structures alike.

2.2.11 Assess the evidence which indicates increases in atmospheric concentration of oxides of sulfur and nitrogen

Don't be afraid to mention the fact that we have relatively little data to work with in calculating any changes in atmospheric concentration. Most information has only been gathered within the last few decades, with any further information being relatively inaccurate, only illustrating basic trends from sources such as trapped gases in the ice caps.

Being soluble in nature, SO_2 and NO_2 levels have not increased dramatically on a global scale. This is because the water cycle effectively 'cleans' the atmosphere of these oxides on a regular basis. Despite this, the average annual concentration of these two gases is still many times over the concentration of clean air, with numerous days appearing each year where emissions completely go past 'safe levels'.

In contrast to the relatively stable levels of SO_2 and NO_2 , N_2O emissions in Australia have been reported to rise 130% since the 1990s. This has largely been attributed to the use of fertilisers.

Despite such figures, any measurements obtained lack any real figures to compare with, as accurate methods of measurement have only been developed within the last few decades.

Remember- Nitrous oxide concentrations have increased significantly in recent decades, whereas sulfur dioxide and nitrogen dioxide levels have remained relatively constant.

2.2.12 Analyse information from secondary sources to summarise the industrial origins of sulfur dioxide and oxides of nitrogen and evaluate reasons for concern about their release into the environment

As the industrial origins of sulfur dioxide and the various oxides of nitrogen have been covered in prior dotpoints, only the reasons for concern about their release will be covered in this dotpoint.

There exists concern about the release of sulfur dioxide due to health and environmental reasons. Sulfur dioxide aggravates existing lung conditions, and can trigger asthma attacks and bronchitis. In addition, the effects of acid rain previously mentioned are a real problem to marine life, forests, and man-made structures made of marble and limestone.

Similarly, nitrogen dioxide can be detrimental to asthmatics, and can irritate the airways. Even in moderate levels, long-term exposure can increase the chances of respiratory illness, and sensitise people to allergies. Nitrogen dioxide is an even more problematic source of acid rain, as the nitric acid formed is a stronger acid than sulfuric acid.

Photochemical smog produced by nitrogen oxides, sunlight and volatile organic compounds is another large problem, as it leaves ozone- a powerful lung irritant and dangerous in even small concentrations- at a level proximate to humans.

As for nitrous oxide, it has been estimated that N_2O has a global warming potential 300 times higher than carbon dioxide. When combined with the alarming growth reported before, of 130% since the 1990s, this has become a problem of great import.

Thus, the evidence is both plentiful and persuasive, making it abundantly clear that concerns regarding the release of sulfur dioxide and the oxides of nitrogen are in fact well-founded.

Remember- Sulfur dioxide is formed during the smelting of sulfur-contaminated metal ores, and nitrogen monoxide and nitrogen dioxide are formed industrially through combustion in automobiles. Environmental concerns include attacks on the respiratory system (lungs and airways) for people, acid rain, and global warming.

2.3 Acids and pH

2.3.1 Solve problems and perform a first-hand investigation to use pH meters/probes and indicators to distinguish between acidic, basic and neutral chemicals

This experiment is relatively simple. Carefully use a pH meter/probe in a solution in order to determine if it is acidic, basic, or neutral. Don't forget the pH ranges for each of these classifications, which are stated again below for your convenience.

An acid has a pH less than 7, whereas a base has a pH greater than 7. Substances with a pH of 7 are neutral substances.

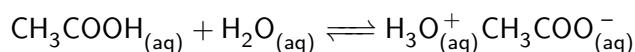
2.3.2 Define acids as proton donors and describe the ionisation of acids in water

Although this dotpoint focuses on acids as proton donors, also keep in mind that this makes a base a proton acceptor. Once you understand that a hydronium ion is equivalent to a proton, you will find acid-base reactions quite simple.

The hydrogen atom consists of one proton and one electron. Thus, when this atom ionises and loses an electron to form H^+ , only a proton remains.

When an acid molecule reacts to release the hydrogen ion H^+ , it is thus donating a proton. In the presence of hydrogen ions, water will convert to a cation known as the hydronium ion (H_3O^+).

An example of the ionisation of an acid in water using acetic acid- an acid known for its presence in vinegar - is shown below:



In the above equation, the donated proton, or hydrogen ion, ionised the water molecule.

It is also useful to note that a monoprotic acid is one that releases one proton. Similarly, a diprotic or triprotic acid releases two or three protons respectively. For example, sulfuric acid is a diprotic acid, and citric acid is a triprotic acid.

Remember- If a substance can 'disassociate' into the hydronium ion and an anion, then we will often recognise it as an acid. This dotpoint simply seeks to illustrate this point, noting that the hydrogen ion is simply a proton. Thus an acid is a proton donor.

2.3.3 Identify acids including acetic (ethanoic), citric (2-hydroxypropane-1,2,3-tricarboxylic), hydrochloric and sulfuric acid

Historical names will not appear often, but systematic names will appear from time to time in examinations. It may also be useful to remember one use for each of these acids, but I would not recommend anything further.

Acetic acid, also known as ethanoic acid, occurs naturally, but is manufactured on an industrial level for the production of items such as soft drink bottles, glue, and vinegar, and is also used as a food additive. Formic acid is methanoic acid, and can be found in ant stings.

Citric acid, also referred to by its systematic name of 2-hydroxypropane-1,2,3-tricarboxylic acid, occurs naturally in citrus fruits. Citric acid is used predominantly as a food additive.

Hydrochloric was historically known as muriatic acid, and finds a wide-scale use industrially due to its nature as a strong acid. One such example is the production of vinyl chloride for its polymer PVC.

Sulfuric acid was historically known as oil of vitriol, and also finds a wide-scale use industrially. A large proportion of sulfuric acid is used in the fertiliser industry, and sulfuric acid is also used as a catalyst in many industrial reactions.

Remember- Acetic acid is ethanoic acid, and formic acid is methanoic acid.

2.3.4 Identify data, gather and process information from secondary sources to identify examples of naturally occurring acids and bases and their chemical composition

Acids:

- Acetic acid ($\text{CH}_3\text{CO}_2\text{H}$)- Acetic acid is produced naturally through bacterial fermentation, and is used in the production of vinegar and as a solvent.
- Citric acid ($\text{C}_6\text{H}_8\text{O}_7$)- Citric acid is produced naturally in citrus fruits, and is used in food both as a preservative, and for flavour.
- Formic acid (HCOOH)- Formic acid can be found in the stings of ants and bees, and is used for a variety of purposes such as esterification and as a preservative.
- Lactic acid ($\text{C}_3\text{H}_6\text{O}_3$)- Lactic acid is produced within the body by muscles during exercise. It can also be found within certain milk products such as yoghurt and various cheeses.

Bases:

- Ammonia (NH_3)- Certain organisms can produce ammonia from atmospheric nitrogen. Ammonia base is used within household cleaners, and in the production of fertilisers such as ammonium sulfate.
- Potassium hydroxide (KOH)- The leaching (The passing through of water resulting in the dissolving of soluble substances) of the certain wood ashes was known historically to produce potash, or potassium hydroxide. Potash can be used to produce soap as well as fertiliser.
- Sodium bicarbonate (NaHCO_3)- Produced naturally in a mineral known as nahcolite, sodium bicarbonate is used for various pest control purposes and antacids. Sodium bicarbonate is also used in the kitchen to make dough rise.

Take care not to confuse ammonia (NH_3) with ammonium (NH_4^+).

Remember- Acetic acid and formic acid are also known as ethanoic acid and methanoic acid respectively.

2.3.5 Describe the use of the pH scale in comparing acids and bases

The pH scale is used to compare the acidity and basicity of substances, making use of a logarithmic scale so as to take into account the great variances of concentration.

For example, a hundredfold change will only result in a change in pH of 2.

This scale measures how acidic or basic a substance is, by assigning it a number to indicate its nature. An acid has a pH less than 7, whereas a base has a pH greater than 7. Substances with a pH of 7 are neutral substances.

This dotpoint will be expanded upon further down in dotpoint 2.3.10 on page 60.

2.3.6 Describe acids and their solutions with the appropriate use of the terms strong, weak, concentrated and dilute

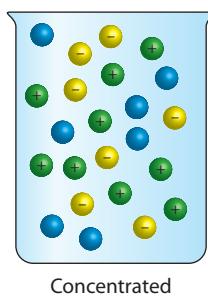
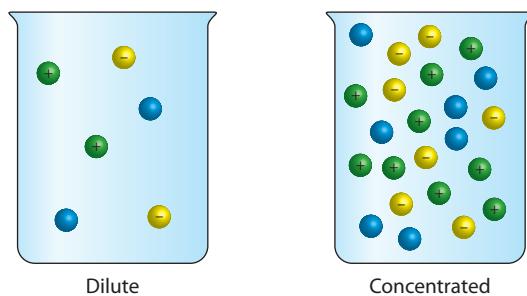
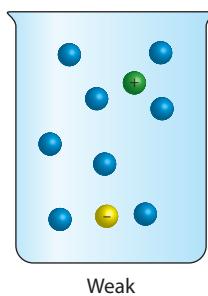
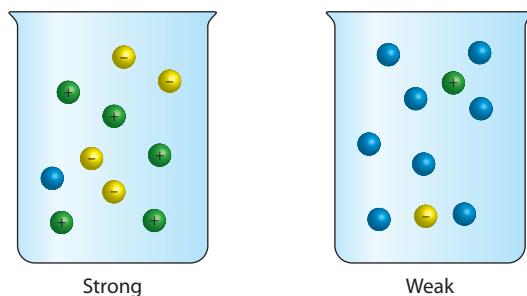
The use of these terms poses some confusion for most students, so take the time to read and understand these definitions.

The terms 'strong' and 'weak' are used to describe the degree of ionisation of an acid. If an acid ionises to completion within solution, then the acid is what is termed a strong acid. However, if neutral acid molecules still remain, meaning only some of the acid molecules have ionised, then the acid is a weak acid. Where an acid reaches equilibrium will determine how we classify it.

Furthermore, an acid can be termed 'concentrated' or 'dilute', referring to the concentration of the solution (mol/L). If the total concentration of the solute is high, then the acid is concentrated. If the total concentration is low, then it is called a weak acid.

Remember- A strong acid is not necessarily concentrated, just as a weak acid is not necessarily dilute. These categories are independent of one another, and a strong acid can be dilute just as a weak acid can be concentrated.

+ positive ionised molecules
- negative ionised molecules
● neutral un-ionised molecules



2.3.7 Plan and perform a first-hand investigation to measure the pH of identical concentrations of strong and weak acids

The approach to this dotpoint will greatly vary depending on your school, but the use of pH probes presents the simplest method. If not, a variety of indicators (Including universal indicator) can be used to narrow down the possible range of pH.

If you do not conduct this experiment, then simply note that the result is essentially 'Equal concentrations of acids do not necessarily yield equal pH measurements.' This is a practical application of the differing degrees of ionisation between acids. The stronger acid will have a lower pH, reflecting its higher degree of ionisation.

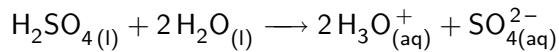
2.3.8 Gather and process information from secondary sources to write ionic equations to represent the ionisation of acids

You will find these equations easy to do once you get used to them. The acid will always be ionising in water, hence these will be the reactants. The products will then be an anion, as well as a hydrogen ion (Or proton as an acid is a proton donor).

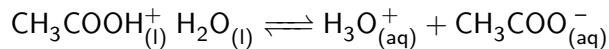
The ionisation of hydrochloric acid (complete):



The ionisation of sulfuric acid (complete):



The ionisation of acetic/ethanoic acid (partial):



2.3.9 Use available evidence to model the molecular nature of acids and simulate the ionisation of strong and weak acids

The diagram introduced in dotpoint 2.3.6 on page 58 may aid with your understanding of the distinctions between acids which are strong or weak, and concentrated or dilute.

2.3.10 Identify pH as $-\log_{10}[\text{H}^+]$ and explain that a change in pH of 1 means a ten-fold change in $[\text{H}^+]$

Simply feed this function through your calculator to obtain the pH, when given the hydrogen ion concentration. Don't forget this function uses a base of 10, and appears as 'log' on your calculator, as opposed to the natural logarithm, 'ln', which uses a base of the constant e.

The relative strength of acids and bases can be measured through the concentration of hydrogen ions present within the molecules. As this concentration is subject to large variances, a logarithmic function is used.

pH = $-\log_{10}[\text{H}^+]$, where the square brackets mean 'concentration'. As such, this can simply be read: 'pH is the negative logarithm of the concentration of hydrogen ions.'

A logarithm is used to effectively provide smaller numbers to work with, as a change in 1 pH represents a tenfold change in hydrogen ion concentration. Similarly, a change in 2 pH represents a hundredfold change.

Thus, an acid with a pH of 2 has a hydrogen ion concentration 10 times more than an acid with a pH of 3.

You also need to be able to work backwards using the logarithmic formula, i.e. work out the hydrogen ion concentration when given the pH. A simple manipulation of the formula provides the result $[\text{H}^+] = 10^{-\text{pH}}$.

Remember- pH = $-\log_{10}[\text{H}^+]$

A point which may be of use which does appear occasionally is the fact that: pH + pOH = 14. Calculations often appear which require this formula, and if you take the time to think about it, it makes sense that adding a measure of hydrogen ion concentration with a measure of hydroxide ion concentration gives you the upper limit of the pH scale. Practice converting from pH to $[\text{H}^+]$ and calculation questions within this area should prove no difficulty.

2.3.11 Compare the relative strengths of equal concentrations of citric, acetic and hydrochloric acids and explain in terms of the degree of ionisation of their molecules

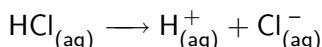
This dotpoint is linked quite closely to the terms 'strong' and 'weak' acid, so re-read the definition for these terms if they have slipped your mind.

Although three different acids may have equal concentrations, they may have different relative strengths due to the degree of ionisation of their molecules.

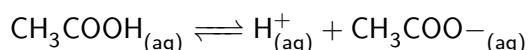
Hydrochloric acid is a strong acid, meaning that virtually 100% of the acid molecules ionise when in solution. In contrast, citric acid and acetic acid only ionise around 1% of their acid molecules, making these weak acids. However, note that citric acid can potentially have each acid molecule dissociate into three protons, as it is a triprotic acid. This makes citric acid the stronger of the two.

This relationship can be represented through the use of the following equations.

Full-ionisation:



Partial-ionisation:



The use of the symbol \longrightarrow demonstrates that the ionisation runs until completion, whereas \rightleftharpoons indicates that the reaction is in a state of equilibrium.

Remember- A weak acid is one where neutral acid molecules are prevalent within the acid, meaning that not all the acid molecules have ionised.

2.3.12 Describe the difference between a strong and a weak acid in terms of an equilibrium between the intact molecule and its ions

The main purpose behind this dotpoint is to point out that the degree of ionisation of an acid is what differentiates a strong acid from a weak acid. Although I would expect that a student have a sound understanding of this difference, the actual definition and calculation of an acid's degree of ionisation is unlikely to appear in an exam.

The degree of ionisation of an acid is the percentage of molecules which are ionised in solution. This is simply the concentration of hydrogen ions divided by the concentration of the acid, as a percentage. For example, a 0.2 mol/L solution of an acid with a pH of 2 has a degree of ionisation of $\frac{10^{-2}}{0.2} = 5\%$

$$\text{Degree of Ionisation} = \frac{[\text{H}^+]}{\text{Concentration}}$$

Remember- A strong acid ionises to completion. A weak acid enters a state of equilibrium between the intact molecule and its ions.

2.3.13 Process information from secondary sources to calculate pH of strong acids given appropriate hydrogen ion concentrations

You will be expected to be able to calculate pH using information about hydrogen ion concentration, as well as calculating hydrogen ion concentration from pH. Simply use the given formula and your calculator to do so. Don't forget to use the 'log' button on your calculator instead of 'ln'.

$$\text{pH} = -\log_{10}[H^+]$$

or

$$10^{-\text{pH}} = [H^+]$$

Using the above equation and its variation, you will be able to switch between hydrogen ion concentration and pH relatively easily.

2.3.14 Gather and process information from secondary sources to explain the use of acids as food additives

Acids find a myriad of uses as food additives. Below are three which I believe would be useful to learn (Also be prepared to give an example of each).

Acids have a wide range of application in the culinary industry, as they can be used as food additives for a variety of purposes. Acids are often used as:

- Preservative- Acids such as acetic acid prevent the growth of microorganisms such as salmonella within food, preserving the food so that it has a longer shelf-life. Phosphoric acid is similarly used in soft drinks to kill bacteria.
- Antioxidant- Some acids such as ascorbic acid are used as antioxidants, preventing spoilage of food due to oxygen.
- Flavouring- Acids are also used to add taste. Citric acid is commonly used to add tartness.

2.4 Acid/base Theories

2.4.1 Gather and process information from secondary sources to trace developments in understanding and describing acid/base reactions

Although covered in detail in later dotpoints, this question appears often enough to warrant a quick overview before launching into the specifics.

- Lavoisier defined acids by taking into account their composition. Defining acids as substances containing oxygen, this clearly excluded acids such as hydrochloric acid.
- Davy developed our understanding of acids by defining them by their reactions. Stating that acids were simply substances that contained hydrogen that could be replaced by a metal. Distinctly acidic metals such as silver are excluded by this definition.
- Arrhenius defined acids and bases through their ability to ionise, contributing greatly to the concept of weak and strong acids and bases. This also served to explain the relative conductivities of acids and bases. However, this definition did not include carbonates which do not dissolve to produce hydroxide ions, yet are distinctly basic.
- Brönsted and Lowry defined acids and bases through their properties relative to their solvents. Explaining much in terms of acid/base equilibria, particularly the pH of salts, this definition still fails to acknowledge the acidic characteristics of metals such as silver.

2.4.2 Outline the historical development of ideas about acids including those of: Lavoisier, Davy, and Arrhenius

It is necessary to learn exactly how each scientist contributed to the development of the acid-base definitions, including what exactly they believed defined acids, as well as any flaws in their definitions.

In 1779, Antoine Lavoisier asserted that acids could be defined as substances which contained oxygen. However, one limitation of this definition was that it clearly excluded acids such as hydrochloric acid, HCl, yet included bases such as calcium oxide, CaO. However, this definition aimed at defining acids through the composition of substances.

In 1815, Humphrey Davy made the observation that acids contained hydrogen which would be replaced with a metal. This definition aided in the development of acid definitions by relating their nature to their reactions with other substances, yet failed to take into account acidic metals such as silver.

In 1884, Svante Arrhenius proposed that acids were substances which ionised to produce hydrogen ions (H^+), and that bases were substances which ionised to produce hydroxide ions (OH^-). Although this definition did much by way of explaining the relative strengths of acids and bases, and their respective conductivities, it excluded carbonates which do not dissolve to produce hydroxide ions, yet are still clearly basic.

Remember- Each individual extended the definition of an acid by relating their definitions to the physical and chemical properties of known acids based upon their observations, and the work of previous scientists.

2.4.3 Outline the Brönsted-Lowry theory of acids and bases

The Brönsted-Lowry theory of acids and bases is not the most recent definition, but is certainly the most common definition you will encounter in the HSC course. It hinges on the nature of an acid as a proton donor.

The Brönsted-Lowry theory of acids and bases was developed independently by two scientists, defining acids as proton donors and bases as proton acceptors. This definition is important as it identifies the importance of the substance's properties relative to the solvent, and did much to explain variances in the pH of salts.

Despite this, the Brönsted-Lowry acid-base definition does not explain the nature of acidic metals such as silver.

Although not strictly a part of this dotpoint, it is important that you understand what is meant by the statement 'An acid is a proton donor.' As such, the concept will be explained again in the following paragraph.

The hydrogen atom, H, has one proton and one electron. When the hydrogen atom is ionised, then it loses an electron, becoming positively charged, H^+ . However, as we already stated that the hydrogen atom has one proton and one electron, then when ionised, only a proton is left. As such, when an acid ionises in solution, it donates a 'proton'.

See the dotpoint 2.4.4 below for an expansion upon the concept of proton donors and proton acceptors.

Remember- The Brönsted-Lowry theory of acids and bases states that acids are proton donors and bases are proton acceptors, where a hydrogen ion is a proton.

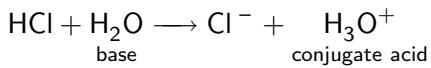
2.4.4 Describe the relationship between an acid and its conjugate base and a base and its conjugate acid

If an acid is a proton donor, then after it donates the proton, what is left can essentially accept a proton, thus defining a base as a proton acceptor. As such, the base that is formed is the acid's conjugate base. A similar relationship exists between a base and its conjugate acid.

If an acid, for example hydrochloric acid, donates a proton, its conjugate base is formed.



Conversely, if a base, for example water, accepts a proton, its conjugate acid is formed



The conjugate acid of a strong base is a weak acid, and the conjugate acid of a weak base is a strong acid. Conversely, the conjugate base of a strong acid is a weak base, and the conjugate base of a weak acid is a strong base.

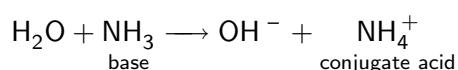
Remember- The base formed when an acid donates a proton is that acid's conjugate base, and turns back into the acid when it accepts a proton. In reverse, a base has a corresponding conjugate acid formed when it accepts a proton.

2.4.5 Identify conjugate acid/base pairs

You can easily identify which substances are paired by identifying the structure of each substance. For example, HCl and Cl⁻ are one pair, whereas H₂O and H₃O⁺ are another pair. It is then simply down to identifying which is a conjugate acid, and which is a conjugate base. Alternatively, you can identify the conjugate acid as the product with the positive charge, and the conjugate base as the product with the negative charge.

With reference to the previous dotpoint 2.4.4 [Describe the relationship between an acid and its conjugate base and a base and its conjugate acid], the conjugate acid/base pairs are identified as HCl/Cl^- and $\text{H}_2\text{O}/\text{H}_3\text{O}^+$ respectively.

Another example is as follows:



Note that in a neutralisation reaction, there will be two conjugate pairs. Don't forget to state both if required.

2.4.6 Choose equipment and perform a first-hand investigation to identify the pH of a range of salt solutions

This experiment simply seeks to demonstrate the effect on the pH of salts when strong and weak acids and bases are used to neutralise one another. As such, prepare the four combinations below, perhaps using the examples provided in dotpoint 2.4.7, and review the expected results below. A pH probe can be used for quick results.

Expected Results:

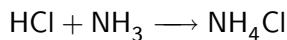
Acid Concentration	Base Concentration	Nature of Salt
Strong	Weak	Weak Acid
Weak	Strong	Weak Base
Strong	Strong	Relatively Neutral
Weak	Weak	Relatively Neutral

2.4.7 Identify a range of salts which form acidic, basic or neutral solutions and explain their acidic, neutral or basic nature

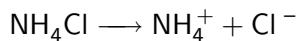
Although the common table salt forms a neutral solution, this is not the case for all salts. Rather, the pH of the salt depends upon the solutions which react to form it. These solutions are in turn an acid and a base. Four basic scenarios are possible, and are listed below.

When a strong acid neutralises a weak base, such as HCl reacting with NH₄OH (Ammonium hydroxide), the cation formed is a weak acid.

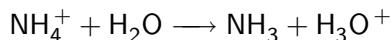
For example:



The salt dissociates in water, forming ammonium cations and chloride anions.



The chloride anion is the weak conjugate base of the strong hydrochloric acid. Thus, it is relatively inert in water. The ammonium cation is the strong conjugate acid of the weak base ammonia. Thus, the ammonium ions react in water, acting as an acid.



The production of hydronium ions in this process is responsible for the acidity of the salt formed.

Remember, conjugate acid-base pairs have inversion in their relative strengths. As such, the conjugate acid of a strong base is a weak acid, and the conjugate base of a weak base is a strong base.

Similarly, when a strong base neutralises a weak acid, such as NaOH reacting with HCOOH (Formic acid), the anion formed is a weak base. The conjugate acid of the strong base is weak, and is thus inert. The weak acid produces a strong conjugate base, which react with water to produce a basic solution.

When a strong acid and a strong base react, the salt formed is relatively neutral. This is because the conjugate acid and conjugate base of a strong base and strong acid respectively are weak, and do not react significantly with water.

When a weak acid and a weak base react, the salt formed is also relatively neutral. Both the conjugate acid and the conjugate base are strong, and react with water to form hydroxide ions and hydronium ions respectively. The net effect of this is the formation of water. As such, the effects are largely cancelled out between the anion and the cation.

Reviewing the results of the experiment in dotpoint 2.4.6:

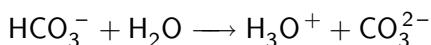
Acid Concentration	Base Concentration	Nature of Salt
Strong	Weak	Weak Acid
Weak	Strong	Weak Base
Strong	Strong	Relatively Neutral
Weak	Weak	Relatively Neutral

2.4.8 Identify amphiprotic substances and construct equations to describe their behaviour in acidic and basic solutions

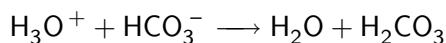
Make sure you don't confuse the term amphiprotic with the term amphoteric, which refers to a substance's ability to act as either an acid or a base. The difference is subtle, but does exist.

A substance that can behave as both a proton donor and a proton acceptor is known as an amphiprotic substance. An example of an amphiprotic substance is the hydrogen carbonate ion, HCO_3^- .

Acting as an acid:

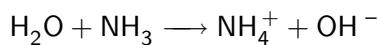


Acting as a base:

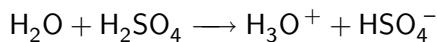


Other examples of amphiprotic substances include water and the hydrogen sulfate ion.

Water acting as an acid:



Water acting as a base:



Remember- An amphiprotic substance is one which can act as either a proton donor or proton acceptor under different conditions.

2.4.9 Identify neutralisation as a proton transfer reaction which is exothermic

This dotpoint is relatively straightforward as we have established an acid as a proton donor and a base as a proton acceptor.

As we have already identified the transferring of a hydrogen ion as the transferring of a proton, it is a given that neutralisation- the reaction between an acid and a base- is a proton transfer reaction. This is because a neutralisation reaction involves a proton donor and a proton acceptor, where a proton is effectively transferred from the acid to the base. This reaction usually produces a heat of neutralisation between strong acids and strong bases of approximately 57 kJ per mole of water formed. Thus this reaction is exothermic (Heat is given off during the reaction).

2.4.10 Describe the correct technique for conducting titrations and preparation of standard solutions

Various methods will exist for conducting a titration and preparing a standard solution. The procedure provided below is but one example, and can be altered in many ways but remain valid. The key for any question is to identify how the method you provide allows for a high degree of experimental reliability and validity.

Titration is an experimental procedure used to determine the concentration of an unknown acid or base. It works by taking the unknown sample, and then neutralising it using a measured quantity of a substance of known concentration. The neutralisation is detected by using an indicator that changes colour at the midpoint of the titration, known as the equivalence point.

As seen in dotpoint 2.4.7 on page 66, when an acid and a base react, they can form a salt that has a pH related to the strengths of the acid and base used. This means that the equivalence point of a titration is not necessarily neutral, with a pH of 7, since the neutralisation reaction may form an acidic or basic salt. Therefore, an indicator must be chosen that changes colour at the expected pH for the equivalence point of the reaction. For example, if a weak base is neutralised by a strong acid, then the solution will be weakly acidic, so the indicator chosen must change colour at a pH of around 3.5, rather than 7.

The premise of titration is the equation $c_1v_1 = c_2v_2$, where c_1 and c_2 are the concentrations of the unknown and known solutions, and v_1 and v_2 are the volumes of the unknown and known solutions. In a titration, everything is known or measured apart from c_1 , for which the equation is solved. This means two things: Firstly, v_1 and v_2 need to be precisely measured, and secondly, c_2 is known with a high degree of accuracy. A standard solution the term given to a highly pure solution that can be used to provide accurate measurements in volumetric analysis techniques such as those used to perform a titration.

Once the reaction reaches its equivalence point and both reactants are completely consumed, the concentration of the unknown solution c_1 can be determined using basic stoichiometric calculations. For example, if 500mL of a 2 mol/L solution of hydrochloric acid is required to completely neutralise a 250mL of sodium hydroxide, then the concentration of sodium hydroxide must be 4 mol/L.

When preparing a standard solution:

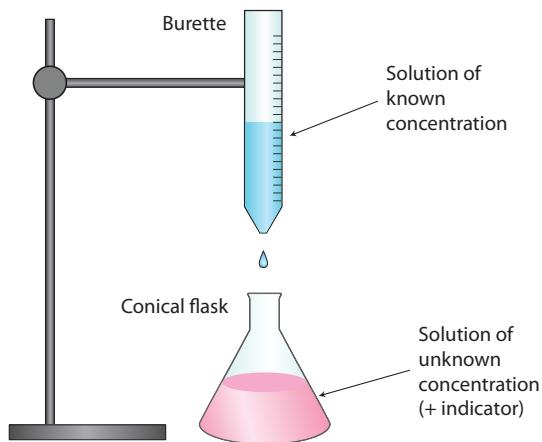
1. Weigh the primary standard- The solid dissolved to form the standard solution. This solid must be as pure and as free from moisture as possible. If necessary, use an oven or desiccator to dry or cool the solid respectively.
2. Dissolve the primary standard in a small amount of distilled water in a beaker.
3. Pour the contents of the beaker into a volumetric flask, rinsing the beaker slightly to ensure all the contents are transferred.
4. Add water to the flask up until the graduation mark. Use a small squirt bottle for the final few drops.

Remember that each step in preparing a standard solution is designed to minimise errors in determining the concentration of the resulting standard solution. If the concentration of the standard solution is incorrect, then the calculated concentration for the unknown solution will also be incorrect.

When conducting titrations:

1. Wash a burette using distilled water, using a small amount of the prepared solution with known concentration to flush out the burette afterwards.
2. Fill the burette with the rest of the solution of known concentration, taking care to ensure the solution is correctly zeroed.
3. Place a solution which is being tested in a flask under the burette.
4. Add two drops of the appropriate indicator to the conical flask.
5. Using one hand to twist the burette to allow its contents into the flask below, use the other hand to constantly swirl the flask.
6. When the solution in the flask changes colour- at its equivalence point- ensure that the burette is no longer dripping, and read the volume of solution discharged by the burette. The first reading should be taken as an approximation, and three further readings should be averaged to determine the result.
7. Use basic stoichiometric calculations to calculate the unknown concentration.

Distilled water is used in this experiment because tap water may contain dissolved salts that can affect the titration.



Remember- $c_1v_1 = c_2v_2$, where c is equal to concentration and v is equal to volume. Watch what units you use!

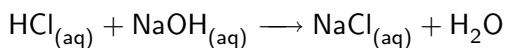
2.4.11 Perform a first-hand investigation and solve problems using titrations and including the preparation of standard solutions, and use available evidence to quantitatively and qualitatively describe the reaction between selected acids and bases

Make use of the methods outlined in the above dotpoint 2.4.10 on page 68 in preparing the standard solution and conducting the titration. Although the process is important, of equal importance are considerations of reliability and validity. Be prepared to answer questions which ask you to improve the experiment, as well as basic risk assessment questions. Although these answers will be common sense, make sure to explain how your recommendation would improve reliability/validity/safety.

You will be required to provide stoichiometric calculations using the formula $c_1v_1 = c_2v_2$, where c is equal to concentration and v is equal to volume. Again, make sure that your two v values are in the same units (mL or L).

For example, imagine you have conducted a titration using- for the sake of simplicity- hydrochloric acid and sodium hydroxide.

1. Write down the equation



2. Identify what variables you know in the equation $c_1v_1 = c_2v_2$. You should be given, or be able to work out, three of the four necessary variables. Let us assume that the concentration of HCl was 1M, and 50mL of the solution was used to react 45mL of NaOH. The concentration of the NaOH is unknown.
3. Work out the number of moles of the substance you have the concentration and volume of using the formula $n = cV$.

The number of moles of HCl was therefore 1×0.05 which equals 0.05 moles.

Keep in mind that V is measured in litres.

4. Work out the ratio of moles between the two substances. In this case, one mole of HCl reacts one mole of NaOH.

If a ratio other than a one-to-one ratio is obtained, then you must take into account the fact that X moles of the acid reacts Y moles of the base, which would clearly influence your later calculations of the unknown variable.

5. Work out the missing variable. At this point you have the number of moles of the substance with an unknown variable, as well as one known variable. Using the $c = n/V$ formula, you can easily work out the last variable. In our example, if $c = n/V$, then $n = \frac{c}{V}$. The number of moles has been determined to be 0.05, and $V = 0.045$. Therefore the concentration must be $0.05 \times 0.045 = 0.00225\text{M}$.

2.4.12 Perform a first-hand investigation to determine the concentration of a domestic acidic substance using computer-based technologies

This experiment simply involves the use of pH probes in solutions. This experiment is unlikely to appear in any questions, but simply remember to calibrate the probe before using it for ideal results.

Procedure:

1. Prepare beakers of a variety of domestic acidic substances such as vinegar, lemon juice and milk.
2. Calibrate a pH probe.
3. Place the probe gently into each beaker in turn, taking care to dip the probe in the buffer in between each test so as to avoid distorted results.
4. Record your results.

2.4.13 Analyse information from secondary sources to assess the use of neutralisation reactions as a safety measure or to minimise damage in accidents or chemical spills

It is highly recommended that not only should you note at least one substance used to neutralise acids/bases, but be able to provide reasons as to why it would be an effective means of minimising danger.

With many acids and bases being highly corrosive, neutralisation reactions form an integral part of safety protocols within scientific and industrial settings. Finding uses ranging from the management of photographic and sewerage wastes, neutralisation reactions offer a quick means of minimising danger if correctly done.

One such example is sodium hydrogen carbonate. Advantages of using NaHCO_3 include:

- It can be stored and transported easily given its stable nature.
- It is amphotropic, meaning that it can be used to neutralise both acid and alkali spills.
- It is relatively cheap, providing an incentive for industrial use.
- It is not so strong that excess usage would be extremely detrimental. If excess NaHCO_3 is used, the result will only be mildly reactive. In contrast, excess HCl or NaOH could serve to only worsen the situation as the result would be strongly acidic or basic.

Remember- The ideal neutralising agent should be safe to store and use, effective against a variety of situations (such as both acid and base spillages), and should be economically viable.

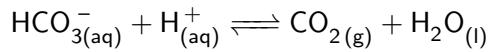
2.4.14 Qualitatively describe the effect of buffers with reference to a specific example in a natural system

Keep in mind Le Chatelier's principle when trying to grasp what a buffer does. This guide will use the bicarbonate ion as an example of a buffer- in particular its use in maintaining the pH of blood.

A buffer is a substance that will work to maintain the pH of a system even if small amounts of acids or bases are added. A buffer is normally a combination of chemicals, commonly consisting of a weak acid and its conjugate base (although it can also be formed from a weak base and its conjugate acid).

An example of a buffer that can be found in a natural system is the bicarbonate ion in human blood. Although this natural system must maintain a pH of 7.4, carbon dioxide frequently makes its way into the bloodstream, potentially reducing pH below 7.0.

In response to this, the bicarbonate ion found in blood plasma forms a state of equilibrium as shown below.



In the above equation, if an acid is added, the increase in hydrogen ions (pH falls) will result in a favouring of the forwards reaction according to Le Chatelier's principle, consuming the additional hydrogen ions (pH rises). Conversely, if a base is added, the hydroxide ions (pH rises) will react with the hydrogen ions, such that the system will work to correct itself by favouring the backwards reaction as stated by Le Chatelier's principle (pH falls).

In both cases, it can be seen that a buffer works to lessen changes in the pH of a system.

Remember- As a buffer is essentially an application of Le Chatelier's principle, it can only minimise the effect on a system rather than completely remove the effects.

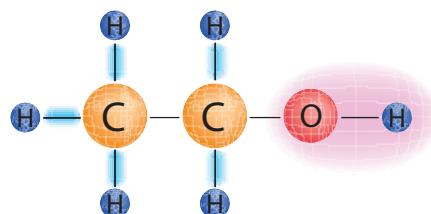
2.5 Esterification

2.5.1 Describe the differences between the alkanol and alkanoic acid functional groups in carbon compounds

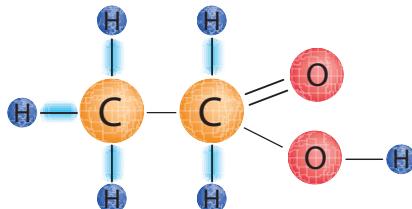
Alkanols such as methanol (CH_3OH) contain the hydroxyl functional group-OH, resulting in high melting and boiling points.

Alkanoic acids such as methanoic acid, also known as formic acid (CH_2O_2), contain the carboxylic acid functional group (-COOH), resulting in even higher melting and boiling points due to the presence of stronger intermolecular forces. This will be explained further in the subsequent dotpoints.

Ethanol



Ethanoic Acid



Remember- Alkanols have a -OH functional group, while alkanoic acids have a -COOH functional group.

2.5.2 Explain the difference in melting point and boiling point caused by straight-chained alkanoic acid and straight-chained primary alkanol structures

Keep in mind that differences in melting and boiling points arise from differences in intermolecular bonding. Dispersion forces arising from polar molecules and hydrogen bonding are examples of intermolecular bonding. Intramolecular forces that bind the atoms that make up the molecule don't affect the melting or boiling point, only the reactivity.

As a quick recap, intermolecular forces are the bonds between molecules, whereas intramolecular forces are the bonds within molecules.

Alkanols have high melting and boiling points due to the polarity and potential for hydrogen bonding present within the molecules.

Similarly, the polarity of an alkanoic acid is one reason why the melting and boiling points of an alkanoic acid are high, a property increased due to the greater molar mass of alkanoic acids relative to alkanols. When combined with the potential for two occurrences of hydrogen bonding to occur on the -COOH chain (and thus much strong intermolecular forces, driving up the melting and boiling point), this means that the melting and boiling points of alkanoic acids are higher than that of alkanols.

Remember- Alkanoic acids have higher melting and boiling points because the -COOH chain increases intermolecular hydrogen bonding.

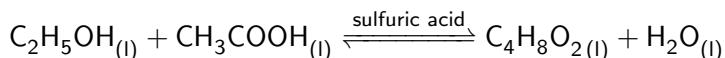
2.5.3 Identify esterification as the reaction between an acid and an alkanol and describe, using equations, examples of esterification

As with most identify dotpoints, this one is fairly straightforward. Simply remember the definition of esterification, and be prepared to use at least one equation to back your definition up. As a note, the formation of ethyl ethanoate, as shown below, is an easy example because it's balanced to begin with.

Esterification is the reaction between an acid and an alkanol, through which an ester is formed. It is a reversible reaction, and the products are the ester and water. Concentrated sulfuric acid is used as a catalyst.



For example, the formation of ethyl ethanoate from ethanol and ethanoic acid is as follows:



Remember- Esterification reacts an alkanol and an alkanoic acid to form an ester and water, in a reversible reaction.

2.5.4 Identify the IUPAC nomenclature for describing the esters produced by reactions of straight-chained alkanoic acids from C1 to C8 and straight-chained primary alkanols from C1 to C8

For naming the esters produced by straight-chained alkanoic acids, the alkyl alkanoate produced follows a very conventional nomenclature. By now you should know the prefix for C1 to C8 (Meth-, eth-, prop-, but-, pent-, hex-, hept-, oct-). All that remains is identifying which part of an 'alkyl alkanoate' is derived from the alkanol, and which part is derived from the alkanoic acid.

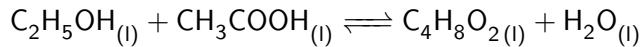
The 'alkyl' portion is named by the alkanol reacted with the alkanoic acid, which names the 'alkanoate' portion. For example, if propanol is reacted with ethanoic acid, the resulting ester is propyl ethanoate (Which has a pear-like smell). Similarly, butanol and butanoic acid react to form butyl butyrate (Pineapple-like scent).

Remember- Alkanol (1) and alkanoic acid (2) react to form alkyl (1) alkanoate (2).

2.5.5 Describe the purpose of using acid in esterification for catalysis

Concentrated sulfuric acid is the most common catalyst used during esterification. This is primarily due to its nature as a strong dehydrating agent, and its subsequent effect upon the reaction. Be sure to explain how the forward reaction is increased through Le Chatelier's principle rather than just stating that it is.

In the process of esterification, concentrated sulfuric acid is used as a catalyst. Acting as a strong dehydrating agent, concentrated sulfuric acid can rapidly remove any water produced during the reaction.



According to Le Chatelier's principle, a system will work to minimise any disturbance to its equilibrium. As such, if water is being removed, then the equilibrium will shift to favour the forward reaction in order to replace the lost water, increasing the amount of ester produced. As such, concentrated sulfuric acid can be used to increase the yield of ester.

Remember- Acid is used during esterification not only to lower the activation energy required for the reaction to take place, but also to increase the output of products.

2.5.6 Explain the need for refluxing during esterification

Refluxing involves the heating of reactants in a flask, from which vapours rise through a cooling condenser. This causes the reactants to condense and fall back into the flask, preventing them from escaping.

According to collision theory, more heat leads to more collisions between molecules, leading to a faster rate of reaction. Normally higher temperatures would also lead to faster evaporation, meaning unacceptable amounts of reactants would be lost to the environment. Refluxing keeps reactants constrained to the flask, where they can continue to react. This means higher temperatures and thus faster reactions can be achieved in a viable manner.

In addition, refluxing also acts as a safety mechanism since the vapours from esterification are flammable.

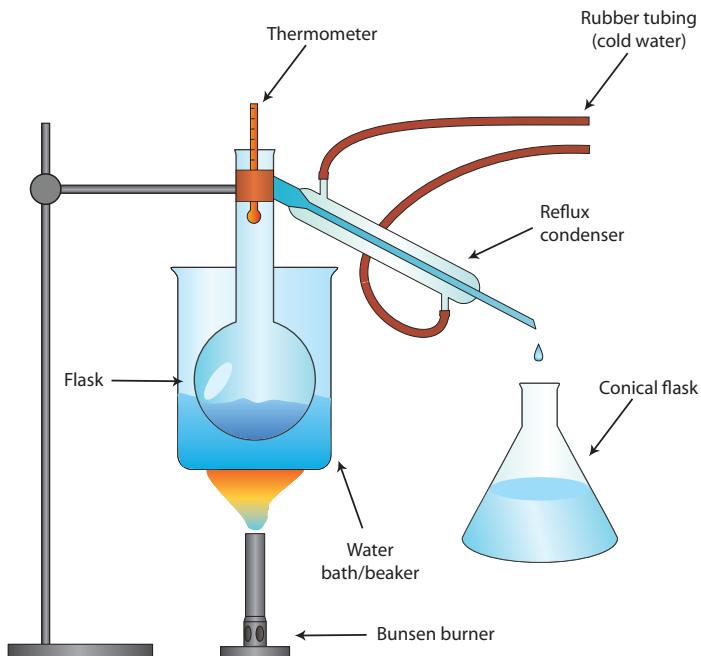
Remember- Refluxing allows for a higher temperature, and thus reaction rate to be achieved while increasing safety given the volatility of the compounds used.

2.5.7 Identify data, plan, select equipment and perform a first-hand investigation to prepare an ester using reflux

This experiment can be time-consuming if excessive amounts of reactants are used. Use only a moderate quantities of reactant and boiling chips, and ensure the mixture is properly refluxed. Excessive reactants or boiling chips will only extend the time taken for the reaction to begin.

Procedure:

1. Place 10-20ml of ethanol and an equal amount of the ethanoic acid into a 50ml flask.
2. Add 1ml of concentrated sulfuric acid into the mixture of reactants, and then add a moderate amount of boiling chips.
3. Gently clamp the flask and refluxing apparatus onto a retort stand, dipping the flask into a water bath.
4. Run a hose through the refluxing apparatus, ensuring water is entering through the bottom and exiting through the top. This ensures that water is always present in the cooling jacket, even if the water supply stops.
5. Heat the water bath using either a Bunsen burner or a hot plate, and reflux the mixture until two discernible layers form.
6. Add 50ml water into the remaining mixture, and pour into a separating funnel. Remove the lower layer. The remaining mixture should be a relatively pure ester. Sodium bicarbonate can be used to neutralise any excess acid if desired.



2.5.8 Outline some examples of the occurrence, production and uses of esters

Remembering one example of an ester may be useful. However, don't waste your effort remembering more than two. In addition to the ones listed below, two more were listed previously. These include propyl ethanoate as pear, and butyl butyrate as pineapple.

Esters can be found naturally in fruits and artificially in flavourings and fragrances such as perfumes and colognes. As esters are cheaper to produce artificially than to extract from natural sources, they are often used to replace natural flavouring and fragrances. Examples of scents and flavours include ethyl butanoate as peach, octyl ethanoate as orange, and methyl butanoate as apple.

2.5.9 Process information from secondary sources to identify and describe the uses of esters as flavours and perfumes in processed foods and cosmetics

This dotpoint simply requires you to remember the applications of a few esters. Learning two or three should suffice.

- Ethyl ethanoate is used industrially as a solvent, with its most common use being a nail polish remover.
- Other esters are also used as flavours or perfumes, given that they are an equally effective, yet cheaper alternative to natural scents and flavours. Examples treated previously are listed again below:

Name of Ester	Flavour/Odour
Methyl Butanoate	Apple
Ethyl Butanoate	Peach
Propyl Ethanoate	Pear
Butyl Butyrate	Pineapple
Octyl Ethanoate	Orange

Chapter 3

Chemical Monitoring and Management

3.1 The work of chemists

3.1.1 Gather, process and present information from secondary sources about the work of practising scientists identifying the variety of chemical occupations and a specific chemical occupation for a more detailed study

Be prepared to list perhaps three different occupations, elaborating on one.

Examples of various chemical occupations include

- Analytical Chemist- Analyses the content of substances, maintaining product quality.
- Nuclear Chemist- Produces and maintains radioisotopes
- Industrial Chemist- Produces industrial products such as ammonia, and various acids and bases.
- Forensic Chemist- Collects and analyses data for use as evidence in legal proceedings.

The forensic chemist analyses data, usually from crime scenes, for the purpose of aiding legal proceedings by introducing evidence. The forensic chemist largely works with the same tools as the analytical chemist, making use of processes such as gas chromatography and spectroscopy.

Using these processes, the forensic chemist would work alongside authorities, identifying substances and matching DNA to incriminate or clear suspects.

3.1.2 Outline the role of a chemist employed in a named industry or enterprise, identifying the branch of chemistry undertaken by the chemist and explaining a chemical principle that the chemist uses

The main aim of this dotpoint is to identify a real-life application of chemistry. In order to simplify matters, the example below draws upon concepts which will be further explored in the next topic. Understanding these concepts later on will spare you the effort of memorising another answer to address this dotpoint.

Do not be concerned if much of the Haber process makes no sense to you at this point. It has only been used in this dotpoint as an example, and will be dealt with much more thoroughly in the next topic. It is only important at this point to be able to identify a role, the branch of chemistry, and a chemical principle that may be used.

Chemist John Smith is a production chemist at a plant which produces ammonia (NH_3) through the Haber process. The branch of chemistry he undertakes is analytical chemistry, which involves the analysis of the chemical contents of substances and mixtures.

A chemical principle that John will employ is that of solubility as he uses gas-liquid chromatography to identify the substances present as the gas is passed over a liquid stationary phase. If a high concentration of noble gases are present, or the ratio of nitrogen to hydrogen is not roughly 1:3, efficiency may be reduced. In addition, if carbon monoxide or sulfur is present, the catalyst may be poisoned', resulting in the need for it to be replaced.

3.1.3 Identify the need for collaboration between chemists as they collect and analyse data

Chemists need to collaborate because:

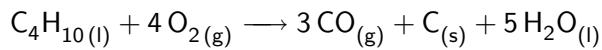
- Efficiency is increased, reducing the time taken, and thus the cost.
- Skills can be amassed and shared between chemists.
- As chemists grow increasingly specialised, collaboration is required to complete a task that may require special expertise in a variety of areas.

3.1.4 Describe an example of a chemical reaction such as combustion, where reactants form different products under different conditions and thus would need monitoring

The combustion of butane is one example of a reaction where the same reactants in different proportions and conditions can form different products.



Shown above, the complete combustion of butane results in 8 moles of carbon dioxide, and 10 moles of water. 13 moles of oxygen are required for this reaction to occur.



In this scenario, where there is insufficient oxygen, incomplete combustion can be seen to occur. In this case, carbon monoxide and soot replace carbon dioxide.

Although carbon dioxide can be harmful, it is largely absorbed into the environment. In contrast, carbon monoxide has the potential to disrupt the human body's ability to transport oxygen by affecting haemoglobin molecules. With such a dangerous product, it is evident why monitoring is required.

Remember- With excess (or sufficient) air, carbon dioxide and water are produced during a combustion reaction. Lacking enough air, the products are carbon monoxide and soot.

3.2 Monitoring in Industry- The Haber process

3.2.1 Identify and describe the industrial uses of ammonia

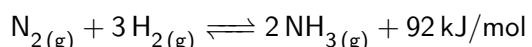
On both a domestic and industrial level, ammonia can be found in cleaning products. Such products are usually used to clean glass surfaces, and as such ammonia can be found in various window cleaners.

Another industrial use of ammonia is the production of fertilisers such as ammonium phosphate, ammonium sulfate, and ammonium nitrate. The vast majority of ammonia produced is used primarily for this purpose.

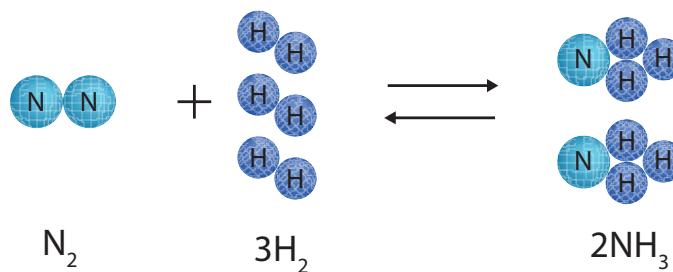
In addition, ammonia plays an important role in the production of nitric acid. This acid can in turn be used to create explosives (such as nitroglycerin), and fertilisers.

Remember- Ammonia can be used in cleaning agents, fertilisers, and the production of nitric acid

3.2.2 Identify that ammonia can be synthesised from its component gases, nitrogen and hydrogen



Observing the above equation, it is apparent that ammonia can be formed through the above reversible, exothermic reaction. The equilibrium usually rests well towards the left, and as such various considerations such as heat, pressure, and the use of a catalyst must be taken into account in order to successfully bring about the reaction.

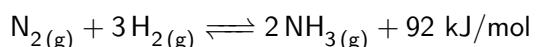


Although not crucial to this dotpoint, it may also help to note how the reactants are obtained.

Nitrogen can be obtained through the fractional distillation of liquefied air. Hydrogen can be obtained through the electrolysis of water.

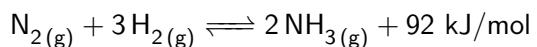
3.2.3 Describe that synthesis of ammonia occurs as a reversible reaction that will reach equilibrium

When considering this dotpoint, also recall Le Chatelier's Principle.



As shown through the symbol for a system in equilibrium (\rightleftharpoons), the above reaction is reversible. The equilibrium lies well to the left under standard temperature and pressure conditions, and as such, under normal conditions the reaction would not be viable source of ammonia.

Remember- The synthesis of ammonia is an equilibrium reaction that lies well to the left under normal conditions.

3.2.4 Identify the reaction of hydrogen with nitrogen as exothermic

For every mole of ammonia produced, 46kJ of energy is released. As such, the reaction is exothermic. Under Le Chatelier's Principle, lower temperatures would thus promote a higher yield of ammonia.

3.2.5 Explain why the rate of reaction is increased by higher temperatures

Remember- Note that when the temperature increases, both the forward and backward reactions increase.

As the temperature increases, particles move increasingly faster, thereby increasing the frequency at which they collide. This increases the chance of successful collisions, and thus increases the rate of reaction in accordance with the collision theory.

Remember- Higher temperatures increase the rate of reaction by increasing the amount of collisions between particles. However, it does not necessarily increase yield.

3.2.6 Explain why the yield of product in the Haber process is reduced at higher temperatures using Le Chatelier's principle

According to Le Chatelier's Principle, a system will act to minimise any disturbance. As such, given that the Haber process- the production of ammonia from its base elements- is an exothermic reaction, an increase in temperature will result in the equilibrium favouring the backwards reaction in an attempt to decrease the amount of heat generated. As such, the yield of ammonia would decrease.

Remember- Exothermic reactions have their yield reduced at higher temperatures. This occurs as the system works to minimise the impact of the disturbance (An increase in heat in this case).

3.2.7 Analyse the impact of increased pressure on the system involved in the Haber process

Under Le Chatelier's Principle, increased pressure would favour the forwards reaction, increasing the yield of ammonia.

3.2.8 Explain why the Haber process is based on a delicate balancing act involving reaction energy, reaction rate and equilibrium

It has been established under Le Chatelier's Principle that, as an exothermic reaction, the higher yield is achieved when the temperature is low.

However, it was also found that the higher the temperature, the faster the rate of reaction and the faster equilibrium is attained.

As such, the Haber process is a delicate balancing act. Although a lower temperature would promote a high yield, it would take far too long for equilibrium to be attained. In contrast, a high temperature would see equilibrium reached quickly, but with a small yield. A temperature between the two extremes must then be used in order to attain a moderate yield in a moderate amount of time.

Remember- Although a higher temperature promotes a higher reaction rate, it decreases yield (Which increases when the temperature decreases). As such, a compromise must be reached in order to attain an equilibrium in the Haber process which maximises yield in a reasonable timeframe.

3.2.9 Gather and process information from secondary sources to describe the conditions under which Haber developed the industrial synthesis of ammonia and evaluate its significance at that time in world history

In the early 1900s, there was a need for ammonia as a means of producing fertiliser to increase the supply of food. Ammonia was primarily gathered from guano- a substance primarily formed by the droppings of sea birds- yet an embargo was placed upon Germany's imports from South America by the Allies.

In 1908, Fritz Haber developed the method of producing ammonia from its base elements, hydrogen and nitrogen. Yet it was only in 1914 that Carl Bosch developed the means to carry out this reaction on an industrial level.

Although the discovery of the Haber process had grave consequences for the wartime effort, as explosives were manufactured from the ammonia produced, the Haber process also allowed for the production of fertilisers, which did much to help feed the German population, and indeed the world population in the post-war period.

Remember- Haber developed the method of producing ammonia in a period where ammonia was drastically needed as Germany's supply (In the form of guano) was cut off. Although the Haber process lead to the production of explosives, it also did much to relieve the German and international population by providing a steady supply of fertiliser, thereby increasing the supply of food.

3.2.10 Explain that the use of a catalyst will lower the reaction temperature required and identify the catalyst(s) used in the Haber process

Catalysts have the ability to offer an alternate reaction pathway that has a lower activation energy, thereby lowering the reaction temperature required without being used up in the reaction.

The catalyst used in the Haber process is porous magnetite, Fe_3O_4 , mixed with several other metal oxides. The high surface area of the catalyst provides a surface for the reaction to take place.

3.2.11 Explain why monitoring of the reaction vessel used in the Haber process is crucial and discuss the monitoring required

Monitoring the reaction vessel is crucial for both reasons of safety, and efficiency. The following is a list of a few reasons:

- Pressure must be maintained at a high level to promote a higher yield of ammonia. However, high pressure is both costly and dangerous to maintain.
- A moderate temperature must be used to promote a moderate yield in an acceptable timeframe.
- The inflow of the reactants hydrogen and nitrogen must be kept at a ratio of 3:1 respectively. Any large variation will result in one reactant building up, reducing the efficiency of the process.
- The presence of other gases must be restricted. Although noble gases will do nothing but reduce the efficiency of the process, the presence of oxygen poses a serious safety concern due to the risk of explosion. In addition, the presence of gases such as carbon monoxide can poison the catalyst, requiring its costly replacement.
- Products must also be monitored to ensure they are continually condensed and removed from the reaction vessel so as to drive equilibrium forward.

The actual conditions used during the Haber process include a temperature of around 400°C and a pressure of 250 atmospheres.

Remember- The reaction vessel requires constant monitoring in order to provide a constant, high yield, while maintaining safety standards by ensuring that no oxygen is present, and the pressure is not too high. The conditions chosen are a compromise between effectiveness, safety, and cost.

3.3 Chemical Analysis

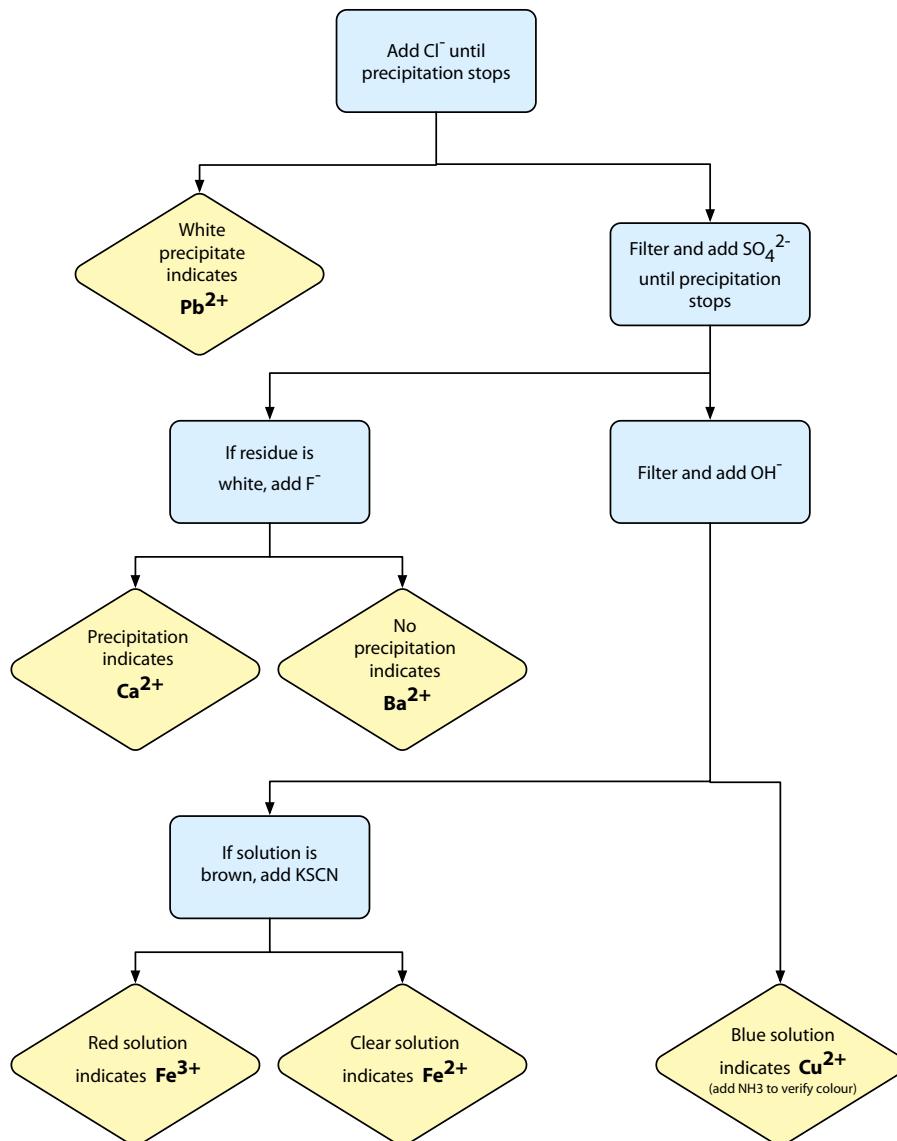
3.3.1 Perform first-hand investigations to carry out a range of tests, including flame tests, to identify the following ions: phosphate, sulfate, carbonate, chloride, barium, calcium, lead, copper and iron

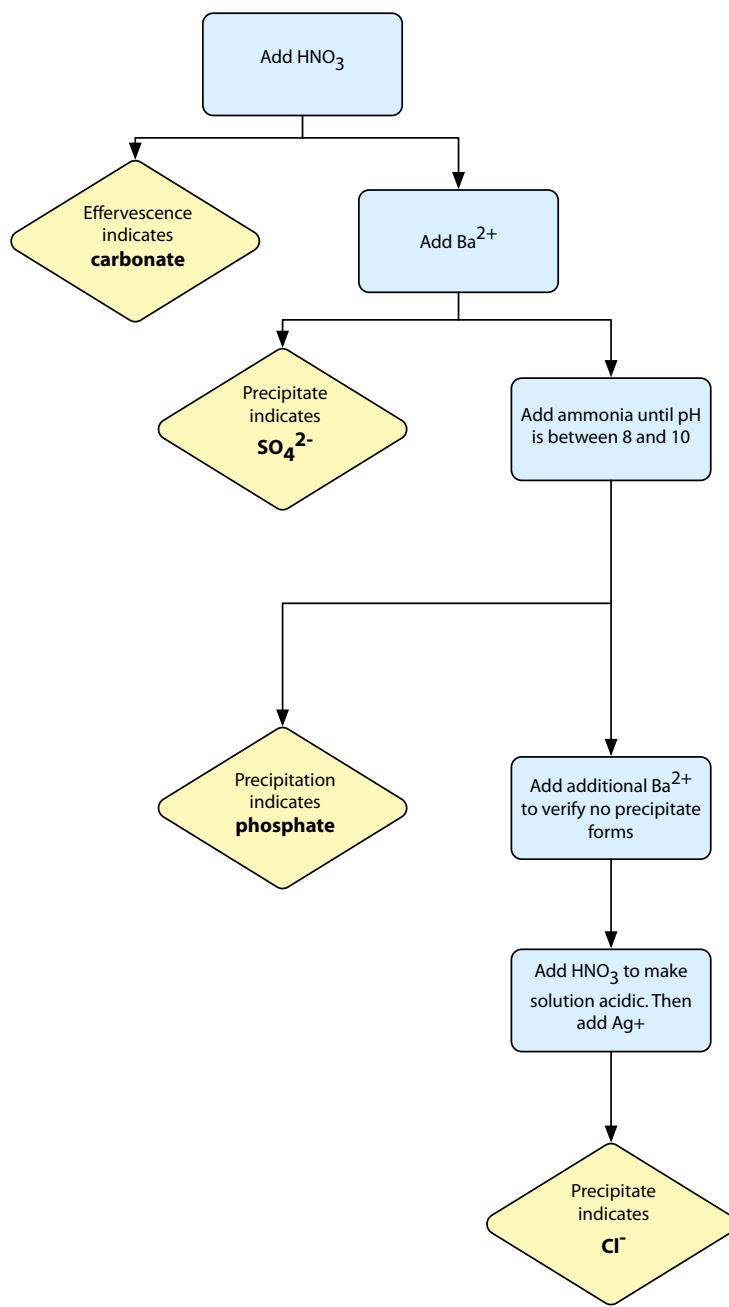
I believe the best way to demonstrate how you can test for ions is to use a flow chart illustrating each successive step given a mixture of ions. With this, you can easily work out how to test for each ion as an individual test. Although the individual tests will be listed after the flow charts to further clarify what is required, understanding the flow chart will save you memorising the table. Finally, this dotpoint will conclude with the procedure and expected results for the flame test.

Testing for Ions

The anions- Phosphate, sulfate, carbonate, and chloride ions

The cations- Barium, calcium, lead, copper, and iron ions





Testing for Individual Ions

Ion	Procedure
Phosphate	Add a base such as ammonia to the solution to increase the pH to an alkaline range (Greater than 7). Add barium nitrate, $\text{Ba}(\text{NO}_3)_2$. If a precipitate (Barium phosphate) forms, PO_4^{3-} is present in the solution.
Sulfate	Add an acid such as hydrochloric acid to the solution to decrease the pH to an acidic range (Less than 7). Add barium nitrate, $\text{Ba}(\text{NO}_3)_2$. If a precipitate (Barium sulfate) forms, SO_4^{2-} is present in the solution.
Carbonate	Use a pH probe to measure the pH of the solution. If the solution is alkaline (Has a pH greater than 7), then carbonate ion is present. Confirm by adding hydrochloric acid. The presence of bubbles (Carbon dioxide gas) will further verify that CO_3^{2-} is present in the solution.
Chloride	Add AgNO_3 to the solution. If a precipitate (Silver chloride) forms, Cl^- is present.
Barium	Add sulfuric acid, H_2SO_4 to the solution. If a precipitate (BaSO_4) forms, then either Ba^{2+} or Ca^{2+} is present in the solution. Use the flame test to confirm. (Light green flame)
Calcium	Add sulfuric acid, H_2SO_4 to the solution. If a precipitate forms, then either Ba^{2+} or Ca^{2+} is present in the solution. Use the flame test to confirm. (Brick red flame)
Lead	Add hydrochloric acid, HCl , to the solution. If a precipitate (PbCl_2) forms, Pb^{2+} is present in the solution.
Copper	Add sodium hydroxide, NaOH , to the solution. If a blue precipitate ($\text{Cu}(\text{OH})_2$) forms, Cu^{2+} is present in the solution. To confirm, the solution should turn into an even darker blue when ammonia, NH_3 , is added to the solution.
Iron	Add sodium hydroxide, NaOH , to the solution. If a white or brown precipitate (Fe(OH)_2 or Fe(OH)_3) forms then either Fe^{2+} or Fe^{3+} is present in the solution. To confirm, add potassium thiocyanate (KSCN) to new sample of the solution being tested. A colourless solution indicates that the ion present is Fe^{2+} , whereas a red solution indicates that the ion present is Fe^{3+} .

Note that Na^+ and NO_3^- are not necessary for the tests, but are simply used to ensure that, when testing for ions, precipitates are not created by accident. This is due to their soluble nature. As such, in tests such as the test for phosphate and sulfate ions, the important ion used to detect the ions in question is the barium ion.

Flame Tests

Procedure:

1. Hold a wire to a Bunsen flame. If the flame does not change colour, the wire is clean and ready to be used. If the flame changes, dip the wire into sulfuric or hydrochloric acid, and then hold the wire to the Bunsen flame.
2. Dip the wire into a solution you wish to test, and then hold the wire to the Bunsen flame. Record the change in colour, if any.

Element	Flame colour
Barium	Light Green
Calcium	Brick Red
Lead	Blue-tinged White
Copper	Blue-Green
Iron	Yellow

3.3.2 Gather, process and present information to describe and explain evidence for the need to monitor levels of one of the above ions in substances used in society

In this dotpoint, you must outline several key points. Keep it concise, but name the ion (Lead in this case), explain where it is found, and then describe its effects and thus why it must be monitored.

In society there is a growing need to monitor levels of various ions. One such example is lead, a heavy metal that has, both in the past and in the present, continued to come into contact with humans. Found in various fuels, infrastructure such as pipes, and paints for items such as houses and toys, lead poses as threat as its concentrations consistently rise as it easily remains inside organisms once inside the food chain.

Lead has been linked to damage to the nervous system, anemia, kidney and brain damage as well as schizophrenia. Deafness, speech impediments, and learning disabilities may also arise from the brain damage.

Remember- There is a need to monitor the levels of lead in a variety of environments, whether it be the soil, paints on imported toys, or marine ecosystems. This is because lead is an accumulative poison, and can be highly detrimental even when in trace levels.

3.3.3 Deduce the ions present in a sample from the results of tests

Unfortunately, this is one area of the course where memorisation plays a rather large rule, as solubility rules must be learnt. For fear of doubling up on dotpoints, this dotpoint will give an extremely brief overview of what is required, as dotpoint 3.3.1 on page 86 will have gone through each test in detail.

The identification of ions can prove important in various scenarios, where a quantitative analysis (Determining how much of the substance is present) is not required, but rather only a qualitative analysis is (Determining which substances are present).

Although this would be much simpler if it was known that a solution had only one particular ion present, this cannot always be assumed. As such, you must frequently assume that a mixture of ions is present, and make use of your knowledge of solubility rules in order to determine which ions are present. If it is only required to determine whether a specific ion is present, then simply run the appropriate test regarding that one ion, validating your answer through a flame test.

A flame test simply involves holding up a substance in a flame and observing the colour of the flame. Different ions will produce different coloured flames. The expected results of a flame test are detailed in dotpoint 3.3.1 on page 86.

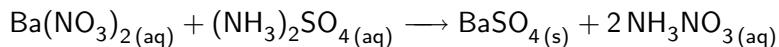
3.3.4 Identify data, plan, select equipment and perform first-hand investigations to measure the sulfate content of lawn fertiliser and explain the chemistry involved

Procedure:

1. Zero an electronic scale with a piece of filter paper on it and weigh 1g of ammonium sulfate fertiliser on the filter paper.
2. Pour the fertiliser into a 250mL beaker, adding 25mL of distilled water.
3. Add 10 drops of concentrated hydrochloric acid into the beaker.
4. Using a hotplate with a magnetic stirrer, bring the solution to boil.
5. Measure out 50mL of a 5% barium nitrate solution using a 250mL beaker, and bring to boil using a hot plate.
6. Stirring consistently with a glass stirring rod, add the barium nitrate solution into the ammonium sulfate solution. Allow the solution to cool.
7. Add 5mL of acetone to the cooling solution.
8. Weigh a sintered glass funnel.
9. Filter the solution through the sintered glass funnel. Rinse once using distilled water, and again using a small amount of acetone.
10. Dry the sintered glass funnel in a desiccator.
11. Weigh the funnel and accompanying precipitate, and calculate the mass of the remaining precipitate (barium sulfate).

Expected Results:

This experiment makes use of the fact that barium sulfate precipitates when barium nitrate is added to ammonium sulfate.



Given the mass of the precipitate at the end of the experiment, the sulfate content of the lawn fertiliser should be relatively easy to determine. Simply calculate the moles of $\text{BaSO}_4\text{(s)}$ formed using the basic formula $n = \frac{g}{M}$, where n is the number of moles, g is the mass of the precipitate, and M is the molecular mass of the molecule BaSO_4 (233.37). Given that the number of moles of the sulfate ion is equal to the number of moles in the barium sulfate precipitate, this figure n will also provide you with the sulfate content of the fertiliser.

For example, if the sintered glass funnel with filter paper weighs 90g, and weighs 92.33g with the precipitate BaSO_4 , then the precipitate weighs 2.33g, and the number of moles of the sulfate ion = $\frac{2.33}{233.37}$, which is roughly 0.01 moles.

Dotpoint 3.3.5 follows on from this dotpoint, as questions regarding methods of improving the reliability of the results of this experiment are not uncommon.

Remember- The aim of this experiment is really to isolate out a precipitate containing the sulfate ion. Once this precipitate (Barium sulfate) is obtained, the fact that the number of moles of the sulfate ion is equal to the number of moles of the barium sulfate precipitate can be used to determine the sulfate content present in the fertiliser.

3.3.5 Analyse information to evaluate the reliability of the results of the above investigation and to propose solutions to problems encountered in the procedure

Reliability is simply the reproducibility of an experiment. The more reliable a set of results, the less differences between the results. For example, if a set of results are 70g, 72g, and 67g, these results are clearly more reliable than 70g, 90g, and 50g. Inaccurate results are more often than not the result of errors which distort the result. This dotpoint seeks to identify examples of such errors in dotpoint 3.3.4 and possible solutions.

Source of Error	Solution
Measurement errors	Use a relatively large amount of fertiliser, and a highly accurate balance to weigh.
Water remaining, influencing the total weight of the sintered glass funnel.	Use acetone to separate any water from the precipitate.
Barium sulfate lost in solution	Cool the solution using ice before filtering. This will reduce solubility.
Carbonates present in the solution result in the precipitation of barium carbonate. When combined with the presence of phosphates, the ending solid is not purely barium sulfate.	The addition of concentrated hydrochloric acid neutralises the carbonate ions while increasing the solubility of the phosphate compounds.

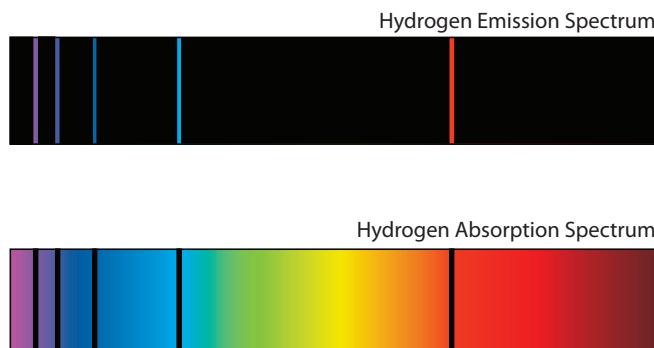
You will notice that these precautions were included in the suggested procedure in dotpoint 3.3.4. If the question asks for a method and then suggested improvements, do not include these steps to begin with, and then mention them as the improvements.

3.3.6 Describe the use of atomic absorption spectroscopy (AAS) in detecting concentrations of metal ions in solutions and assess its impact on scientific understanding of the effects of trace elements

Although the detail below regarding emission and absorption spectra is not required, it may help in giving you a better grasp of the theory behind the process in AAS.

Every substance can be said to have its own set of 'fingerprints'. When a metal atom or ion is heated, the electrons enter an excited state, moving up onto higher orbitals further from the nucleus. After a while, the electrons fall back to their original starting point, or ground state. It is possible to measure exactly how much energy, released as light energy, is absorbed by the electrons and subsequently released. When this energy, measured as electromagnetic radiation at each frequency, is recorded, the emission spectrum and the absorption spectrum of the metal atom or ion can be obtained. Each metal has a unique spectrum.

It may help to realise that a metal's emission spectrum is the opposite of its absorption spectrum. The emission spectrum appears as coloured lines on a black strip, whereas the absorption spectrum appears as black lines on a rainbow strip. When overlaid, the lines on the emission spectrum would fit exactly where the gaps are in the absorption spectrum.



The process of AAS is as follows:

1. A light source in the form of a lamp is directed at the flame. This light source emits radiation at a wavelength specific to the metal being tested. As such, every metal has its own lamp.
2. A sample of the metal being tested is placed in a flame. The concentration of this sample must be known.
3. The energy absorbed is measured. This data is used to calibrate the machine given the known concentration.
4. The metal being tested is once again fed into the flame. The light shines through the vapour, through a lens, and into the monochromator, which is a prism which splits the light into different wavelengths. A photomultiplier is used to pick up and analyse the refracted light, providing the concentration of the metal.

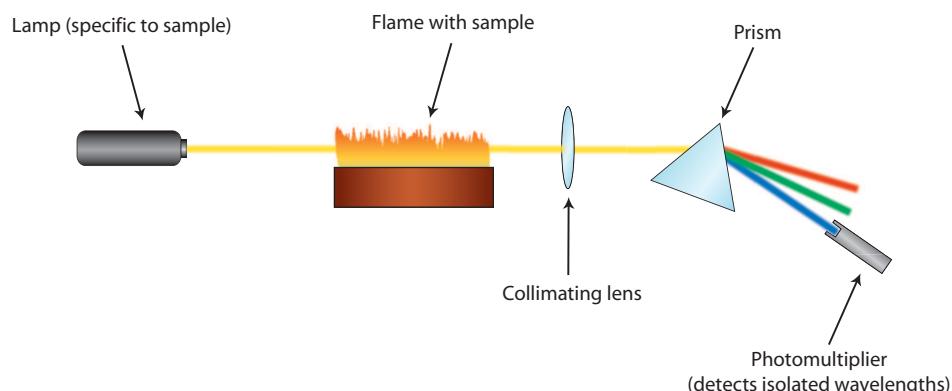
Atomic Absorption Spectroscopy has had a large impact upon scientific understanding of the effects of trace elements. The amount of trace amounts of metals within all lifeforms can now be measured, ranging from blood samples from humans, cultures from micro organism, and even soil. Pollutants and contaminants that are dangerous in even small amounts can also be picked up with great accuracy. A practical application of this is monitoring nutrient content of soil, as well as monitoring air and water pollution.

If required, provide a practical scenario of the above listed examples. For example, detecting traces of lead in human blood, or other heavy metals within soils.

Despite such advances, the limitations of AAS are also an important part of this dotpoint.

AAS, while accurate, is only effective if a quantitative reading is required. If a qualitative result is desired, i.e. simply identifying whether or not a metal is present rather than its concentration, then analysing the emission and absorption spectra is a much more viable option.

In addition, the equipment to operate AAS is extremely expensive, so start-up fees are normally rather high. On top of this, only one metal can be tested at a time, and the machine must be calibrated between, making the process relatively slow as well.



Remember- In AAS, a light is shone through the vapour of sample which has been fed into a flame. This light then goes through a lens, and refracts off a monochromator, where the light is then analysed by a photomultiplier. Each metal takes a unique amount of energy to make its electrons jump' an orbital. By measuring the difference between the quantity of energy put into the flame and the energy outputted onto the photomultiplier, it is then possible to accurately measure the concentration of the metal by figuring how many times the electrons jumped'.

This process, while highly accurate and extremely important in analysing trace amounts of certain substances, is extremely costly to setup. It is also a relatively slow process, as the concentration of only one metal can be measured at a time.

3.3.7 Gather, process and present information to interpret secondary data from AAS measurements and evaluate the effectiveness of this in pollution control

This dotpoint essentially requires you to be able to name one example of the use of AAS, and then list the advantages and disadvantages of such a process.

In addition, don't forget to take a stance after considering the above arguments. Personally I would be inclined to state that the benefits to society of being able to measure trace amounts of potentially harmful metals outweighs the costs.

One simple example is the detection and prevention of heavy metal poisoning in the oyster industry. By monitoring the levels of metals such as copper, zinc, and lead, then the adverse effects of allowing excess amounts of such metals to enter the food chain can be minimised if not outright prevented. As such, AAS plays a prominent role in pollution control on both an industrial and environmental level.

Evaluating the effectiveness of this method:

Advantages of AAS include the fact that trace amounts of a metal can be measured with a high degree of accuracy through a relatively simple process.

However, it must also be noted that the process is extremely costly to set up initially, and metals can only be tested one at a time, making the process relatively slow.

Remember- While effective in measuring trace amounts of any metal, AAS is extremely costly to setup. And while the process of determining the concentration of a sample is not too drawn out, only one substance can be measured at a time, resulting in a relatively slow process.

3.4 Atmospheric chemistry and ozone

3.4.1 Describe the composition and layered structure of the atmosphere

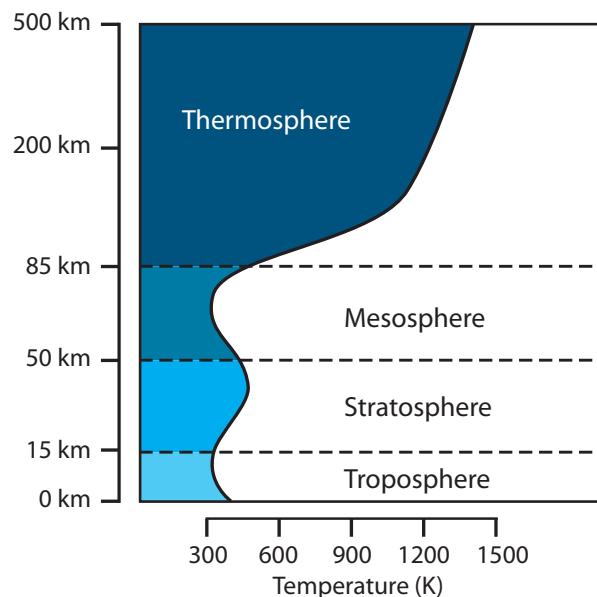
The atmosphere can be split into four main layers, namely the troposphere, stratosphere, mesosphere and thermosphere.

The lowest layer of the atmosphere is known as the troposphere, and extends from 10km above sea level to 15km above sea level. The troposphere contains approximately 75% of the atmosphere's mass, and is predominantly filled with nitrogen gas (78%) and oxygen gas (20%), as well as carbon dioxide and inert gases such as argon to a lesser extent. In addition, water vapour can be found at the lower end of the troposphere as the temperature decreases with height.

The next layer, the stratosphere, spans from 15km to 50km and mainly contains nitrogen gas, oxygen gas, and ozone. Due to the presence of this layer of ozone, the temperature increases with height as ultraviolet radiation is absorbed.

The mesosphere is the next layer after the stratosphere, extending between 50km to 85km. This layer contains many gas particles and gas ions that could not normally exist at lower altitudes. Examples include O_2^+ and O, as well as metal atoms present to the meteors which burn up as a result of colliding with the gas particles present. Temperatures decrease as the height increases.

The thermosphere spans from 85km to as high as 500km, and has very few gas particles present. However, due to the trace amounts of oxygen present, temperatures increase as height increases as solar radiation is absorbed.



Remember- The atmosphere is formed by the troposphere, stratosphere, mesosphere and thermosphere. In the lower atmosphere, nitrogen gas and oxygen gas are prevalent, with small amounts of carbon dioxide, water vapour and inert gases such as argon. A layer of ozone can then be found slightly higher up. Higher up still, gas particles and gas ions can be found here that would normally exist closer to the Earth's surface, such as O_2^+ and O. The presence of atoms and ions decreases gradually as the altitude continues to increase.

3.4.2 Identify the main pollutants found in the lower atmosphere and their sources

This dotpoint doesn't require much, as later dotpoints going into further detail. As such, the table below will be used simply to name several pollutants as well as how they enter the atmosphere.

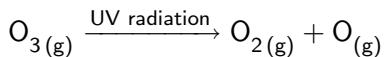
Pollutant	Sources
Carbon Monoxide	Incomplete combustion of fossil fuels.
Carbon Dioxide	Respiration from organisms, combustion
Ozone	Photochemical smog, combination of free oxygen radicals and oxygen.
Chlorofluorocarbons (CFCs)	Aerosols, refrigerants, fire extinguishers.

3.4.3 Describe ozone as a molecule able to act both as an upper atmosphere UV radiation shield and a lower atmosphere pollutant

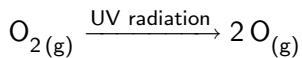
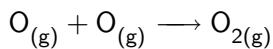
The presence of ozone is a double-edged sword, as it has a beneficial aspect despite its clear dangers. This dotpoint requires you to explore both sides. Again, be sure to include equations where appropriate.

Ozone acting as an upper atmosphere UV radiation shield

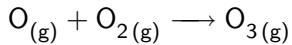
When in the stratosphere, ozone works to absorb ultraviolet (UV) radiation, using its energy to decompose into oxygen gas and an oxygen free radical.



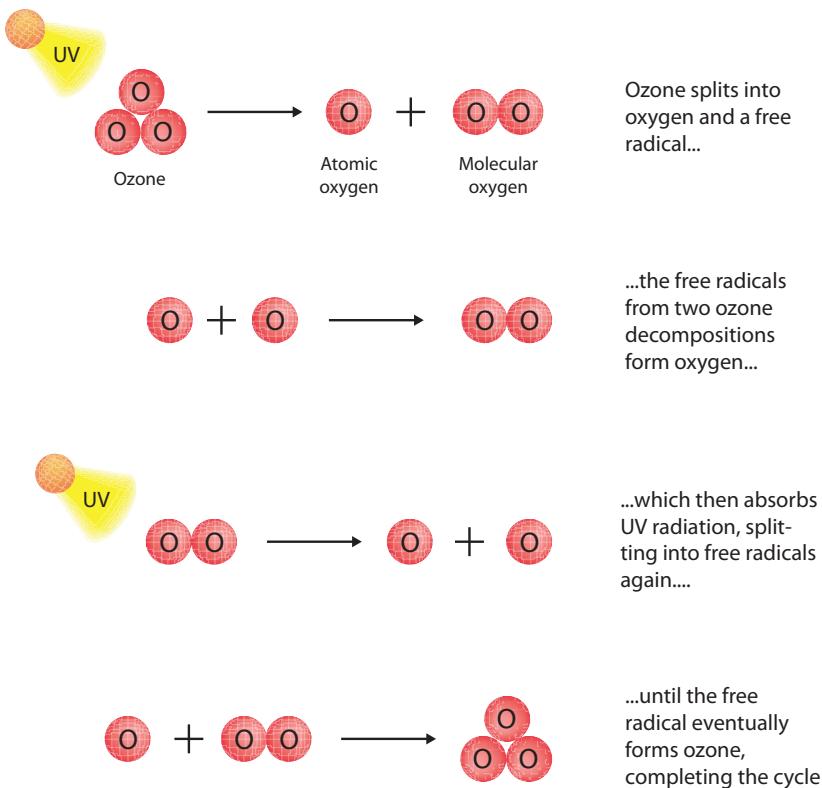
Simultaneously, oxygen free radicals combine to form oxygen gas, which in turn absorbs UV radiation when it breaks back into its component oxygen free radicals.



These radicals then react with oxygen gas to reform the decomposed ozone, completing the cycle.



The effectiveness of ozone as an upper atmosphere UV radiation shield revolves around the fact that ozone is replenished after it decomposes. As such, should any ozone be removed from the system, there could be devastating ramifications.



Ozone acting as a lower atmosphere pollutant

Even in small concentrations (even as low as 0.1ppm), ozone has the potential to cause respiratory distress in humans. Not only increasing the severity of symptoms presented by asthmatics, ozone may also induce fatigue and the onset of headaches. As such, ozone is a lower atmosphere pollutant.

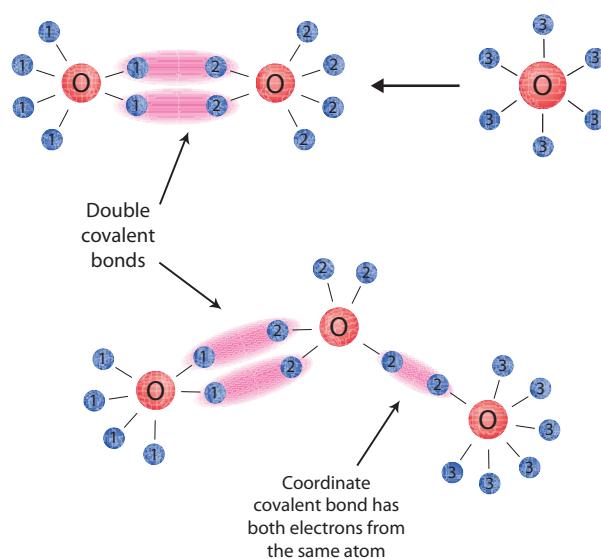
Remember- Ozone is a powerful upper atmosphere UV radiation shield due to its ability to be recycled'. However, it is always a lower atmosphere pollutant as it can cause respiratory problems and aggravate asthma attacks.

3.4.4 Describe the formation of a coordinate covalent bond

Take care to note that although a coordinate covalent bond is formed differently from a regular polar covalent bond, they have the same characteristics once formed.

When two species share electrons in order to bond, a covalent bond is formed. When one of the species provides all the electrons necessary for that bond (Rather than, say, one each, one species might provide both electrons), then a coordinate covalent bond is formed.

One example of a molecule which is formed through coordinate covalent bonding is ozone.



Remember- A coordinate covalent bond is formed when two species share electrons that are given by only one of the species.

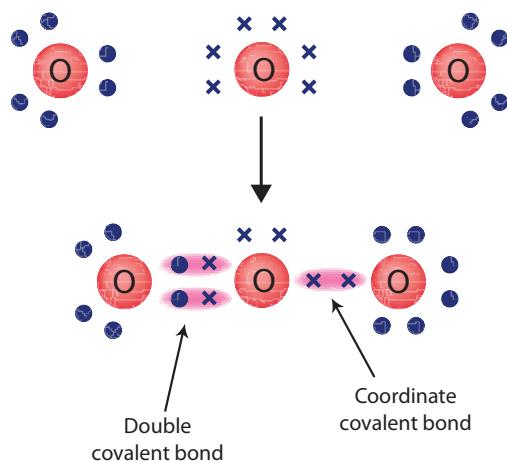
3.4.5 Demonstrate the formation of coordinate covalent bonds using Lewis electron dot structures

With an atomic number of 16, oxygen has an outer valence shell of 6 (2-8-6) and can combine with two other oxygen atoms to form O_3 , or ozone.

The bond between the first two oxygen atoms is a classic example of a (double) covalent. Both oxygen atoms contribute two electrons each to complete their outer valence shells and form a strong double covalent bond.

However, the third oxygen atom is unable to form a double covalent bond with one of the now-bonded oxygen atoms, as both that atom and the middle oxygen atom would then have 10 electrons in their outer valence shells. As such, the middle oxygen atom provides both of the electrons necessary to fill the outer valence shells on both oxygen atoms. This is done through a coordinate covalent bond.

Thus ozone is formed through one double covalent bond and one coordinate covalent bond.



3.4.6 Compare the properties of the oxygen allotropes O_2 and O_3 and account for them on the basis of molecular structure and bonding

As with most dotpoints in Chemistry, you must not only be able to state the physical property in question, but also relate it back to the structure and bonding of the molecule.

Where oxygen has a boiling point of $-183^\circ C$, ozone has a boiling point of $-111^\circ C$. This reflects the presence of stronger intermolecular forces due to the bent structure of the ozone molecule. Due to the presence of slight dipole-dipole forces in ozone because of its bent structure, ozone is also more soluble than oxygen.

Given the presence of a single bond in ozone, ozone is less stable than oxygen, which contains only a double bond. As such, ozone is more reactive than oxygen and can decompose to form oxygen and a free radical.

In addition, ozone is denser than oxygen. This is because ozone, having three oxygen atoms rather than two, has more mass per unit volume.

Remember- Ozone has stronger intermolecular forces but weaker intramolecular forces than oxygen. This translates to higher boiling and melting points, yet lower stability (and thus higher reactivity) for ozone relative to oxygen.

3.4.7 Compare the properties of the gaseous forms of oxygen and the oxygen free radical

Note that the gaseous forms of oxygen generally refer to both oxygen and ozone. Don't forget that the oxygen free radical is a single oxygen atom that stands by itself. It is generally formed through the decay of oxygen in the upper atmosphere due to UV radiation, but can also be produced through lightning in the lower atmosphere.

Where ozone is generally more reactive than oxygen due to the presence of a single coordinate covalent bond, the oxygen free radical is even more reactive as it attempts to complete its outer valence shell.

In addition, oxygen and ozone do not conduct when in solution. When in aqueous or molten form, the oxygen free radical does.

Remember- The oxygen free radical is highly reactive as has an incomplete outer valence shell. For this reason, it cannot exist for long in the lower levels of the atmosphere.

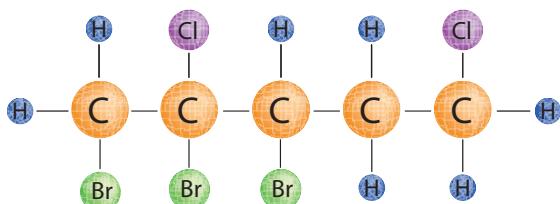
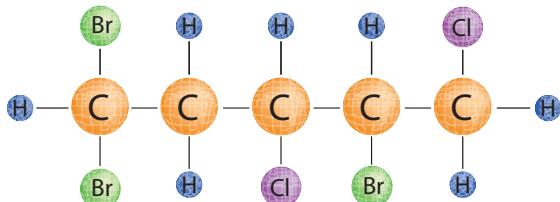
3.4.8 Identify and name examples of isomers (excluding geometrical and optical) of haloalkanes up to eight carbon atoms

This dotpoint can cause endless confusion if you do not grasp the simple definition of an isomer. Take care to note once again that a haloalkane is just a broad category of alkanes with at least one halogen attached to the carbon chain.

There exist many different molecules that have exactly the same chemical formula when simplified down. However, these molecules can have entire different structures, and look nothing alike. When a molecule same atoms as another molecule, but an entirely different structure (The atoms are put together differently), then we say that these two molecules are isomers.

Naming an isomer can be tricky at first, but becomes quite easy once you get the hang of it. First of all, if you have 2 of one atom, you normally prefix the name of the atom with di after separating all their locations on the chain with a comma. If you have 3, use the tri prefix. Secondly, if you have multiple atoms added to the chain, such as chlorine and bromine, write them down alphabetically, separating the different atoms with hyphens (-). Thirdly, make sure that you don't double up on isomers. A molecule is not a fixed thing, and can be rotated around. Therefore a ethane chain with a bromine atom on its left end is the same as a an ethane chain with a bromine atom on its right end.

Applying the above principles, here are two isomers of $C_5H_7Br_3Cl_2$:



Which are 1,1,4-tribromo-3,5-dichloropentane and 1,2,3-tribromo-2,5-dichloropentane respectively.

Other possible isomers could include 1,1,1-tribromo-2,2-dichloropentane or 1,3,3-tribromo-4,5-dichloropentane, among many other combinations.

Remember- List the different atoms alphabetically, use commas to separate the different locations, use hyphens to separate the different atoms, use di, tri prefixes, and name the carbon chain as you would a regular alkane/alkane/alkyne. Also remember that an isomer can be rotated. So 5-chloropentane should really be called 1-chloropentane and should not be double-counted!

3.4.9 Gather, process and present information from secondary sources including simulations, molecular model kits or pictorial representations to model isomers of haloalkanes

So long as you are able to name isomers, you will achieve the outcome of this dotpoint. However, do not ignore this dotpoint entirely, as viewing a physical representation of various isomers may give you a clearer picture of what is meant by a different structure in the definition of an isomer, as stated in dotpoint 3.4.8.

When a molecule same atoms as another molecule, but an entirely different structure (The atoms are put together differently), then we say that these two molecules are isomers.

3.4.10 Identify the origins of chlorofluorocarbons (CFCs) and halons in the atmosphere

A haloalkane is a regular alkane where at least one hydrogen atom has been replaced by a halogen.

Chlorofluorocarbons, or CFCs, are essentially haloalkanes where all hydrogen atoms are replaced by either chlorine or fluorine. CFCs were once used in propellants, refrigerants, foams, and cleaning solvents, as their properties (Chemically inert, moderate thermal conductivity, non-flammable, non-corrosive and non-toxic) made them an ideal replacement for substances such as ammonia.

Halons are haloalkanes where bromine atoms are present in addition to chlorine and fluorine atoms. Halons largely arose from their use in fire-extinguishers. Up until as late as 1996, halons could still be found in sources such as BCF extinguishers, but their use has since been made illegal.

Remember- CFCs originated from their use in propellants, refrigerants, foams and cleaning colvents, whereas halons primarily originated from their use in fire extinguishers.

3.4.11 Analyse the information available that indicates changes in atmospheric ozone concentrations, describe the changes observed and explain how this information was obtained

This dotpoint can go through a variety of ways. For the sake of simplicity, I will explain two various methods of measuring changes in atmospheric ozone concentrations, and explain the changes that have been observed with each. Of course, different methods do exist. However, taking note of these two should really be more than sufficient.

First developed in the mid 1920s by Gordon Dobson, the Dobson Ozone Spectrophotometer is one of the many ground-based methods of total column ozone. Using the sun or the moon as a light-source, the absorption of differing strengths of UV light radiation is measured to present an accurate presentation of total column ozone.

Total column ozone is the total ozone per unit area above a specific location, measured in Dobson Units (DU). One DU is equal to 0.01mm of pure ozone above the said location at atmospheric pressure.

Total Ozone Mapping Spectrometer, or TOMS, is the name of a satellite instrument used to map global ozone levels. The Nimbus project obtained complete readings of global levels between 1978 to 1994 through the combined effort of several satellites.

Despite such different methods, the results obtained through both ground and satellite-based methods of measurement indicate that the ozone layer has indeed diminished in thickness. Although the use of synthetic substances which damage the ozone layer, such as CFCs, have been greatly reduced, these substances have rather long lifetimes. This is due to the fact that they are repeatedly recycled through their ozone-destroying processes. As such, it will take much time before the ozone layer is restored.

It is said that 220 DU (Ozone has a depth of 2.2mm above the said location) is the base unit for which a 'hole' in the ozone layer is said to exist. This measurement occurs with some degree of consistency over parts of Antarctica. This occurs as rare clouds form in the winters of Antarctica, forming a surface upon which various oxides of nitrogen can combine with chlorine monoxide to form ClONO_2 inside ice crystals. When spring comes, these crystals melt, releasing the accumulated ozone-destroying molecules and causing weakening in the ozone layer above Antarctica.

Thus, although there are many indications that ozone levels have drastically fallen over time, this level should be gradually restored. However, this process will be rather drawn-out as CFCs slowly leave the system.

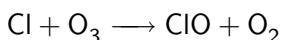
Remember- Ground-based methods such as the Dobson Ozone Spectrophotometer, as well as satellite-based methods such as Total Ozone Mapping Spectrometer (TOMS) are available to scientists for the purpose of measuring total column ozone. Both methods have indicated that although the production and use of CFCs have drastically reduce, the atmosphere will require much more time to remove itself of CFCs. However, over time, holes' in the ozone layer should be less common, and the ozone layer should be gradually restored.

3.4.12 Present information from secondary sources to write the equations to show the reactions involving CFCs and ozone to demonstrate the removal of ozone from the atmosphere

Although this dotpoint has been elaborated upon in much detail in previous dotpoints, the importance of these equations warrants a quick review.



CFCs present in the atmosphere allow the reaction between chlorine monoxide and oxygen free radicals, forming oxygen gas and a chlorine free radical.

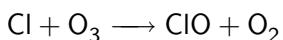
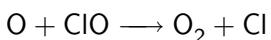


The chlorine free radical reacts with ozone in the atmosphere. This depletes the ozone layer, forming oxygen gas, as well as the original chlorine monoxide, allowing for the cycle to begin afresh.

3.4.13 Discuss the problems associated with the use of CFCs and assess the effectiveness of steps taken to alleviate these problems

You will find that although there are numerous problems associated with the use of CFCs, the majority of these are linked to the destruction of the ozone layer by CFCs. Learn the equations which illustrate this and you will find a sound foundation for any other explanations.

Although posing no real direct harm to humans, CFCs can cause a great deal of harm indirectly through their effect on the ozone layer. For example, if a single chlorine monoxide molecule is formed due to the presence of CFCs, then a chain reaction can cause a devastating amount of damage to the ozone layer. This is evident through the equations below. UV light can break apart the chlorine monoxide molecule, combining with an oxygen free radical to form oxygen and a chlorine free radical. This chlorine free radical can then in turn combine with ozone, forming back oxygen as well as the original chlorine monoxide molecule.



Although this reaction seems harmless enough, you must remember that ozone acts as an upper atmosphere UV radiation shield, breaking apart to form oxygen and an oxygen free radical. If the ozone is used up by chlorine free radicals, then we effectively lose our UV radiation shield. In addition, this problem is made worse by the fact that chlorine monoxide is still present at the end of the cycle, effectively freeing itself to start the cycle over again, removing vast amounts of ozone from the ozone layer.

With higher levels of UV radiation in the lower levels of the atmosphere, several dangers exist. The first is the danger of global warming as temperatures will increase slightly. Perhaps more dangerous still are the direct effects of prolonged exposure to UV radiation: Cell and tissue damage leading to tumours and/or cancer, formation of eye cataracts, damage to certain plants and organisms dependent upon sunlight but vulnerable to UV radiation, such as various forms of plankton.

Steps taken to alleviate the problem

Do not forget to assess the problem, i.e. make a judgment, if the question requires (As this dotpoint does).

One of the most significant steps towards alleviating the CFC problem was the ratification of the Montreal Protocol in 1987, presently with the support of close to 200 nations. The Montreal Protocol was really drafted to restrict the production and use of various haloalkanes with the purpose of preserving the ozone layer.

Of less importance but still worthy of mention are the efforts of nations in creating trade agreements for the purpose of restricting trade in substances which are dangerous to the ozone layer, as well as the research and production of chemicals to replace the extinguishers, refrigerants, propellants and other such substances once dominated by dangerous haloalkanes.

Although there is no doubt that the ozone layer was damaged immensely through years of thoughtless use of synthetic chemicals such as CFCs, the damage is not irreparable, and steps such as the Montreal Protocol will see that the problem is not made worse. With time, the ozone layer will replenish.

Remember- CFCs are dangerous because they can cause devastating harm upon the ozone layer, allowing potentially deadly UV radiation to impact upon life on Earth. The Montreal Protocol, as well as various trade and research initiatives, seeks to minimise this damage to the ozone layer by restricting the production and use of these harmful chemicals.

3.4.14 Present information from secondary sources to identify alternative chemicals used to replace CFCs and evaluate the effectiveness of their use as a replacement for CFCs

Following the introduction of CFCs, various alternatives have been identified and tested in an attempt to provide a substance with far less devastating effects upon the ozone layer than the CFCs found in propellants, refrigerants and various other compounds.

Halons have been considered by some as one alternative, yet it was quickly determined that the bromine-containing substances are far more detrimental to the ozone layer given that halons are even more reactive under UV light than conventional CFCs.

Hydrochlorofluorocarbons, or HCFCs, were then introduced as yet another alternative to CFCs. Although highly reactive, HCFCs were found to break down in the lower altitudes of the troposphere, generally breaking down completely before reaching the ozone layer. However, despite causing an estimated 10% of the damage that CFCs do, chloride ions do still enter the ozone-depleting cycle.

Following HCFCs were hydrofluorocarbons, or HFCs- substances with no trace of either bromine or chlorine. To this date, it is maintained that HFCs have absolutely no impact upon the ozone layer. However, the permanence of this solution is questionable, as members of the scientific community have found that HFCs have further worsened global warming. Despite this, it is undeniable that HFCs have been effective in offering an ozone-friendly replacement for CFCs.

Remember- Following CFCs and halons was the introduction of HCFCs, which were only 10% as harmful to the ozone layer as CFCs. The introduction of HFCs have virtually eliminated the introduction of chlorine ions into the atmosphere, paving the way for the gradual repair of the ozone layer.

3.5 Monitoring the Water Supply

3.5.1 Identify factors that affect the concentrations of a range of ions in solution in natural bodies of water such as rivers and oceans

When answering this dotpoint you will notice that a variety of answers are appropriate. Feel free to use your own subheadings, fill them up by considering how many different ways the ionic content of water can be changed. In essence, ask yourself where everything in the water can come from and you will find you are more than equipped to answer this dotpoint.

Source of water

Depending on the source of a system's supply of water, the concentrations of ions within the sample of water will vary considerably.

If the water mainly comes from rainwater overland, various nitrates and phosphates are likely dissolve in the water as it travels to the body of water.

If the water makes its way underground, then minerals such as calcium and magnesium ions, as well as sulfates and carbonates are likely to accumulate in the water supply.

Should the water combine with sewerage or waste, then the ionic content will again differ. Where phosphates and nitrates are likely to result from the sewerage, the presence of wastes can lead to the presence of various ions. Farm waste is likely to see an increase in the concentration of phosphates, nitrates, and sulfates, whereas industrial wastes are more likely to lead to an increase in the concentrations of heavy metals.

Level of nearby human activity

If significant industrial activity is nearby, then acid rain is a possibility. Should acid rain occur, then various minerals such as calcium and magnesium are likely to leach from the soils and into the waterways. Other metals are also likely to be found.

Should agricultural activity be nearby, then fertilisers are also likely to impact upon the water supply, increasing the phosphate, sulfate, and nitrate content of the water.

Should infrastructure such as stormwater drains lead to the water supply, or land be cleared nearby, then a variety of ions- both cations such as calcium and magnesium and anions such as sulfates and chlorides- may also be found.

Old or current rubbish dumps also present problems of leaching, with old batteries, wastes, metals, cleaning equipment and various other chemicals and compounds potentially making their way to sources of water. This also applies to rubbish which may be dumped into the bodies of water.

Remember- A wide variety of ions can make their way into our supplies of water in a myriad of ways. Although leaching can occur naturally, the ions introduced are not usually dangerous, with the only real consequence being the occurrence of algae blooms due to the presence of nitrates and phosphates. However, ions introduced by human activity are likely to be more dangerous to the environment, with increased human activity not only increasing the range of ions entering our supplies of water, but also the rate at which they do so.

3.5.2 Identify that water quality can be determined by considering concentrations of common ions, total dissolved solids, hardness, turbidity, acidity, dissolved oxygen and biochemical oxygen demand

Water quality can be determined by giving consideration to a variety of factors. Examples include concentrations of common ions, total dissolved solids, hardness, turbidity, acidity, and dissolved oxygen and biochemical oxygen demand.

Concentrations of common ions

The concentrations of common ions such as phosphate and nitrate ions are an important indication of water quality. Often present in waterways due to the use of fertilisers, these ions can result in the formation of algae blooms.

Total dissolved solids

Total dissolved solids (TDS) is simply the mass of solids dissolved in a sample of water, measurable in parts per million (ppm, which is equal to mg/L). Generally drinking water should have a TDS below 100ppm. Anything higher than 500ppm is not suitable for drinking by humans.

Hardness

Hard water includes water that has a high concentration of calcium or magnesium ions (Ca^{2+} and Mg^{2+}). Due to the presence of these ions, the water is unable to form a lather with soap, instead forming a grey scum which limits the effectiveness of using the soap to clean.

Turbidity

Simply the cloudiness of a sample of water, turbidity is the result of solids present in the water. If the solid is organic in nature, the oxygen content of the water will decrease as oxygen is used to decompose the solid. In addition, plants in the water are likely to suffer from less chemical energy as photosynthesis is impeded by a reduced source of light due to the presence of suspended solids.

Acidity

Acidity can simply be measured as the pH of a sample of water. Many marine ecosystems are highly dependant upon a stable and narrow pH range. With increasing industrialisation, acid rain has become more common, and threatens such marine ecosystems, both animal and plant life alike.

Dissolved oxygen and biochemical oxygen demand

The dissolved oxygen content of a sample of water is quite simply how much oxygen can be found in a sample of water (in ppm). If DO is too low, the growth of harmful bacteria which reply upon anaerobic respiration will be promoted.

Biochemical Oxygen Demand (BOD) follows directly from DO by measuring the DO necessary to completely breakdown organic matter in the water through aerobic bacteria. The lower the BOD, the higher the quality of the water as there is a higher DO level. BOD is also measured in ppm, with a BOD lower than 5ppm consistent with unpolluted water.

Remember- Concentrations of common ions, total dissolved solids, hardness, turbidity, acidity, and dissolved oxygen and biochemical oxygen demand all help in determining the quality of a sample of water when weighed up against one another.

3.5.3 Perform first-hand investigations to use qualitative and quantitative tests to analyse and compare the quality of water samples

Rather than outside a single test, this guide will go through one test for each of the headings in dotpoint 3.5.2, ensuring that you have various examples at your grasp should an exam question demand it. The method for each will be briefly outlined below.

Keep in mind that qualitative tests are used to determine a characteristic or attribute without a measurement. (E.g., that a sample of water is acidic). Quantitative tests provide a numerical result (E.g., that a sample of water has a pH of 4.3).

Concentrations of common ions

Gravimetric analysis can be used to quantitatively determine the concentration of specific ions. Water is first added to the sample, and the sample is then filtered. The filtered solution is then heated. The remaining solid is then weighed and stoichiometric calculations can be used to determine the concentration of the ion in question.

Total dissolved solids

Gravimetric analysis can also be used to determine Total Dissolved Solids (TDS), but this is not as desirable as the amount of solids in the sample is often quite low, meaning an accurate result is not likely. Instead, electrical conductivity is a preferred alternative. The dissolved solids are often salts, meaning that they are capable of conducting electricity. As such, a conductivity metre can be used to determine the TDS of a sample.

Hardness

Don't forget that hardness refers to the amount of calcium and magnesium in a sample of water.

The level of calcium and magnesium can be measured using gravimetric analysis. However, a more accurate method can be found through the use of EDTA titration.

You are not expected to know how to conduct titrations using EDTA. Simply be aware that such a procedure exists.

Turbidity

Turbidity can be measured through the use of a turbidity tube. Water is poured into the tube, and the turbidity is measured by examining which lines marked on the tube can be seen. The units of measurement are known as NTUs, or Nephelometric Turbidity Units.

Acidity

The pH of a sample of water can be quickly determined using a pH metre. If more accurate measurements are required, a pH probe can be used (After it has been calibrated).

Dissolved oxygen

Dissolved Oxygen (DO) can be determined through the use of a specific titration procedure known as the Winkler Method.

Again, you are not expected to know the science behind the Winkler Method. You are only expected to know that it is used to measure dissolved oxygen content.

Biochemical oxygen demand

Biochemical Oxygen Demand (BOD) is relatively easy to calculate once the principles behind DO are understood. Given that BOD is simply the amount of oxygen required for the complete breakdown of organic matter by aerobic bacteria, the BOD would simply be the oxygen consumed over the length of time required for complete breakdown (Generally less than five days) in the absence of light (So that no photosynthesis can occur and thus no new oxygen can be produced). Keep this explanation in mind when considering the relatively simple process below.

In order to measure BOD, a sample of water is divided evenly into two. The dissolved oxygen content of one sample is then measured, while the other sample is placed in a location away from any source of light, such as a cupboard. After five days, the dissolved oxygen content of the second sample is then measured. The difference in DO between the first and second samples is the BOD.

3.5.4 Gather, process and present information on the range and chemistry of the tests used to identify heavy metal pollution of water and monitor possible eutrophication of waterways

Tests used to identify heavy metal pollution of water

Heavy metals can be loosely defined as metals between copper (Atomic number 29) and bismuth (Atomic number 83) on the periodic table.

Qualitatively, tests such as the flame test and precipitation tests can be used to determine if a heavy metal is present in a sample of water. If a quantitative result is required, then colorimetry or Atomic Absorption Spectroscopy (AAS) may be used.

If you've forgotten what a flame test or precipitation test was, refer to dotpoint 3.3.1 on page 86. For information regarding AAS, refer to dotpoint 3.3.6 on page 92.

Tests used to monitor possible eutrophication of waterways

Eutrophication is a process whereby water is converted into a swamp, and then into firm land. This process is a naturally-occurring process, resulting from the presence of algal blooms and dead animal and plant matter. However, in recent years it has become evident that human activity has contributed to the speed of the process and as such it has become increasingly important that the situation is carefully monitored.

On a broader, observational level, eutrophication is often evident through the presence of algal blooms (Evident from a distance through the green tinge in the water), the smell of hydrogen sulfide, a degree in the strength of the current (stagnation), and a reduction in the presence of marine organisms in the water.

However, the potential for eutrophication can be more accurately measured through an analysis of the presence of the two key substances present during algal blooms- nitrates and phosphates.

Both nitrates and phosphates can be measured semi-quantitatively through the use of spectrometry. If 1ml of the sample is heated gently (In a fume cupboard) with a copper strip and three drops of sulfuric acid, nitrogen dioxide gas will be produced if the nitrate ion is present. With the phosphate ion, the addition of sodium molybdate solution and hydrazine sulfate should turn the solution blue if the ion is present. In both cases, the tests are semi-quantitative as the amount of gas and the strength of the blue will indicate the concentration of the nitrate or phosphate ions respectively.

If these samples are tested using a colorimeter, a quantitative sample can also be determined to a degree of accuracy after the colorimeter has been calibrated. This is possible as the colorimeter can analyse the amount of light passed through the sample against the concentration of the sample.

It is generally accepted that the nitrate ion is found in such abundance that it is in fact the phosphate ion which determines the rate at which eutrophication progresses. Typically, a healthy sample will have a nitrate:phosphate (n:p) ratio of 20:1. A relatively clean sample will have a n:p ratio of 10:1.

Remember- Precipitation tests and flame tests can be used to qualitatively determine the presence of heavy metals in a sample of water. Colorimetry and AAS can be used to quantitatively determine the concentration of these heavy metals.

Eutrophication can be monitored through the levels of nitrates and phosphates in a sample of water. The ratio of these ions is known as the n:p ratio, and is not to be confused with the neutron:proton ratio used in nuclear chemistry. The lower this ratio, the more polluted the sample, and the higher the likelihood and rate of eutrophication.

3.5.5 Gather, process and present information on the features of the local town water supply in terms of catchment area, possible sources of contamination in this catchment, chemical tests available to determine levels and types of contaminants, physical and chemical processes used to purify water, chemical additives in the water and the reasons for the presence of these additives

Many of the tests used in this dotpoint are simply a review and application of the tests outlined in previous dotpoints. Be sure to go over these dotpoints if you do not understand them, as it is important that you understand what exactly they serve to do in a practical context as much as a theoretically context.

Warragamba Dam is Sydney's main storage dam, accounting for 80% of Sydney's water with its catchment area of approximately 9,000km².

Resting on the Warragamba River, there exist many possible sources of contamination in this catchment. Examples include nearby agricultural zones, as fertilisers and chemicals may leach through the soil and into the river system. Wildlife also present another source of contamination, with faecal matter and corpses promoting harmful bacteria in the water. However, the largest influence is perhaps the presence of residential zones in the area, as sewerage and rubbish may find its way into the river system and thus into the dam.

Various tests outlined in dotpoint 3.5.3 on page 108 may be used in order to identify the quality of the water in Warragamba Dam. However, in addition to these tests various bacteria are also tested for. The first two, *Cryptosporidium* and *Giardia*, are easy to comprehend as these two organisms have been heavily publicised for their ability to invade water supplies and cause health crises. The next pathogen is *E. Coli*. Not terribly dangerous by itself, *E. Coli* levels are frequently measured as it is what is known as an indicator organism'. Generally speaking, if *E. Coli* levels rise, there is a strong likelihood that the levels of other bacteria have also increased. As such, there are strict guidelines regarding levels of *E. Coli* (Among tests for other organisms) which must be abided by.

Regarding physical and chemical processes used to purify the water from the dam, the same process of screening, flocculation, clarification, filtration, chlorination and fluoridation outlined in dotpoint 3.5.6 on page 112 are used.

Chemical additives in the water are partially covered by the above paragraph, as it was noted that chlorine is added in order to kill off bacteria present in the water and inhibit their growth in the pipes. Fluorine is added in accordance with legislation for dental health reasons. Not mentioned above, but added from time to time are substances designed to regulate the pH of the water supply. Generally hydrochloric acid and sodium hydroxide are added in small concentrations if a large change in pH is required. However, more generally sodium carbonate is added due to its amphoteric nature (Ability to be used as both an acid and base).

*Remember- The Warragamba Dam is the largest source of Sydney's water at over 9,000km². Agriculture, wildlife and residential areas all influence the quality of the water in the dam. Although various tests are used to monitor the quality of the water, *Cryptosporidium*, *Giardia*, and *E. Coli* examples of bacterium specifically tested for in order to maintain health standards. Screening, flocculation, clarification, filtration, chlorination and fluoridation is the generally procedure followed in purifying the water supply, as well as the addition of substances such as sodium carbonate periodically in order to maintain pH levels.*

3.5.6 Describe and assess the effectiveness of methods used to purify and sanitise mass water supplies

Various methods exist for the purification and sanitisation of water supplies. The method outlined below is but one example of one of the more general procedures. Don't forget to make a judgment on the effectiveness of the method if the question requires, as suggested by the assess' dotpoint.

The first stage in the preparation of water is known as screening. Water is passed through a series of large screens in order to remove larger foreign objects from the water supply.

The next stage is known as flocculation. Coagulants such as alum ($\text{Al}_2(\text{SO}_4)_3$) mixed with lime ($\text{Ca}(\text{OH})_2$) are added to the water, resulting in the formation of larger particles as the solids, iron, and bacteria present in the water are absorbed into the coagulant.

Clarification then follows, where the coagulated particles are allowed to settle at the bottom.

The water is then passed through a series of filters to remove any remaining particles in a process known as filtration. Note that these filters are much finer than those used in the first screening process, and are most commonly sand, gravel, or anthracitic coal. More advanced water purification industries will make use of devices known as membrane filters to further filter out any bacteria or particles that were not already caught.

Chlorine is then added into the water to kill off any surviving bacteria in a process known quite simply as chlorination. A further step known as fluoridation is then taken at some plants (All in NSW in accordance with legislation) where fluoride is added into the water supply for dental health reasons.

Despite public health scares which arise with a degree of frequency regarding outbreaks of *Cryptosporidium* and *Giardia*, the NSW water purification and sanitisation system is rather effective, with the vast majority of supplies across NSW having 100% of samples meeting criteria regarding the presence of micro-organisms. Very rarely will any of the catchment readings of *Cryptosporidium* and *Giardia* exceed 0/100L, indicating that no organisms are found in a sample of 100 litres. Although there is no doubt that the installation of membrane filters would provide an even higher quality of water within NSW, the high costs of such technology cannot be justified with the minimal improvements in an already-effective system of purification and sanitisation.

Remember- The steps are Screening, Flocculation, Clarification, Filtration, Chlorination and Fluoridation.

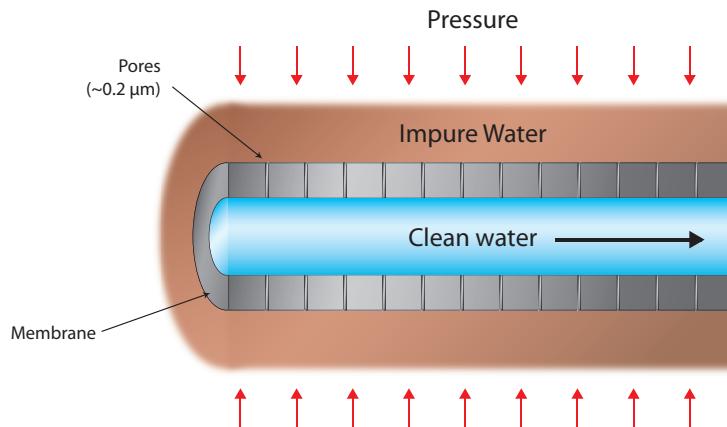
3.5.7 Describe the design and composition of microscopic membrane filters and explain how they purify contaminated water

A membrane filter is an extremely fine filtering system most commonly made of synthetic polymers such as polypropylene and polytetrafluoroethylene (Although glass, metal, and even ceramics have been used in the past). Formed off uniform-sized pores roughly $0.2\text{--}0.5\ \mu\text{m}$ in size, these filters are often layered upon each other in order to provide an even finer level of filtration.

Pressure, developed through either a pump, gravity, or even a vacuum, is used in order to force the water through the filter. In doing so, all particles, bacteria, and even viruses greater than the size of the pores are trapped outside the filter.

However, substances that are dissolved in the water such as dissolved anions cannot be filtered as they pass through the filters within the water.

Although membrane filters offer the advantage of a much higher level of filtration, as well as a degree of reusability, their costs can reach rather high levels as clogged filters require replacement and the filters require constant flushing in order to ensure that they function efficiently.



Remember- Water is pushed from the outside of a membrane filter and the result is collected from the insides. Membrane filters are so effective as their extremely small pores allow water to pass through while still denying access to virtually all other foreign solids. Dissolved substances cannot be filtered.

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Chapter 4

Industrial Chemistry

4.1 Resources and Replacements

4.1.1/4.1.2 Identify data, gather and process information to identify and discuss the issues associated with the increased need for a natural resource that is not a fossil fuel and evaluate the progress currently being made to solve the problems identified (including 'Discuss the issues associated with shrinking world resources with regard to one identified natural product that is not a fossil fuel, identifying the replacement materials used and/or current research in place to find a replacement for the named material')

Due to the strong link between the above two dotpoints, they have been joined together for the purpose of this study guide so that a single, more concise response is possible. Once again, if this material appears as an exam question, make sure that proper attention is given to the verb used, as although the content will be provided below, it is up to you to ensure that it is used in a manner relevant to the question.

Rubber is a natural resource that has grown increasingly popular over time as its unique elastic structure has proved a ready solution to many problems, resulting in a myriad of uses.

Originally native to South America, Rubber occurs naturally as latex in the sap of plants and trees such as the Para rubber tree. Introduced into South Asia by the English, rubber was grown predominantly in large plantations in these countries, and today the production of natural rubber in Indonesia, Malaysia and Thailand alone accounts for well over half of the world's natural rubber production.

With the arrival of automobiles, the demand for rubber soon escalated rapidly as rubber was seen as the ideal material for the production of tyres. An added increase in tyres for military vehicles during World War II, on top of ever-increasing demand related to the production of gloves, adhesives, conveyer belts, pencil rubbers, rubber bands amongst many other products soon made the global population realise that current production rates of rubber were simply insufficient for future needs. Thus although natural rubber sources were renewable, their availability was in fact shrinking relative to global demand.

It is under this backdrop that American and German scientists developed synthetic rubber. More chemically resistant and physically durable than natural rubber, synthetic rubber was found to be a much more desirable alternative as it provided a superior material at a cheaper price. Common synthetic rubbers found today include butadiene rubber, chloroprene rubber, ethylene-propylene-diene rubber, and styrene-butadiene rubber (SBR), with SBR being the most common. Collectively, 80% of global rubber production now consists of synthetic substitutes.

Yet although lower manufacturing costs for a product that is both chemically and physically superior to its natural counterpart is clearly appealing, it must be noted that synthetic rubbers are derived from fossil fuels, a non-renewable resource. If an alternative source of energy can be harnessed, such as the use of ethanol converted from biomass, then there is no doubt that the potential for the use of rubbers will be unlimited. Indeed, as is there is no doubt that the advantages of producing synthetic rubber clearly outweighs any disadvantages, and as the manufacturing process is continually refined, so too will the synthetic rubber products improve.

Remember- Due to a myriad of uses because of the chemical and physical properties of rubber, the demand for rubber soon outpaced its supply. Today, over 80% of rubber is synthetic, as the material is chemically and physically superior while cheaper to produce. However, synthetic rubber is derived from fossil fuels, and as such cannot be sustained indefinitely. One solution would be the arrival of fuel such as ethanol derived from biomass.

4.2 Equilibrium and the Equilibrium Constant

4.2.1 Identify data, plan and perform a first-hand investigation to model an equilibrium reaction

This experiment really just aims at showing you what equilibrium 'looks' like by using a model. Provided you understand Le Chatelier's Principle thoroughly, I would not overly concerned about this experiment.

Materials:

- Two 50mL measuring cylinders
- Two pipettes of differing diameters (e.g. one 10ml, one 5ml)

Procedure:

1. Fill one 50mL measuring cylinder with water, leaving the other measuring cylinder empty.
2. Place one pipette into first measuring cylinder, letting the pipette lightly touch the bottom. Place your finger over the top of the pipette and move its contents to the second measuring cylinder, letting your finger off the top and releasing the water. Do so carefully to ensure minimum spillage.
3. Place the other pipette into the second measuring cylinder, repeating the second step by emptying the contents of the second pipette into the first measuring cylinder.
4. Note the amount of water in both measuring cylinders and tabulate the results. These results represent the results after 'Cycle 1'. Repeat steps 2 & 3 until Cycle 30 is recorded. Graph the results against each other, with volume on the y-axis and cycle-count on the x-axis.
5. Repeat steps 1 to 4, but transfer 10mL of water from the second measuring cylinder to the first after the 15th cycle and continue as normal. Graph the results using the same axis as before.

Expected results:

You will find that equilibrium will be seen as the water levels on both cylinders tend towards, but do not reach, the halfway mark- 25ml of water in each measuring cylinder. The change will be relatively quick at first, but will eventually slow down to the point where the water level does not change considerably.

Step number 5 will lead to a temporary spike in the volume of water in the first measuring cylinder, and a decrease in the second measuring cylinder. However this will also be rapidly corrected at first, and then more gradually after a while.

4.2.2 Choose equipment and perform a first-hand investigation to gather information and qualitatively analyse an equilibrium reaction

Given that the nature of this experiment is qualitative, this experiment is quite easily accomplished, as all you need to do is disturb a system in equilibrium, and note the changes. Clearly you must pick a disturbance which has a visible effect, and the simplest disturbance is simple a change in temperature.

Select two ampoules of nitrogen dioxide (NO_2) gas. Place one ampoule in a beaker of warm-hot water, and one in a beaker of cold water. You will soon notice that one ampoule- the one which was heated- has turned a reddish brown, whereas the other- the one that was placed in cold water- has become almost completely colourless.

In effect, you are simply observing the following equilibrium:



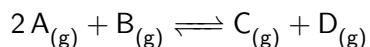
As NO_2 is a reddish brown, and N_2O_4 is colourless, we can effectively deduce that this equation is endothermic (Note that if the reaction was reversed such that nitrogen dioxide gas was a reactant it would be exothermic).

Confirm this by stating the effects as per Le Chatelier's Principle. If the equation is endothermic, then an increase in heat should see a shift in the equilibrium towards the right, making the gas appear more reddish as more NO_2 is produced. If there is a decrease in heat, then the gas would become colourless as more N_2O_4 is produced. The results are consistent with this hypothesis.

4.2.3 Explain the effect of changing the following factors on identified equilibrium reactions: Pressure, volume, concentration and temperature

This dotpoint is simply a review of Le Chatelier's Principle. As such, please refer to dotpoint 2.2.4 on page 47 for a detailed treatment of the impact of disturbances to equilibria. However, a quick summary will be provided below.

Using the following equation as an example, and assuming it is an exothermic reaction:

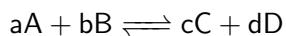


An increase in pressure would shift the equilibrium to the right. An increase in the volume of the substances is essentially an increase in pressure (Same space, more atoms). As such the equilibrium would shift once more to the right to account for this. An increase in the concentration of one substance will result in the system working to minimise this change. For example, an increase in B would result in a decrease in A and B and an increase in C and D. Therefore the equilibrium has shifted to the right. An increase in temperature in an exothermic reaction will result in a shift in the equilibrium to the left as the system works to reduce the amount of heat produced.

Remember- A closed system will always work to minimise the impact of a disturbance upon the system.

4.2.4 Interpret the equilibrium constant expression (no units required) from the chemical equation of equilibrium reactions

Let the following equation be any given chemical equation, where A, B, C, and D are any given substances, and a, b, c, and d are their respective molar ratios:



When this equation is in equilibrium, a constant, K also known as the equilibrium constant, can be obtained using the following expression:

$$K = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

It helps to keep in your mind that the products are always on top of the reactants when calculating K. One way to remember this is that P comes before R alphabetically.

K provides much information if it can be read successfully, as it tells us how far to completion an equilibrium is currently at. Think of it as at a scale. At one end no reactions have occurred and there are only reactants and no products. At the other end, everything has been reacted and there are no reactants, only products. A point of equilibrium always lies between these two points. The value of K tells us exactly where on this scale the equilibrium is currently at.

The smaller the value of K , the lower on the scale you'll find the equilibrium- Mostly reactants, very little product. (Generally where $K < 10^{-4}$).

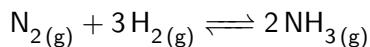
The larger the value of K , the higher up on the scale you'll find the equilibrium- Mostly products, very little reactant. (Generally where $K > 10^4$).

In between these two values, there is a moderate mixture of products and reactants, with the ratio depending upon the exact value of K , and therefore where on the scale the equilibrium can be found.

When a reaction is not at equilibrium, Q replaces K in the same expression. If Q is less than K, then the reaction is lower on the scale than the point equilibrium, and thus the concentration of products must be increased in order to achieve equilibrium. Conversely, if Q is larger than K, then the reaction is higher on the scale than the point of equilibrium, and thus the concentration of reactants must be increased in order to achieve equilibrium.

Remember- K is the equilibrium constant, calculated by multiplying the concentrations of the products to the power of their respective number of moles. A lower value of K indicates an equilibrium where very few products are produced, whereas a high value of K indicates that there is an equilibrium which is close to completion.

Example:



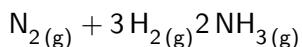
If the above equation is at equilibrium when there are 3mol of nitrogen gas, 2mol of hydrogen gas, and 2mol of ammonia, then:

$$K = \frac{3 \times 2^3}{2^2} = 6$$

4.2.5 Process and present information from secondary sources to calculate K from equilibrium conditions

The easiest way to grasp how to calculate K is really to go through a worked example.

Continuing from the Haber process example, let us begin with the equation:



At the beginning of an experiment, there were 2.1mol of nitrogen gas, and 6.9mol of hydrogen gas. The reaction was allowed to proceed to equilibrium in a 10L container, at which point 1.2mol of N₂ was remaining. What is the value of K, assuming a fixed temperature?

The first step is to note the concentration of each item individually at equilibrium. If 1.2mol of N₂ is remaining, then 0.9mol must have been converted. Therefore 3 x 0.9mol of H₂ must also have been converted, leaving 0.3mol of H₂ (For every one mole of nitrogen gas, 3 moles of hydrogen gas are converted, as per the chemical equation). In addition, there must now be 2 x 0.9mol of NH₃.

We now know the concentrations of each substance. The next trick is to note that, when dealing with gases, you must account for the size of the container, as it is very rarely a simple 1L in exams. You must do this because concentration is proportional to pressure, and the size of a container influences the pressure.

Once the 10L has been taken into account, K can be calculated:

$$K = \frac{1.2^{0.1} \times 0.3^{0.3}}{1.8^{0.2}} = 0.6310$$

Note that the indices have been adjusted (divided by 10) to account for the 10L container.

4.2.6 Identify that temperature is the only factor that changes the value of the equilibrium constant (K) for a given equation

If volume, concentrations or pressure change, then the numerator and denominator used in the calculation of the equilibrium constant K shift correspondingly, cancelling out the effect of one another. As such, these disturbances do not impact upon K.

However, when the temperature is changed, then K does in fact change.

For endothermic reactions, if the temperature increases then K increases (Remember how equilibrium 'shifts to the left' according to Le Chatelier's Principle. This means it moves up the scale.) Thus for exothermic reactions, if the temperature increases then K decreases.

Remember- A change in temperature is the only factor which changes the value of K. Changes in volume, concentration, and pressure all have no effect on the value of the equilibrium constant.

4.3 Sulfuric Acid

4.3.1 Outline three uses of sulfuric acid in industry

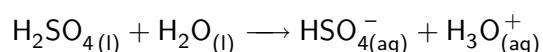
This dotpoint is largely introductory, so don't feel that you need to go into a great deal of detail for three separate uses- the keyword is 'outline'.

- Dehydration: Sulfuric acid is a powerful dehydrating agent. It is used to increase yield and act as a catalyst during esterification, and is also used in the dehydration of ethanol into ethylene and water.
- Cleaning metals: Sulfuric acid is used to clean pieces of metal, a task that is necessary before a piece of metal can be galvanised or coated in general.
- Production of fertilisers: Sulfuric acid plays an important role in the production of a few different fertilisers. In a coke oven, sulfuric acid can extract ammonia from a mixture of gases by reacting with the gas. This can later be used to produce ammonium sulfate fertiliser. In addition, sulfuric acid can convert insoluble calcium phosphate into a soluble mixture which can then be used in phosphate fertilisers.

Remember- Sulfuric acid can be used in the dehydration of ethanol and alkanols & alkanoic acids during esterification. Sulfuric acid can also be used to clean metals, as well as produce ammonia sulfate and superphosphate fertilisers.

4.3.2 Describe and explain the exothermic nature of sulfuric acid ionisation

The ionisation of sulfuric acid is a highly exothermic reaction, meaning that a great deal of heat is released upon ionisation. This is because concentrated sulfuric acid is virtually 100% molecular, meaning there are very little ions. Upon ionisation, the breaking of the covalent bonds present in the acid releases a vast amount of energy, thus leading to an exothermic reaction.



$$\Delta H = 90 \text{ kJ/mol}^{-1}$$

4.3.3 Identify and describe safety precautions that must be taken when using and diluting concentrated sulfuric acid

Safety is a point that has really been pushed in the new HSC. As such, make sure you are prepared to answer safety questions on any topic, in any module. Fortunately, safety procedures are fairly common sense, as the following list will prove.

- Wear full protective lab gear when conducting this experiment. This includes a lab coat and goggles. Appropriate footwear and gloves are also recommended.
- Have a supply of water, such as a tap, nearby should you need to flush a part of the body with running water.
- Always add acid to water, never the other way around.
- Store and use sulfuric acid in volumes smaller than 1 litre.
- Ready a base in the case that a spill must be neutralised. Sodium carbonate is one such example.
- Avoid conducting experiments at eye level, or near the face.

4.3.4 Use available evidence to relate the properties of sulfuric acid to safety precautions necessary for its transport and storage

As mentioned previously, concentrated sulfuric acid is virtually 100% molecular. For this very reason steel containers can be used to transport concentrated sulfuric acid, as there are no ions present to attack the metal.

If dilute sulfuric acid is being transported however, glass is the preferred medium of transport, followed by plastic.

In general, regardless of the method of transport, sulfuric acid should be split into many small containers, minimising the danger on the off-chance that a container breaks or there is a contamination of concentrated sulfuric acid with water.

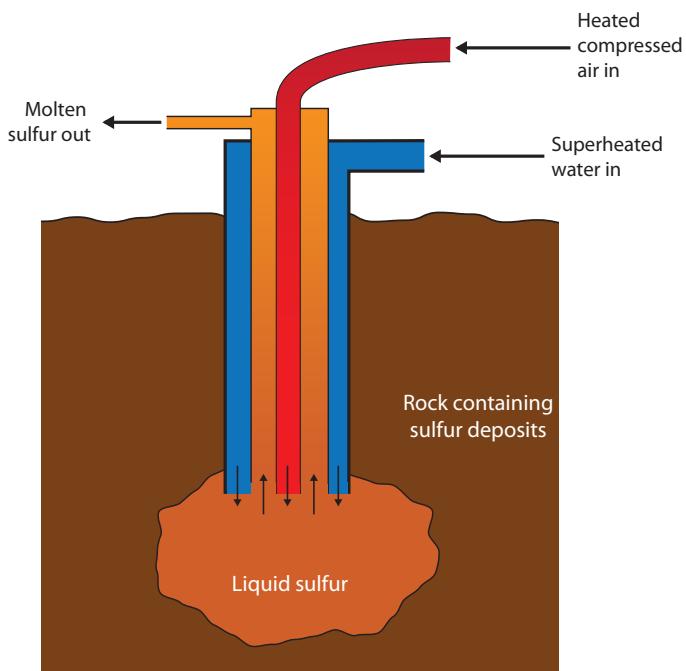
In addition, excess amounts of a substance to neutralise the spill is highly recommended. One example is sodium carbonate, which even when used in excess is not likely to create a situation of danger.

4.3.5 Describe the processes used to extract sulfur from mineral deposits, identifying the properties of sulfur which allow its extraction and analysing potential environmental issues that may be associated with its extraction

Note that this dotpoint is essentially three points in one. To make sure you cover all possible examinable materials, as well as to break this dotpoint into more manageable portions, this guide will treat each point separately under different subheadings.

Processes used to extract sulfur from mineral deposits

One of the more common methods of extracting sulfur from mineral deposits is known as the Frasch Process. Using this process, one large pipe is placed into a sulfur deposit. Within this pipe are three pipes, all placed concentrically (All pipes form circle with equal centres, but differing diameters). Water superheated to 160°C is funnelled down the outermost pipe, melting the sulfur in the deposit, which combines with the water to form an emulsion. Compressed air is then sent down the innermost pipe, pushing the emulsion up to the surface through the middle pipe.



Upon cooling, 99.5% pure sulfur is obtained as it separates from the water due to its insolubility.

Properties of sulfur which allow for its extraction

Two properties of sulfur allow for its extraction through the Frasch Process. The first is its relatively low melting point of 113°C, which allows the sulfur in the deposits to dissolve in the water and be pushed up to the top. The second property is the insolubility of sulfur, which allows for the easy extraction at the surface.

Environmental issues associated with the Frasch Process

The sulfur brought to the surface can be readily oxidised or reduced to form sulfur dioxide or hydrogen sulfide respectively. Both of these pollutants can have severe consequences even when in small concentrations, and as such the water used to flush the sulfur from the deposit must be reused, or recycled before it can be disposed of.

In addition, should the sulfur dioxide find its way into the atmosphere, large concentrations will result in the formation of acid rain, which can have serious repercussions upon both manmade infrastructure as well as flora and fauna, particularly marine ecosystems.

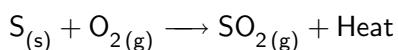
Remember- Sulfur can be extracted from mineral deposits through the Frasch Process. Superheated water is flushed down the outermost pipe of a set of concentric pipes, and compressed air sent down the innermost pipe sends the emulsion to the surface, where the sulfur can be gathered due to its insolubility. The process must be carefully monitored, as sulfur dioxide and hydrogen sulfide are serious pollutants, with the former having the potential to lead to acid rain.

4.3.6/4.3.7/4.3.8 Outline the steps and conditions necessary for the industrial production of H_2SO_4 from its raw materials (including 'Describe the reaction conditions necessary for the production of SO_2 and SO_3 ' and 'Apply the relationship between rates of reaction and equilibrium conditions to the production of SO_2 and SO_3 ')

These three dotpoints are better approached as one, as they all lead to the production of sulfuric acid. Make sure you break down the production process into three parts, as you'll find that the contact process (As it is called) makes a lot more sense when you can picture it as three big steps. Also keep in mind equilibrium considerations (Le Chatelier's Principle and rates of reaction).

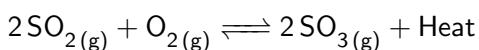
The current method used to produce sulfuric acid was discovered in the early 1800s in a process known as the contact process. Under this process, three main steps can be identified: The production of sulfur dioxide, sulfur trioxide, and sulfuric acid. You will note how these tie into one another as the process progresses.

The production and preparation of sulfur dioxide



This step is simply a combustion reaction, where liquid sulfur is sprayed into an excess of dry air. These conditions are used to ensure a maximum output of sulfur dioxide while taking care to avoid producing sulfuric acid through the reaction of sulfur dioxide with water. Alternatively, water can be used to wash the gas, and then sulfuric acid is used to dry the gas once more. When produced, the sulfur dioxide reaches temperatures of 1000C. However, this must be cooled down to a temperature between 400-450C for the next stage of the process.

The catalytic conversion of sulfur dioxide and oxygen into sulfur trioxide

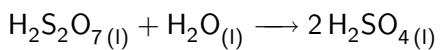
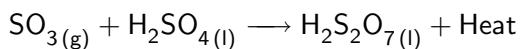


The sulfur dioxide previously produced is combined with oxygen as it is passed over a catalyst bed at approximately 450°C , converting round 70% of the sulfur dioxide in the mixture into sulfur trioxide. The sulfur trioxide is then removed after the mixture has been cooled, after which the mixture is passed over yet another catalyst bed. This provides a yield of 99.7%, with the remaining sulfur dioxide at safe concentrations for release into the atmosphere.

The ideal conditions for this stage of the process are as follows:

- Low temperatures: The reaction is exothermic, and thus lower temperatures would promote a higher yield. This is the reason why the temperature is dropped from 1000°C to around 450°C .
- Catalyst: Catalyst beds of vanadium pentoxide are used as the lower temperatures used to promote a higher yield also promote a slower rate of reaction. Without these catalysts, the reactions is not likely to ever reach a state of equilibrium.
- Pressure: Pressures slightly above atmospheric pressure are used as higher pressure promotes higher yield in this reaction. However, the effect is minimal, thus the pressure is not raised any higher as the costs and risks of doing so outweighs any gain in yield.
- Excess oxygen: A small excess of oxygen is also used to promote a higher yield of sulfur trioxide.

The conversion of sulfur trioxide into sulfuric acid



Although sulfur trioxide can theoretically be added directly to water to form sulfuric acid, this is undesirable as the reaction is so exothermic that acid vapour is formed rather than a liquid.

Thus, to get around this problem the sulfur trioxide is added to existing sulfuric acid, producing $\text{H}_2\text{S}_2\text{O}_7$, a substance known as oleum. The oleum is then reacted with water, producing sulfuric acid.

Remember- Liquid sulfur is first sprayed into an excess of dry air to form sulfur dioxide, which is added to oxygen and passed over catalyst beds of vanadium pentoxide twice to form sulfur trioxide. The sulfur trioxide is then added to sulfuric acid to form oleum, which is reacted with water to produce more sulfuric acid.

In the production of sulfur trioxide, lower temperatures and high pressure promotes increased yield. However, the rate of reaction and safety and cost must be given consideration. As such, a compromising $400\text{-}450^\circ\text{C}$ is used with a vanadium pentoxide catalyst and a pressure of a little over atmospheric pressure.

4.3.9 Gather, process and present information from secondary sources to describe the steps and chemistry involved in the industrial production of H_2SO_4 and use available evidence to analyse the process to predict ways in which the output of sulfuric acid can be maximised

This dotpoint has already been covered to a large extent in dotpoint 4.3.8 on page 124. As a note, I will mention that any question where you are required to note how to maximise yield is essentially a practice in Le Chatelier's Principle. As such, go through temperature and pressure. However, with these two, note that a low temperature reduces the rate of reaction, so for exothermic reactions a catalyst is likely to be required. In addition, higher pressure will not influence yield much if the imbalance in the moles of gas on the reactants side does not vary much from the moles of gas in the products side (In the case of the production of sulfur trioxide in the contact process, the ratio is only 3:2). As such, the cost and safety concerns of higher pressure might be given a larger weighting. Lastly, consider what products can be removed and what reactants can be used in excess so as to further push the equilibrium to the right. In the case of producing sulfur trioxide, sulfur trioxide can be removed and excess oxygen can be used.

4.3.10 Perform first-hand investigations to observe the reactions of sulfuric acid acting as: An oxidising agent, and a dehydrating agent

For this dotpoint there is a very simple experiment which illustrates the use of sulfuric acid as both an oxidising agent and as a dehydrating agent.

Materials:

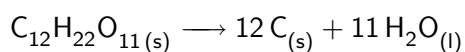
- One 250mL beaker
- Sucrose
- Concentrated sulfuric acid

Procedure:

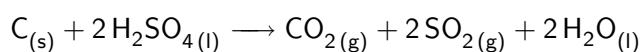
1. Pour sucrose into an empty beaker, filling it up to one third of capacity.
2. Place the beaker in a fume cupboard.
3. Pour 25mL of concentrated sulfuric acid into the beaker, and step away.

Expected Results:

It may take a few seconds for the beaker to show any change, but once it does the result will be rather quickly achieved. The sucrose should be replaced by a much larger carbon structure, which demonstrates sulfuric acid agent as a dehydrating agent:

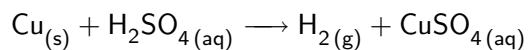


Next, the carbon is oxidised, forming carbon dioxide, sulfur dioxide and water. These fumes are dangerous in quantity, so make sure the experiment is in fact conducted in a fume cupboard.

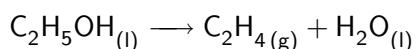


4.3.11 Describe, using examples, the reactions of sulfuric acid acting as: An oxidising agent, and a dehydrating agent

Sulfuric acid acts as an oxidising agent when a metal is added to sulfuric acid. In the example below, the copper metal is oxidised, losing two electrons and gaining a 2+ charge.



The use of sulfuric acid as a dehydrating agent is evident when examining its use during esterification, and during the dehydration of ethanol, the latter of which is shown below. The concentrated sulfuric acid is not shown as it is in fact a catalyst in the reaction.



Remember- Oxidisation Is Loss, Reduction Is Gain (OILRIG). Sulfuric acid acts as both an oxidising agent and as a dehydrating agent.

4.4 Sodium Hydroxide

4.4.1 Explain the difference between galvanic cells and electrolytic cells in terms of energy requirements

The difference between galvanic cells and electrolytic cells causes endless grief for students if they attempt to think of it in terms of positive and negative anodes and cathodes. Rather, keep thinking in terms of AnOx (Anode Oxidation).

In galvanic cells, a spontaneous chemical reaction converts chemical energy into electrical energy, whereas in an electrolytic cell, electrical energy is converted into chemical energy so as to bring about a chemical reaction.

In both the galvanic and electrolytic cells, oxidation still occurs at the anode and reduction occurs at the cathode.

Remember- A galvanic cell converts chemical energy into electrical energy. An electrolytic cell converts electrical energy into chemical energy.

4.4.2 Identify data, plan and perform a first-hand investigation to identify the products of the electrolysis of sodium chloride

This experiment is relatively simple, and simply confirms the fact that sodium ions do not reduce in the production of sodium hydroxide as readily as water. A concentrated sodium chloride solution is used as a molten sodium chloride solution is not as practical.

Procedure:

1. Prepare one beaker of dilute sodium chloride solution and one beaker of concentrated sodium chloride solution.
2. Set up two electrolytic cells using the two beakers. A current should be run through the beakers on at a time.
3. Place a test tube at both electrodes, taking care to trap any gas formed during the reaction.
4. Place a pH meter at both the electrodes of the dilute sodium chloride solution.
5. Light a splint, gently blow it out, and place near each of two electrodes in the dilute sodium chloride solution beaker in quick succession.
6. Repeat steps 3 to 5 for the beaker containing the concentrated sodium chloride solution.
7. Waft (Ideally, conduct the experiment in a fume cupboard) any gases formed during the experiment carefully towards yourself to identify its contents.
8. Place a piece of litmus paper each at both electrodes of the concentrated sodium chloride solution.

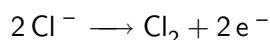
Expected Results:

- The pH was below 7 at the anode, and above 7 at the cathode. This indicates the self-ionisation of water brought about by the unresponsiveness of sodium ions to reduction where water is present.
- A re-lighted splint indicates the presence of oxygen.
- A popping sound produced when gas caught in a test tube is ignited indicates the presence of hydrogen gas.
- Chlorine gas has a distinct odour which is immediately identifiable.
- The bleaching of litmus paper at the anode of the concentrated sodium chloride solution confirms the presence of chlorine gas.

4.4.3 Analyse information from secondary sources to predict and explain the different products of the electrolysis of aqueous and molten sodium chloride

In an aqueous sodium chloride solution:

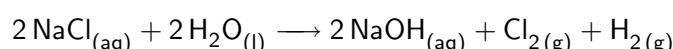
At the anode:



At the cathode:



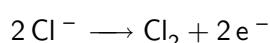
Full equation:



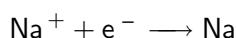
The sodium ion is more stable than water, and as such, water is reduced instead.

In a molten sodium chloride solution:

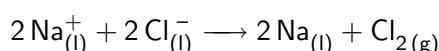
At the anode:



At the cathode:



Full equation:



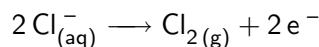
Although water is less stable than a sodium ion and thus more likely to be reduced, there is no water in molten sodium chloride! As such, the sodium ions must be reduced.

Remember- Where water exists, it will be reduced more readily than a sodium ion. Where no water exists, the sodium ions must be reduced. Chlorine ions are always oxidised in the formation of sodium hydroxide.

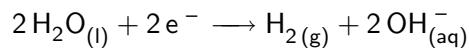
4.4.4 Outline the steps in the industrial production of sodium hydroxide from sodium chloride solution and describe the reaction in terms of net ionic and full formulae equations

In the industrial production of sodium hydroxide, a solution known as brine (really just concentrated NaCl solution) is electrolysed.

At the anode, chlorine ions are oxidised (and therefore lose electrons becoming more positive) to form chlorine gas.

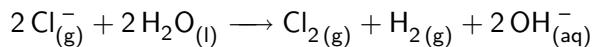


At the cathode, sodium ions are not reduced as would be expected. Rather, because the sodium ions are simply too stable to undergo this process, the water in the solution is reduced instead, forming hydroxide ions and hydrogen gas.

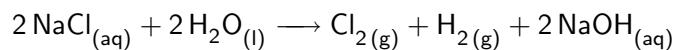


Through these processes, sodium hydroxide is formed.

Net ionic equation:



Full equation:



Remember- Chlorine ions are oxidised at the anode and water, not sodium ions, are reduced at the cathode to form sodium hydroxide from brine.

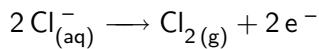
4.4.5 Distinguish between the three electrolysis methods used to extract sodium hydroxide: The mercury process, diaphragm process, and membrane process by describing each process and analysing the technical and environmental difficulties involved in each process

After examining the general gist of how electrolysis occurs and then how sodium hydroxide is produced, this dotpoint focuses specifically on three of the more common processes by which these reactions are brought about. Take the time to understand each of these equations, as they may prove important in your exams.

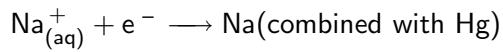
Mercury Process

In this process, the anode is a segmented titanium plate and the cathode is mercury.

At the anode, chlorine ions are oxidised to produce chlorine gas.

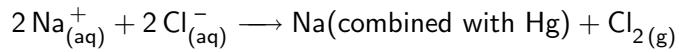


At the cathode, sodium ions are reduced to produce sodium, which forms a mixture with the mercury in the cell.

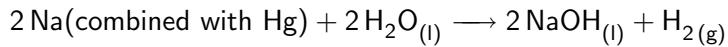


This mixture is frequently referred to as an 'amalgam', which literally means an alloy of mercury and another metal, or metals.

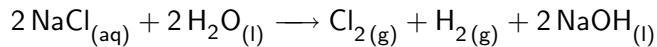
Net ionic equation:



The amalgam of sodium and mercury then flows into the next area of the cell, where the sodium and water combine to form sodium hydroxide

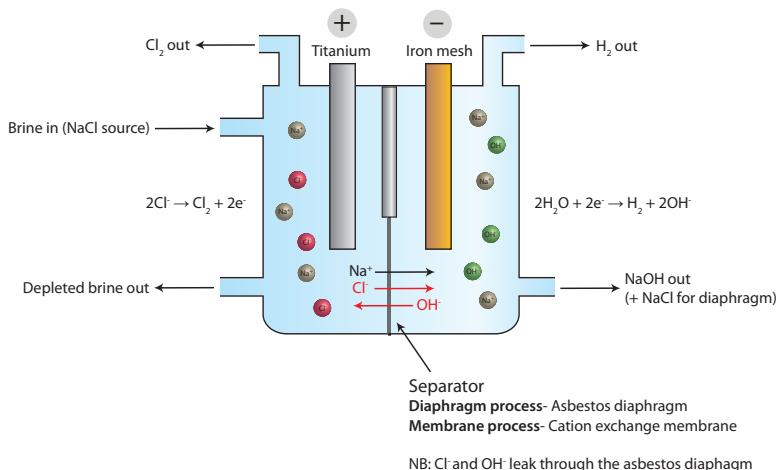


Full equation:



Difficulties:

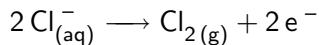
- Technical: Although achieving a high degree of purity, this process is extremely costly, particularly when taking into account the cost of replacing the mercury lost in the process.
- Environmental: Mercury, if mishandled, clearly poses both environmental and general safety problems. Mercury can enter the food chain and continue its journey through organisms through a process known as bioaccumulation.



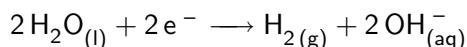
Diaphragm Process

In this process, the anode is once again titanium, yet the cathode is an iron mesh.

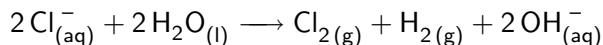
At the anode, brine enters the cell from the side closest to the anode, where the chlorine ions are attracted to the anode (which is positive in an electrolytic reaction), where they become oxidised to form chlorine gas.



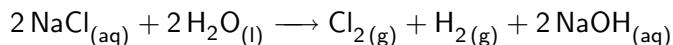
At the cathode, sodium ions are drawn towards the cathode (which is negative in an electrolytic reaction), travelling through the asbestos diaphragm. However, given its relatively stable nature, these sodium ions are not reduced. Instead, water is reduced to form hydroxide ions and hydrogen gas.



Net ionic equation:



Full equation:



Given the soluble nature of sodium hydroxide, water can be removed to crystallise out any remaining chlorine ions. Hydrogen and chlorine gas are tapped off separately.

Difficulties:

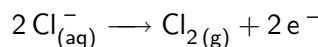
- Technical: Purity is affected as it is extremely difficult to remove all traces of chlorine ions from the sodium hydroxide product. In contrast, in the mercury process, chlorine ions are separated before the sodium hydroxide is formed, therefore resulting in a product that is still relatively pure.

- Technical: If the hydroxide ions come into contact with chlorine ions, hydrochlorous acid (HOCl) may be formed. As such, care must be taken to prevent this.
- Environmental: Asbestos is equally as problematic, if perhaps more controversial than, mercury. If asbestos fibers come into contact with humans, mesothelioma may result amongst other diseases.
- Environmental: One potential product, the hypochlorite ion (ClO^-), is an extremely potent oxidant which can have devastating repercussions upon the environment.

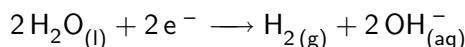
Membrane Process

This process is virtually identical to the diaphragm process. The one distinguishing feature is that where the diaphragm process makes use of an asbestos diaphragm, the membrane process relies upon a polytetrafluoroethylene (PTFE) membrane.

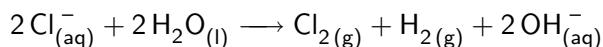
At the anode:



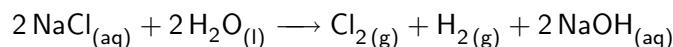
At the cathode:



Net ionic equation:



Full equation:



Difficulties:

- Technical: As with the mercury process, the membrane process is rather costly given its reliance upon membranes formed by expensive petrochemical derivatives.
- Environmental: The petrochemicals used to produce the membrane can have an adverse effect upon the environment by way of the greenhouse effect.

Remember- While the cathode half-equations may change, the reaction at the anode is always the same: Chlorine ions are oxidised to form chlorine gas.

4.5 Saponification

4.5.1 Perform a first-hand investigation to carry out saponification and test the product

Procedure:

Saponification:

1. Place 5g of NaOH pellets in a 100mL beaker.
2. Add 30mL of water.
3. Add 5mL of coconut oil.
4. Using a Bunsen burner, bring the solution to a boil, taking care to stir constantly so as to provide a uniform temperature.
5. Allow the solution to cool once the layer of oil has fully disappeared.
6. Add 10g of NaCl and bring the solution once more to a boil, again taking care to stir constantly.
7. Allow the solution to cool, at which point lumps of soap should appear.
8. Decant the solution taking care to keep the product.
9. Flush the beaker with a NaCl solution 2-3 more times, decanting after each successive flush.
10. Allow the soap to dry in the fume cupboard, upon filter paper.

Testing:

1. Take six test tubes, filling two with the soap, two with synthetic commercial detergent, and leaving the remaining two untouched.
2. Fill all six test tubes halfway with water, and a third as much of oil.
3. Stopper all six test tubes, and shake each one vigorously, taking care to keep the contents within the test tube.
4. Record your observations for each test tube five, ten, and fifteen minutes after shaking. In particular, note whether or not the liquids have formed an emulsion or separate layers, and the height of foam each time.

Expected results:

- The soap and detergent observations are likely to show varying amounts of foam depending upon the length of shaking and the detergent used, as well as the purity of the soap. However, in general, you will find that the detergent produces the most foam.
- The soap and detergent observations result in an emulsion of water and oil, whereas the test tube of simply water and oil is likely to seemingly form an emulsion at first, but settle into separate layers relatively quickly

4.5.2 Describe saponification as the conversion in basic solution of fats and oils to glycerol and salts of fatty acids

Saponification is the hydrolysis of an ester in an alkaline environment, where fat or oil is combined with hydroxide ions to form an alcohol and a carboxylate ion (COO^-).

Remember- In particular, saponification is the process of converting fats and oils in sodium hydroxide into glycerol (Propane-1,2,3-triol) and the salts of fatty acids.

4.5.3 Gather, process and present information from secondary sources to identify a range of fats and oils used for soap-making

This dotpoint is best served by learning around three different oils and two different fats.

Oils:

- Avocado Oil- Primarily used in cosmetics.
- Castor Oil- Produces a soap which is relatively mild.
- Olive Oil- Produces a durable and hard soap.

Fats:

- Lard- Derived from pig fat and produces a mild soap which is used for laundry.
- Tallow- Derived from the solid fat of beef cattle and produces soap which is yellow in colour.

4.5.4 Describe the conditions under which saponification can be performed in the school laboratory and compare these with industrial preparation of soap

In the school laboratory, esters are boiled in a solution of sodium hydroxide to produce a homogenous mixture. Given the volatility of the substances being handled- in particular the ester- this process must take place in reflux. Given that alkanoic acids have higher melting and boiling points than their alkanol counterparts- a characteristic attributable to the stronger dispersion forces of a longer carbon chain- several options are then possible:

- If the carboxylate ion has more than four carbon atoms, its solubility is likely to be relatively low. As such, it may be precipitated out in its acidic form upon being acidified, and then filtered off.
- If the carboxylate ion has an extremely low degree of solubility, it may simply be removed upon being cooled with ice.
- If the carboxylate ion has a boiling point which is relatively different, distillation may be used to separate the components.

The key characteristics of this process are:

- A water bath and hotplate are used.
- The full reaction can be brought about within an hour.
- Refluxing is employed.
- The product is relatively impure, with quantities of sodium hydroxide present.
- Process is dependent upon different boiling and melting points.

In the industrial preparation of soap, vegetable oils and fats are heated to high temperatures in a 30% sodium hydroxide solution. The product has all surface fat skimmed upon cooling, and the remaining mixture is mixed with brine. By increasing the concentration of ions present in the mixture, the soap is precipitated out in a process known quite simply as 'salting out'. Colouring and perfumes are then combined with the product which is then packaged by being cut and pressed into a mold. The brine solution is combined with acid to neutralise any excess sodium hydroxide, and then distilled to produce the glycerol.

The key characteristics of this process are:

- Significantly higher temperatures are used.
- Although the process may be completed in a day, the reactions can take several days, if not weeks to go to completion.
- The fat and/or oil are not as volatile, resulting in no need for refluxing.
- Any remaining sodium hydroxide can be neutralised and the remaining water distilled, leaving a relatively pure product.
- Process is dependant upon the 'salting out' method.

4.5.5 Perform a first-hand investigation to gather information and describe the properties of a named emulsion and relate these properties to its uses

Many different emulsions can be found in everyday situations. For the sake of simplicity, this study guide will simply describe the emulsion of French dressing.

French dressing is a common salad dressing generally containing vinaigrette among other ingredients. When combined with small amounts of mustard powder and shaken vigorously, a fairly basic emulsion is formed.

French dressing can be described as a water-oil emulsion which is fairly evenly dispersed. When coupled with commercial emulsifiers, this mixture can remain as an emulsion for extended periods of time, resisting any attempt to separate into two immiscible layers.

4.5.6 Perform a first-hand investigation to demonstrate the effect of soap as an emulsifier

For your ease, this study guide uses the same procedure for both this dotpoint, and for the testing component of dotpoint 4.5.1 on page 135. The steps are repeated again for your convenience.

Procedure:

1. Take six test tubes, filling two with soap, two with synthetic commercial detergent, and leaving the remaining two untouched.
2. Fill all six test tubes halfway with water, and a third as much of oil.
3. Stopper all six test tubes, and shake each one vigorously, taking care to keep the contents within the test tube.
4. Record your observations for each test tube five, ten, and fifteen minutes after shaking. In particular, note whether or not the liquids have formed an emulsion or separate layers, and the height of foam each time.

Expected results:

- The soap and detergent observations are likely to show varying amounts of foam depending upon the length of shaking and the detergent used, as well as the purity of the soap. However, in general, you will find that the detergent produces the most foam.
- The soap and detergent observations result in an emulsion of water and oil, whereas the test tube of simply water and oil is likely to seemingly form an emulsion at first, but settle into separate layers relatively quickly

4.5.7 Account for the cleaning action of soap by describing its structure

Soap is an effective surfactant, or surface active agent, due to its unique tadpole-like structure which allows for a two-pronged cleaning effect. The anionic 'head' is described as hydrophilic, or water-loving, whilst the hydrocarbon 'tail' is said to be hydrophobic, or water-fearing.

By working together, soap is able to act as a bridge in carrying off particles as the 'tail' attaches itself to oil and the anionic 'head' dissolves in the water. By effectively encircling these oil particles, the particle now carries a negative charge, and as such is unable to resettle on the surface as the particles repel one another.

Particles that are not oily can still be removed as the hydrophilic 'head' dissolves in the water and the hydrophobic 'tail' sticks up into the air, thereby breaking surface tension. This effectively allows particles to become more wet, and thus wash off.



Remember- The anionic head and hydrophobic tail work together to remove both oily and non-oily particles. This is done by circling and making the oily particle negatively charged, as well as by lowering the surface tension of the water.

4.5.8 Explain that soap, water and oil together form an emulsion with the soap acting as an emulsifier

An emulsion can be defined as a dispersion of one liquid in another in the form of droplets. While oil and water are immiscible- that is, they do not mix when combined- when soap is added they form an emulsion as the oil breaks up into small droplets within the water. This occurs as the hydrophobic 'tails' of soap particles attach themselves to oil particles and the hydrophilic anionic 'heads' of soap particles dissolve in water, bridging the oil particle and the water, and dispersing the oil particles throughout the mixture.

Remember- An emulsifier is a substance that produces an emulsion from two or more otherwise immiscible solutions. Soap acts as an emulsifier between oil and water due to its tadpole-like structure.

4.5.9 Distinguish between anionic, cationic and non-ionic synthetic detergents in terms of: Chemical composition, and uses

Anionic detergents have a structure similar to that of soap, with a long hydrocarbon tail and an anionic, sulfonate 'head' (OSO_2O). These surfactants are generally more effective as cleaning agents than soap, and as such are used in dishwashing detergents as well as laundry detergents. Generally, anionic detergents produce a degree of foam. However, this foam has no impact- both beneficial and detrimental- upon the performance of the synthetic detergent.

Cationic surfactants are derivatives of the ammonium ion in which the hydrogen atoms are replaced by alkyl groups. These surfactants can be formed by two long alkyl group chains in between ten and twenty carbon atoms in length, and two to three methyl groups with a charged nitrogen 'head'. Cationic detergents are particularly useful in the cleaning of fabrics, textiles, and even hair, as they are absorbed into the material, reducing friction and static charges. Another use includes both domestic and commercial cleaning agents for plastic, as well as antiseptics and biocides.

Non-ionic detergents similarly have a long hydrocarbon tail, but the opposite end is a long polar oxygen component ending in an alcohol group. The presence of the ethoxy groups ($\text{CH}_2\text{CH}_2\text{O}$) helps in producing this polarity when combined with the hydrogen end. Non-ionic detergents find much use in cosmetics, paints and adhesives, as they tend to produce far less foam than their anionic counterparts.

4.5.10 Distinguish between soaps and synthetic detergents in terms of: The structure of the molecule, chemical composition, and its effect in hard water

The Structure of the Molecule

Soaps may be distinguished from synthetic detergents by way of their molecular structure, as where soaps have a hydrocarbon tail and an ionic or polar 'head' which is usually anionic, synthetic detergents, despite having a similar structure, can have an anionic, cationic or non-ionic 'head'.

Chemical Composition

Where soaps are made of metal (usually sodium or potassium) salts of long alkanoic fatty acid chains, synthetic detergents consist of hydrocarbons and either a sulfonate end (anionic), nitrogen end (cationic), or ethoxy group end (non-anionic).

Effect in Hard Water

Hard water is simply water high in calcium and/or magnesium cations. When soap comes into contact with hard water, it creates 'scum', a greyish precipitate which limits the cleaning ability of the soap.

In contrast, synthetic detergents do not form precipitates with hard water, with cationic and non-ionic surfactants experiencing no effect at all. However, although synthetic detergents do not react with hard water, anionic surfactants may form soluble complexes with the calcium and magnesium cations, effectively neutralising the anionic surfactant and limiting the cleaning potential of the synthetic detergent.

4.5.11 Solve problems and use available evidence to discuss, using examples, the environmental impacts of the use of soaps and detergents

Early detergents were branched anionic alkylbenzene sulfonates chains which had poor biodegradability, causing excessive foaming in our waterways. Newer detergents have solved this issue, but have brought new concerns as well.

In order to combat hard water, detergents often contain 'builders' such as sodium tripolyphosphate ($\text{Na}_5\text{P}_3\text{O}_{10}$). These work to reduce the hardness of water while simultaneously increasing pH to the optimum level at which the surfactant operates. However, phosphates are a key reagent in the eutrophication of our waterways, and as such, the discharge of detergents has led to high algal bloom rates causing the death of marine ecosystems.

Cationic synthetic detergents have also proved problematic as they act as biocides, causing much damage if the runoff continues to a sewage treatment plant, as these plants rely upon bacteria to decompose the sewerage- bacteria which are effectively killed off.

In contrast, soaps have had very little impact upon the environment, as they readily break down into the harmless components of carbon dioxide and water.

4.6 The Solvay Process

4.6.1 Identify the raw materials used in the Solvay process and name the products

Before beginning this topic, it is useful to outline what exactly the Solvay process is. Defined simply, it is the major means of production of sodium carbonate (Na_2CO_3) in the modern industrial world. Sodium carbonate is also referred to as soda ash from time to time.

In the Solvay process, ammonia, calcium carbonate and sodium chloride are reacted in various stages to produce calcium chloride and sodium carbonate. This process is not brought about spontaneously, and has multiple stages of reactions in which the raw materials are introduced.

4.6.2 Describe the uses of sodium carbonate

As this dotpoint is virtually endless, for the purposes of any possible questions three major uses of sodium carbonate will be briefly elaborated on, with a few more uses listed.

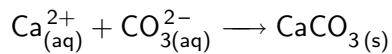
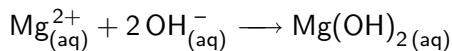
- Making glass- Sodium carbonate is combined with silica (as well as other ingredients specific to the intended purpose of the glass) in order to produce glass.
- Regulating pH- Sodium carbonate is used in order to regulate pH in many different settings. Examples include something as simple as a neutralising agent in school laboratories, as well as cooking.
- Softening water- Sodium carbonate is used in order to soften hard water- water containing calcium and/or magnesium ions. This is particularly useful in laundry products.
- Other applications include: Making paper, a primary standard in titrations, producing sodium hydrogen carbonate, and reducing sulfur dioxide emissions from industrial processes.

4.6.3 Identify, given a flow chart, the sequence of steps used in the Solvay process and describe the chemistry involved in: Brine purification, hydrogen carbonate formation, formation of sodium carbonate, and ammonia recovery

Brine Purification

Brine is first obtained from a variety of sources, including underground brine deposits, sea water evaporation projects, or simply flushing underground salt deposits. Sodium carbonate is added to this solution to precipitate out any calcium and sodium hydroxide is added to precipitate out any magnesium.

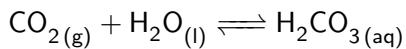
The precipitates then undergo flocculation before being filtered off upon settling.



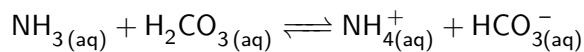
A 30% pure solution of sodium chloride is sufficient to begin the Solvay process.

Hydrogen Carbonate Formation

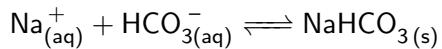
Carbon dioxide gas is first bubbled through brine which is saturated with the base ammonia to produce a weak carbonic acid.



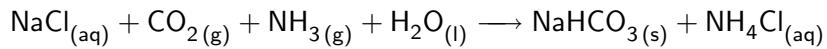
The carbonic acid then reacts with the weak base ammonia to form ammonium and bicarbonate ions.



Sodium ions then combine with bicarbonate ions to form sodium hydrogen carbonate.



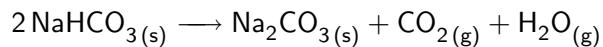
The full equation is thus:



The sodium hydrogen carbonate produced is then filtered off.

Formation of Sodium Carbonate

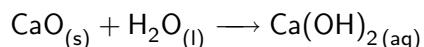
Sodium hydrogen carbonate is then decomposed into its baser components of sodium carbonate, carbon dioxide gas and water vapour by introducing heat.



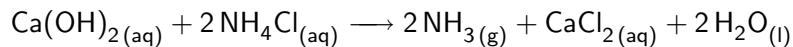
The sodium carbonate is then washed and dried, and the remaining carbon dioxide recycled for re-use in the Solvay process, along with new carbon dioxide gas produced by heating calcium carbonate (limestone).

**Ammonia Recovery**

The calcium oxide produced by heating limestone is then combined with water to produce calcium hydroxide.



This solution of calcium hydroxide is combined with ammonium chloride solution and heated to recover the valuable ammonia, as well as calcium chloride and water.



4.6.4 Discuss environmental issues associated with the Solvay process and explain how these issues are addressed**Thermal pollution**

The Solvay process is overall an exothermic process, producing considerable quantities of heat which are released into the environment. Plants near water may discharge this heat into water sources, damaging ecosystems, yet the biggest concern is inland processing plants. This concern arises as although thermal pollution may prove damaging, in general the heat is fairly well dissipated if enough water is present. In contrast, where no water is present, these plants must large quantities of heat must explore (often costly) methods of safely releasing this excess heat. Examples of such methods include the construction of pools designed specifically for the purpose of cooling ejected hot air.

Release of Ammonia

Ammonia is an air pollutant which poses a possible health threat, particularly to marine ecosystems. Well-designed monitoring systems to combat this problem are a costly, yet relatively simple solution to this problem.

Disposal of Calcium Chloride

Calcium chloride is a product which must be disposed of safely, yet cannot be done so in bulk for fear of damaging the environment. Although other countries find a use of this product for the purpose of removing the ice from streets, this use is negligible in Australia. Disposal of calcium chloride in the water sources closely connected to the ocean is an option as the sea water can effectively neutralise any residual alkalis, yet inland plants must again seek alternate means. One possibility is the burial of dried calcium chloride, yet attention must also be paid to the cost of this procedure, as well as the leaching of ions into nearby water sources.

Obtaining Calcium Carbonate and Brine

In obtaining the calcium carbonate (limestone) and brine necessary for the Solvay process, the environment can be adversely impacted by mining projects. As such, proper research and procedures (Such as an Environmental Impact Statement in some countries) should be followed to minimise long-term harm to the environment.

Disposal of Calcium Hydroxide

The calcium hydroxide remaining from the process must be safely disposed of. This is a relatively simple process, as the substance may be neutralised and disposed of relatively easily.

4.6.5 Perform a first-hand investigation to assess risk factors and then carry out a chemical step involved in the Solvay process, identifying any difficulties associated with the laboratory modelling of the step

Any step in the Solvay process can be chosen for this dotpoint. One of the simplest is perhaps the decomposition of sodium hydrogen carbonate into its base components. Keep in mind that this reaction, when done in a school lab, is nothing like the industrial process which relies on much higher temperatures. However, the chemistry is still relatively similar, if not the same.

Procedure:

1. Place a small amount of NaHCO_3 on a piece of filter paper and record its net weight.
2. Carefully pour the NaHCO_3 into a test tube after weighing the test tube separately.
3. Stopper the test tube with a rubber tube connected to the stopper.
4. Clamp the test tube to a retort stand, and light a Bunsen burner below the test tube such the tip of the flame just touches the base of the test tube. The rubber tube should trail down into a beaker of lime water.
5. Record all observations over a five minute period.
6. Allow the test tube to cool, and weigh.

Expected results:

- Water vapour should be seen to form around the mouth of the test tube.
- The cloudiness of the lime water indicates the presence of carbon dioxide gas.
- The test tube would have decreased in mass, indicating the loss of water vapour and carbon dioxide gas.

Also note that although we chose to identify the presence of carbon dioxide gas in this experiment, realistically the carbon dioxide gas would be carefully monitored and captured for reuse in the Solvay process. Also, take the time to consider safety precautions for this experiment such as the use of protective eyewear, laboratory coats and gloves.

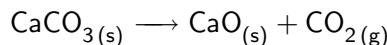
4.6.6 Process information to solve problems and quantitatively analyse the relative quantities of reactants and products in each step of the process

This dotpoint is easy to brush aside, but is important to understand as stoichiometric calculations such as the ones involved in this dotpoint will form a vital part of the HSC Chemistry course. As such, a full line of explanation of the reasoning behind stoichiometric calculations in the Solvay process is provided below.

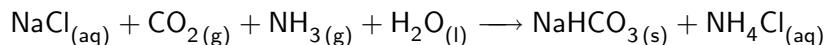
Initially, one mole of CaCO_3 is present (100.09g).

This implies:

- One mole each of calcium oxide (56.08g) and carbon dioxide gas (44.01g) are produced from one mole of CaCO_3 :



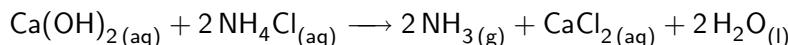
- One mole each of sodium hydrogen carbonate (84.01g) and ammonium chloride (53.49g) are produced from one mole of NaCl :



- Half a mole each of sodium carbonate (53.00g), carbon dioxide gas (22.01g) and water vapour (9.01g) are produced from one mole of NaHCO_3 :



- One mole each of ammonia (17.03g) and water (18.02g), as well as half a mole of calcium chloride (55.49g) are produced from one mole of ammonium chloride:



Remember- Select the appropriate equation, solve the number of moles for the appropriate substances, and calculate the mass of these substances as required.

4.6.7 Use available evidence to determine the criteria used to locate a chemical industry using the Solvay process as an example

In selecting the location for a chemical industry using the Solvay process, several criteria must be weighed up against one another and prioritised:

- Raw materials- Brine and limestone are two raw materials which should ideally be found close to a chemical industry using the Solvay process. Ammonia is not as problematic as it is recycled, bringing its costs down substantially.
- Labour- The cost and skill of the labour force in the area are relevant factors which must be given consideration.
- Transportation costs- The proximity of the chemical industry in relation to key locations such as suppliers and buyers must be taken into account.
- Disposal of waste- The site should ideally be located near a water source connected the ocean for the disposal of calcium chloride and excess heat.

Remember- The key concerns in choosing a location for a plant using the Solvay process are raw materials, labour, transportation costs and the disposal of waste.

Chapter 5

Extra Content

5.1 Balancing Formulae

Being able to write and balance chemical formulae quickly and accurately is integral to getting the top marks in any Chemistry exam, as you will find that a well-placed formula can support a point you are trying to make and complement any in-depth theory you wish to explore.

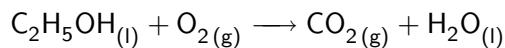
For the sake of simplicity, I have generalised the process into five basic steps:

1. Identify the reactants and products.
2. Balance the element with the smallest amount first. This is essentially a value judgment on your side, and it isn't too much of an issue which element you choose – it simply serves to simplify the process.
3. Next balance the next simplest element (ideally any elements attached to the element you just balanced). Even if you have to multiply a molecule by a factor such as a half in order to get it to balance, do so! At the very end, you can just multiply everything by two if you need to.
4. Repeat step three until all elements have been dealt with, taking care to recheck any elements which you subsequently change.
5. Count each elements on both sides (reactants and products) to double check.

To make this process clearer, two fully worked solutions are as follows.

The Complete Combustion of Ethanol

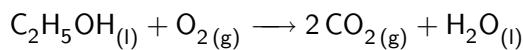
Firstly, the reactants in the combustion of ethanol are clearly ethanol, as well as oxygen. The products are water, and given that it is a complete combustion, carbon dioxide gas as well. Carbon (soot) and carbon monoxide are only reactants if incomplete combustion occurs. Write your starting equation, disregarding the fact it may not (and does not in this case) balance.



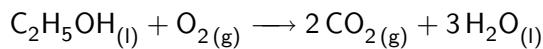
The element with the smallest amount would seem to be carbon, and as such, it may be easier to begin balancing the carbons.



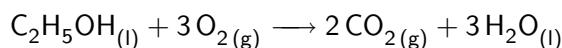
The next simplest element to balance is oxygen, as it is joined to carbon on the products side, and by itself in oxygen on the reactants side. To do this, count up how many oxygens on both sides, and play around a bit with the numbers on both sides until they balance. This usually takes a while to get the hang of it. As it happens, there are exactly three oxygens on both sides, so we can leave them.



The next element would then be hydrogen. There are six on the left, and only three on the right. Clearly we must double the water molecule, but doing so changes the oxygen we already balanced!



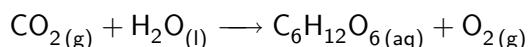
Carbon and hydrogen are thus balanced, and there are three oxygens on the left and seven oxygens on the right (a difference of four). There are thus two alternatives. We either add four ethanols, or add two oxygen gases. Noting that if we multiply the ethanol, we change carbon and hydrogen as well, clearly the best route is to add two oxygen gases.



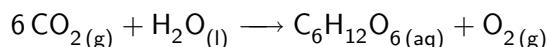
There are now two carbons on both sides, six hydrogens on both sides and seven oxygens on both sides.

Photosynthesis

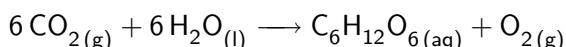
Firstly, you require carbon dioxide and water in order for photosynthesis to occur, so these are the reactants. Glucose and oxygen gas are produced, so these are the products. Unbalanced, the formula is thus:



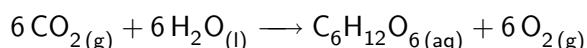
Carbon seems the simplest to balance, having a total of six atoms on both sides, unbalanced, so starting with this might be best.



Next, we can either balance oxygen or hydrogen. Now, oxygen seems easiest, but note that there is an oxygen gas molecule by itself on the products side. As such, perhaps balancing hydrogen first is best, as this oxygen gas can be used to balance the oxygens easily at the end.



We now have 18 oxygens on the left, and eight on the right. Adding five more oxygens to the right results in:



There are now six carbons, 18 oxygens and 12 hydrogens on both sides.

5.2 Common Ions

Throughout the HSC Chemistry course you will come into various ions quite often. The purpose of this list is to provide you with a list of the more common ones so as to familiarise you with them to the point where you can use their names and formulae interchangeably.

You will often find that many metals have ions with multiple valences. Although they will often be referred to as tin(II) and tin(IV) or lead(II) and lead(IV), situations may arise where to see them as stannous, stannic, plumbous, or plumbic ions. The trick to remembering which is which is that, more often than not, -ous will be the suffix for the ion with a lower valence, and ic will be the suffix for the ion with the higher valence. E.g. Cu^+ and Cu_2^+ are known as the cuprous and cupric ions respectively.

Empirical Formula	Name
CH_3COO^-	Acetate
NH_4^+	Ammonium
CO_3^{2-}	Carbonate
CrO_4^{2-}	Chromate
HCO_3^-	Hydrogen carbonate or bicarbonate
HPO_4^{2-}	Hydrogen phosphate
HSO_4^-	Hydrogen sulfate or bisulfite
HS^-	Hydrogen sulfide
OH^-	Hydroxide
ClO^-	Hypochlorite
NO_3^-	Nitrate
NO_2^-	Nitrite
O^{2-}	Oxide
MnO_4^-	Permanganate
O_2^{2-}	Peroxide
PO_4^{3-}	Phosphate
PO_3^{3-}	Phosphite
SO_4^{2-}	Sulfate
SO_3^{2-}	Sulfite
SCN^-	Thiocyanate

Chapter 6

Exam Verb Guide

6.1 HSC Exam Verbs

Throughout the HSC course you will notice a variety of verbs used to phrase questions. These verbs play a surprisingly important role in directing how the question is to be answered, and as you will find it is far easier to get the marks if you know exactly what is expected of you, it is definitely worth the effort going over exactly what these verbs mean. Write beyond the scope of the question and you will waste precious time, and write without sufficient focus and you may miss the point of the question entirely. Read the question carefully, and you might find that the extra time spent in reading and planning each question is returned by way of higher marks.

Identify

Identify is the most basic verb that you will encounter. All it requires is for you to recognise and name the required subject. However, as this is extremely simple, it is uncommon for *identify* questions to be used by themselves, but rather they will be used in conjunction with other verbs such as *explain*. Unfortunately, this is often impliedly expected from the question, and as such, it can be difficult to ascertain the precise degree of detail a question may require. For example, *Identify instruments and processes that can be used to detect radiation* requires you to name several instruments used to detect radiation. However, this question is likely to expect, at the very least, a brief description of each named process. For example, a base answer for a one-mark question may name the Cloud Chamber, and then state that this device actually consists of an air space with supersaturated water or alcohol vapour which condenses into water droplets with the passing of radiation. Your best bet is to touch on all the main parts of a question briefly at the very least so as to cover all your bases. However, do not waste time expanding upon such question if it means sacrificing other questions for the sake of a one-mark question. *Identify* is meant to be an extremely straightforward verb.

Explain

Explain essentially picks up where *identify* left off, making a subject known in detail rather than simply recognising and naming it. This might mean a variety of things in different contexts. It may require you to explain 'Why' in some questions and 'How' in others, but its intention will always be to make a certain issue clear. As a note, it often helps to define your terms before launching into an in-depth explanation of a subject. For example, *Explain the formation and effects of acid rain* would ideally begin with an explanation of what acid rain is first before leading on to how it arises largely due to the solubility of gases such as sulfur dioxide and nitrogen dioxide in the atmosphere. Its effects should then be explained by identifying the effects and explaining why they occur for example, limestone structures gradually disintegrate as the acid reacts with the calcium carbonate.

Define

If *identify* and *explain* existed on a scale, *define* would lie somewhere in between. While it doesn't require the degree of detail required by *explain*, *define* requires you to not only to recognise and name the meaning of the given subject, but also to identify its essential characteristics. For example, *Define the molar heat of combustion of a compound and calculate the value for ethanol from first-hand data* should ideally be answered with the sentence 'The molar heat of combustion is the amount of energy released in the form of heat when one mole of a substance is combusted to form products in their standard states (solid, liquid or gas) at 100 kPa and 25°C (298K)'. This not only states the meaning of the required term, but also provides its essential features.

Compare

Compare is an extremely simple verb to grasp if you can answer in your mind two questions before writing anything down:

- What similarities do the subjects have in common?
- What differences exist between the subjects in question?

For example, *Compare the properties of the oxygen allotropes O₂ and O₃ and account for them on the basis of molecular structure and bonding* provides a rich opportunity to name key differences in boiling points, structure and density between oxygen and ozone. It is of the utmost importance that you learn to link your observations together so that you demonstrate to the examiner that you are noting the characteristics of one subject in relation to one or more other subjects, and not simply rattling off unrelated comments. To do this, make use of phrases such as 'whereas', 'in contrast', and 'similarly'. Note that *contrast* questions are simply *compare* questions without any attention to similarities. Your focus in a *contrast* question is simply on the characteristics between two or more subjects which are different if not altogether opposite. Providing similarities will not give you any marks in a *contrast* question.

Discuss

This is perhaps the most common verb used as its comprehensive nature allows for the examiner to test vast areas of knowledge in depth. Broadly, it consists of two parts. The first is a simple identification of issues, and the second part requires you to provide arguments for and against the issues being discussed. Note that this is not for *or* against. You must thus take the time to first identify relevant issues and then develop opposing lines of argument. For example, *Assess the potential of ethanol as an alternative fuel and discuss the advantages and disadvantages of its use* states quite specifically that you should identify both the advantages and disadvantages of using ethanol as an alternative fuel after identifying its appropriate as an alterative fuel as the issue to be resolved.

Ethanol is renewable and cleaner-burning. However, it requires a large area of arable land, leads to large wastage of biomass, and can currently consume more energy to produce than it outputs.' Straight away we establish an argument for, and an argument against. Clearly in an exam these points would be elaborated upon, and possibly even an equation provided for the combustion of ethanol in comparison with octane.

As a note, you will do yourself a favour if you structure your response appropriately, separating arguments for and against so that they are distinguishable from one another. This makes it far easier for a marker to read your response and allocate marks accordingly.

Analyse

Analyse is an odd fusion of *identify* and *explain* while going a little bit further. You not only need to identify the main components and explain the relationship between these components and their essential features, but you also need to make use of data and draw implications and/or conclusions. For example, *Analyse the position of these non-metals in the Periodic Table* requires you to examine the Periodic Table and then make an inference as to the nature of each metal based on their location on said table. Once you identify the oxides of most elements on the right are acidic, you can infer that non-metal oxides are generally acidic. Pay particular attention to any information that is provided, and where possible, factor that into your response.

Evaluate

Evaluate is an extremely straightforward dotpoint which requires you to *discuss* an issue and by implication arguing for and against an issue you've taken time to identify and then taking an actual stance supported by your arguments. For example, *Gather, process and present information to interpret secondary data from AAS measurements and evaluate the effectiveness of this in pollution control* requires you to argue for any against the effectiveness of the use of AAS in pollution control, and then actually state whether or not you believe it should be used in this context. Where appropriate, include a real life example such as the detection of heavy metal poisoning in seafood. For example, you could argue that although AAS is an extremely costly and slow process, it is highly accurate and easy to understand, and has enormous potential to prevent health emergencies from occurring. As such, you would state that the process is effective and should be continued. The key to this question is identifying what the issue to be resolved really is, and for this, the answer lies in the question. Note that this is extremely similar to the verb *assess*. The only difference is that where *evaluate* calls for a general judgment to be made, *assess* targets a judgment made on value, quality, outcome, or results. Despite this, the difference is usually quite negligible.

General Tips

It should now be obvious that every verb has a specific direction which it seeks to take you. Your best bet in the exam is to actually take the time to consider exactly what the question wants from you on a case-by-case basis and then do your best to answer it, using the verb as guidance.

Do this, and you will find that you give your response structure, and avoid the time-wasting waffle that most students do in an attempt to fill up the provided writing space and hopefully get the marks. Conciseness and relevance are the key to maximising marks.

Chapter 7

Exam Technique

7.1 In-exam hints

There are many sources of information on how best to approach exams, but this section examines hints that are specific to the HSC Chemistry exam.

Read the question first

It seems like a logical enough statement, but it's surprisingly common for a student to entirely miss the point of a question. With criteria-based marking, this can have devastating effects on a student's mark as they go off entirely on a tangent. Studying the dotpoints above doesn't mean anything if you can't adapt your response to suit the question, removing irrelevant parts while including others that are.

Pick out key verbs

As stated in the verb guide, target the verbs in the question. They will give your response structure and direction, providing a concise response which answers the question and filtering off any other unnecessary material else you may be tempted to add.

Double check answers

This is a simple enough point, but the trick lies in how you double check. Don't just read your answer and be done with it. Forget what you've written, read the question, and then ask yourself what points you need to have covered. Only now should you go through your answer and check that you've answered all these points. Doing so provides a fresh perspective on an answer and can often help make a great answer even better.

Add equations where appropriate

Equations, equations, equations. These will support your answer and save you a great deal of explanation as the equation itself is evidence of the point you are trying to make. As such, take the time to learn how to write and balance an equation.

Use the lines as a guide as to how much to write

You will often feel that you have not been given enough space to write a suitable answer, but what you must understand is that you have been given *exactly* enough space to do so. If you repeatedly go over the lines, practice writing concisely. You are not only more likely to get higher marks, but you'll save valuable time which you can use for answering other questions.

Draw diagrams in pencil

Drawing a diagram in pen only to realise you drew something incorrectly is the most frustrating and unnecessary mistake. Draw in pencil! This provides a diagram that is easier to distinguish from writing around it while providing something a lot easier to change should you feel the need to.

When discussing equilibrium, phrase your response carefully

Too often will a phrase such as equilibrium shifts to the right' present the opportunity for much confusion in equilibria questions particularly questions relating to Le Chatelier's principle.

Instead, state that the forward reaction is favoured *as a result* of a change in the system, and that due to Le Chatelier's principle, equilibrium will attempt to shift backwards in an attempt to correct the situation. By the very nature of dynamic equilibrium, the reaction literally goes both ways. Specify exactly what is happening and you will save yourself from losing marks that you might very well have deserved.

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