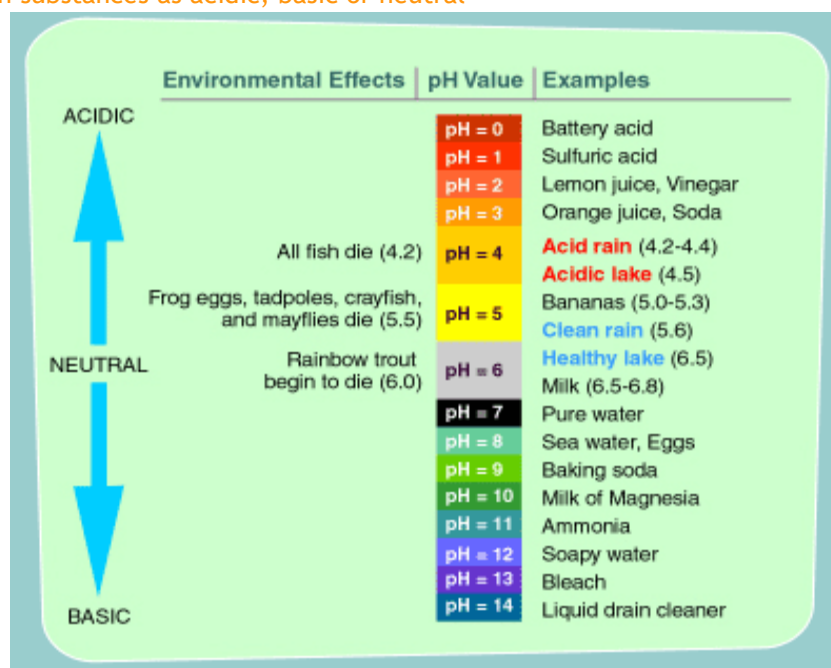


# HSC Chemistry Notes - The Acidic Environment

## 9.3 - 1. Indicators were identified with the observation that the colour of some flowers depends on soil composition

### 1. classify common substances as acidic, basic or neutral



An **acid** is a substance which in solution **produces** hydrogen ions,  $H^+$  or more strictly  $H_3O^+$ , sometimes called **hydronium ions**.

A **base** is a substance which either **contains** the oxide  $O^{2-}$  or **hydroxide ion**  $OH^-$  or which in solution **produces** the hydroxide ion.

### 2. 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

There are a range of indicators which will react and show visually different results in acids, bases or neutral substances. Indicators such as litmus, phenolphthalein, methyl orange and bromothymol blue can be used to determine the acidity or basicity of a material over a pH range. **The range is identified by a colour change in the indicator.**

### 3. identify and describe some everyday uses of indicators including the testing of soil acidity/basicity

- **Testing acidity/basicity of soils:** because some plants or flowers require acidic soils and others, including vegetables require basic pH levels in the soil composition for optimal growth
- **Testing acidity/basicity of water:** It is important to always test the pH levels in swimming pools to ensure bacteria cannot thrive. Monitoring the pH levels in local rivers are important for ecologists to maintain an ecological balance.

### 1. perform a first-hand investigation to prepare and test a natural indicator

**Red cabbage** can be used as a natural indicator and **will change according to the level of pH** of a substance. The juice from the cabbage must first be extracted by crushing the leaves to remain with the natural indicator purple juice.

In practical experiment we notice that in **acid solution** it is **red**: as the hydrogen ion concentration is decreased, the colour changes to purple and then to green: if placed in a **basic hydroxide ion solution** this dye **goes yellow**.

### 2. identify data and choose resources to gather information about the colour changes of a range of indicators



- Dissolved **non-metal** oxides are **acidic**
- Dissolved **metal** oxides are **basic**
- Aluminium, Zinc, Tin and Lead oxides are **amphoteric** (can be either basic/acidic depending on environment)
- **Noble gases** are **neutral oxides**

Common acidic, basic, amphoteric and neutral oxides:

Acidic oxides	Basic oxides	Amphoteric oxides	Neutral oxides
CO <sub>2</sub>	Na <sub>2</sub> O	ZnO	CO
NO <sub>2</sub>	CuO	Al <sub>2</sub> O <sub>3</sub>	NO
P <sub>2</sub> O <sub>3</sub>	K <sub>2</sub> O	PbO	N <sub>2</sub> O

### 3. define Le Chatelier's principle

In 1885, Le Chatelier formulated a simple generalisation for predicting *how an equilibrium will respond to a disturbance*; it is known as Le Chatelier's principle.

If a system at equilibrium is disturbed, then the system adjusts itself so as to minimise or balance out the change.

Using this principle one can determine which way the reaction goes (left or right) in order to balance.

### 4. identify factors which can affect the equilibrium in a reversible reaction

- **Pressure** is altered
- **Temperature** is changed during the reversible reaction
- **Volume**
- **Concentration** of one chemical is altered in the reaction

### 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

We are able to dissolve CO<sub>2</sub> into water to become H<sub>2</sub>CO<sub>3</sub> (carbonic acid) but is limited by Le Chatelier's principle in the respect that only certain concentrations/amounts can dissolve dependent on pressure, temperature and volume.

- CO<sub>2</sub>(g) + H<sub>2</sub>O(l) ↔ H<sub>2</sub>CO<sub>3</sub>(aq)** - If the pressure of CO<sub>2</sub> is increased by pumping more into a reaction vessel, **some will dissolve/go into the water as H<sub>2</sub>CO<sub>3</sub> to counteract the increase**. We say that the **equilibrium equation moves to the right**. This also works in reverse - if pressure of CO<sub>2</sub> is decreased then some dissolved H<sub>2</sub>CO<sub>3</sub> will decompose to CO<sub>2</sub> and thus the equilibrium equation **moves to the left**.

If were to increase the overall pressure of the system (with say nitrogen) without decreasing the volume, this would have no effect on the equilibrium since we would not be changing the pressure of CO<sub>2</sub>.

- CO<sub>2</sub>(g) ↔ CO<sub>2</sub>(aq) + heat - The reaction is **exothermic**, which means **heat is liberated as the system moves from the left to the right**. However, if we were to add heat to the CO<sub>2</sub> - H<sub>2</sub>CO<sub>3</sub> equilibrium the reaction would move to the left to counteract the addition of heat because that is the side which **absorbs the heat**. Thus abiding my Le Chatelier's principle.

As temperature increases, the solubility of carbon dioxide into water decreases.

### 6. identify natural and industrial sources of sulfur dioxide and oxides of nitrogen

	Natural Sources	Industrial Sources
SO <sub>2</sub>	Volcanoes/Decomposition of organic matter	Burning of fossil fuels/Extracting metals from sulfide ores
NO <sub>x</sub>	Lightning (nitric oxide)	Combustion of fuels (both power stations and cars)

7. describe, using equations, examples of chemical reactions which release sulfur dioxide and chemical reactions which release oxides of nitrogen

- **Sulfur Dioxide** - Coal contains *sulfur* and is impractical to extract it from coal before burning. The chemical reaction for the combustion of coal is:  $S(\text{in compounds}) + O_2(g) \rightarrow SO_2(g)$
- **Oxides of Nitrogen (nitrogen dioxide)** - Lightning, generated at high localised temperatures and pressures, oxygen and nitrogen gases combine to form *nitric oxide*:  $O_2(g) + N_2(g) \rightarrow 2NO(g)$ . This nitric oxide then combines again with existing oxygen in the atmosphere to form nitrogen dioxide:  $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$

Oxides of nitrogen include nitrogen monoxide, nitrogen dioxide etc.

8. assess the evidence which indicates increases in atmospheric concentration of oxides of sulphur and nitrogen

Direct Evidence:

- **No conclusive direct measurements** of sulfur dioxide or oxides of nitrogen over a long period of time because:
  - overall levels of  $SO_2$  and  $NO_x$  are extremely low, approximately 0.01ppm.
  - Analysis techniques sensitive enough to detect such low levels of pollutant in the air as only been available within the last 30 years
  - $SO_2$  and  $NO_x$  are washed out by rain

Indirect Evidence:

- Acid rain is shown to have increased during times of **high manufacturing activity** which release oxides of sulfur and nitrogen - shown in precipitation records
- Effects of **acid rain have become more pronounced**
  - damage to vegetation and buildings
  - **lowering of pH in lakes**
  - **deaths in Europe during the Industrial Revolution** which corresponded with extremely high levels of local emissions. With regulations for the emission of  $SO_2$  and  $NO_x$  such incidents declined markedly

Compared to the quality of air and levels of  $SO_2$  and  $NO_x$  pollutants in the air during the Industrial revolution, I do not believe that there is enough hard evidence to say that there has been a significant or noticeable increase in the atmospheric concentrations of oxides of sulfur or nitrogen. Most pressing is the fact that we do not have data on the levels of such pollutants beyond 30 years ago - indicative of insufficient evidence to back claims of an increase in atmospheric concentrations.

It may also be not be fully right to state that there has not been an increase due to the higher occurrence of acid rain, a result of absorbed  $SO_2$  and  $NO_2$  in rain. If there has been an increase in the atmospheric concentrations of either or both of  $SO_2$  or  $NO_2$  it has not been of very concern or affect on society and nature as this would be indicated through lowered pH levels and erosion.

9. calculate volumes of gases given masses of some substances in reactions, and calculate masses of substances given gaseous volumes, in reactions involving gases at  $0^\circ\text{C}$  and 100kPa or  $25^\circ\text{C}$  and 100kPa  
The volume of gases can be calculated using the data sheet information and **number of moles** and **mass**. One mole of an ideal gas at 100kPa at  $0^\circ\text{C}$  is 22.71L and at  $25^\circ\text{C}$  is 24.79L. See blue dot point 1.

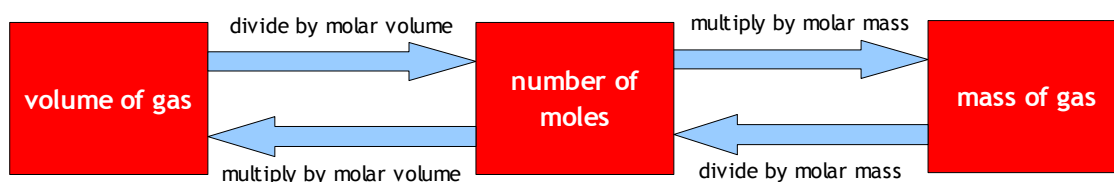
Example:

What mass of gas is present in a 515 mL flask containing carbon dioxide at  $25^\circ\text{C}$  and at a pressure of 100.0 kPa?

Using the molar volume of a gas at  $25^\circ\text{C}$  and 100.0 kPa:

$$\text{Number of moles of } CO_2 \text{ present: } \frac{0.515 \text{ L}}{24.79 \text{ mol/L}} = 0.0208 \text{ mol}$$

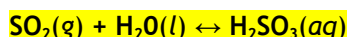
$$\text{Molar mass of } CO_2: 12.0 + 16.0 \times 2 = 44.0 \text{ g/mol} \quad \text{Mass of } CO_2: 0.0208 \times 44.0 = 0.914 \text{ g}$$



### 10. explain the formation and effects of acid rain

**Acid rain** is rain that has a **higher hydrogen ion concentration than normal** - higher than about  $10^{-5}$  mol/L. Due to the natural and human/industrial emissions of pollutant gases such as sulfur dioxide ( $\text{SO}_2$ ), carbon dioxide ( $\text{CO}_2$ ) or nitrogen dioxide ( $\text{NO}_2$ ). These **pollutants are then converted into sulfuric acid** (the main concern) which are **removed/dissolved by rain**. This leads to slightly acidic rain.

Chemical Formation - The most common and simple (reversible) reaction occurring:



We know that  $\text{H}_2\text{SO}_3(\text{aq})$  is acidic as it ionises in water given by,  $\text{H}_2\text{SO}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) \leftrightarrow \text{H}_3\text{O}^+ + \text{HSO}_3^-$

Effects of this acid rain include:

- **Increasing acidity of lakes** - detrimental effects on fish and plant life in the seas and lakes by altering the delicate balance in pH and oxygen solubility levels.
- **Erosion of Corrosion of Marble and Buildings** - this can be extremely dangerous where buildings are made from marble or stone; the acid rain gradually eats away and erodes the stone and therefore depleting the strength of building and its foundations.

1. identify data, plan and perform a first-hand investigation to de-carbonate soft drink and gather data to measure the mass changes involved and calculate the volume of gas released at  $25^\circ\text{C}$  and 100kPa

2. analyse information from secondary sources to summarise the industrial origins of sulphur dioxide and oxides of nitrogen and evaluate reasons for concern about their release into the environment

**$\text{SO}_2$ :**

- After the **Industrial Revolution** of the early 1800s there was a great **increase in emissions of sulfur dioxide** (and particulates) to the air in the growing industrial cities from burning coal and extracting metals.
- **Air quality deteriorated significantly which led to the introduction of regulations** to control emissions from factories and these caused a dramatic improvement in air quality, particularly with respect to  $\text{SO}_2$  and particulates.

**$\text{NO}_x$ :**

- Serious pollution from  $\text{NO}_x$  did not develop until the **twentieth century** as **electricity generation and use of motor cars expanded** dramatically (since it is high temperature combustion that produces  $\text{NO}$ ).
- **Photochemical smog** (ozone in the air formed from  $\text{NO}_x$  and sunlight) **first became a problem in the 1960s**. This led to the **introduction of emission controls on motor cars**.
- These introduced regulations, however at the time being to increase the quality of air, no longer are able to have the same effect compared to approximately 50-100 years ago. **The annual average concentration of  $\text{SO}_2$  and  $\text{NO}_2$  in worldwide major cities is 0.01ppm which is 10 times the amount of normal 'pure' air** - however this is still not scientifically harmful.

We can recognise and compare to data that there has been an increase of about 30% over the past 150 years in  $\text{CO}_2$  while there is a lack of data for the concentration of  $\text{SO}_2$  and  $\text{NO}_2$  before the 1950s. However the cause for concern about sulfur dioxide and oxides of nitrogen are still present.

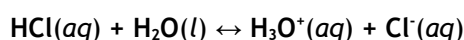
### 3. Acids occur in many foods, drinks and even within our stomachs

#### 1. define acids as proton donors and describe the ionisation of acids in water

An acid is a substance that can donate one or more protons (hydrogen ions, because a hydrogen atom that has lost an electron only has the nucleus which is made up of solely one proton). Acids ionise in water to produce hydrogen ions,  $\text{H}^+$ .



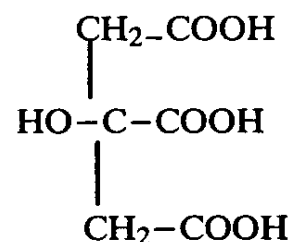
Hydrochloric acid + water  $\leftrightarrow$  hydronium ion + Chlorine ion



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

**Natural acids:**

- **Hydrochloric acid (HCl):** This is produced by glands in the lining of our stomachs. Completely ionises in  $\text{H}_2\text{O}$  and is therefore a strong acid.  $\text{HCl}(aq) + \text{H}_2\text{O}(l) \leftrightarrow \text{H}_3\text{O}^+(aq) + \text{Cl}^-(aq)$ . This is however made commercially in large quantities used to clean metals before galvanising and adjusting pool acidity.
- **Acetic (ethanoic) acid ( $\text{CH}_3\text{COOH}$ ):** Acetic acid is present in vinegar which is commonly made from wine by oxidation of the ethanol. Partially ionises in water and is therefore a weak acid.  $\text{CH}_3\text{COOH}(aq) \leftrightarrow \text{CH}_3\text{COO}^-(aq) + \text{H}^+(aq)$
- **Citric acid (2-hydroxypropane-1,2,3-tricarboxylic):** occurs in citrus fruit. It is also made industrially and widely used as a food additive (for flavour and as a preservative). The molecular formula of citric acid is  $\text{C}_6\text{H}_8\text{O}_7$ . Partially ionises in water and is therefore a weak acid.



**Synthetic acids:**

- **Sulfuric acid ( $\text{H}_2\text{SO}_4$ ):** The most factored chemical worldwide. Sulfuric acid finds its use in fertilisers, synthetic fibres and car batteries. Completely ionises in  $\text{H}_2\text{O}$  and is therefore a strong acid.  $\text{H}_2\text{SO}_4(aq) \leftrightarrow 2\text{H}^+(aq) + \text{SO}_4^{2-}(aq)$

3. describe the use of the pH scale in comparing acids and bases

The pH scale is a negative logarithmic measure of the concentration of  $\text{H}^+$  ions in solution. This scale thus determines the acidity or basicity of the substance based on the  $\text{H}^+$  concentration. The  $-\log_{10}[\text{H}^+]$ , square brackets symbolise the concentration of a certain substance.

**Note:** Acidic (0)  $\rightarrow$  Neutral (7)  $\rightarrow$  Basic (13).

4. describe acids and their solutions with the appropriate use of the terms strong, weak, concentrated and dilute

The strength of an acid is determined by its ionising ability, the strength does not change depending on its concentration.

**Strong acid:** is a solution in which *all neutral molecules have ionised to hydrogen ions*

**Weak acid:** in which only some of the acid molecules have completely ionised to hydrogen ions and thus there remains neutral molecules.

**Concentrated solution:** is one in which the total concentration of solute species is high – say, above about 5 mol/L.

**Dilute solution:** is one in which the total concentration of solute species is low – say, *less than about 2 mol/L*.

We may talk about the degree of an acid whether it is weaker or stronger than another depending on the concentration of the acid and its nature - however there is no measure of the 'strongness' of an acid. The fact of the matter is that if the acid entirely ionises then it is strong - end.

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

$\text{pH} = -\log_{10}[\text{H}^+]$ ; where  $[\text{H}^+]$  is the concentration of hydrogen ions in solution.

The pH number of a substance represents the negative power of the negative logarithmic function  $[\text{H}^+]$ . e.g. pH of 11 means  $10^{-11}$ .

As this is a log function if the pH changes by 1 then the  $[\text{H}^+]$  changes by a rate of 10. (it adds an extra 0)

6. 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

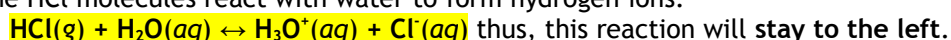
Acid	Concentration	pH	Strength	Degree of Ionisation
citric	0.1 mol/L	2.1	weak	8.00%
acetic	0.1 mol/L	2.9	weak	1.30%
hydrochloric	0.1 mol/L	1	strong	100.00%

Thus, the degree of ionisation determines an acid's strength - as either weak or strong.

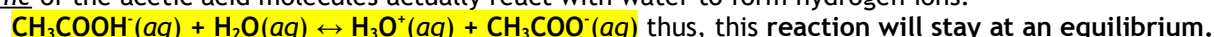
7. describe the difference between a strong and a weak acid in terms of an equilibrium between the intact molecule and its ions

We can describe the differences between strong and weak acids in terms of the equilibrium between its molecules and ions.

Hydrochloric acid is a **strong** acid. This means that when we dissolve hydrogen chloride gas in aqueous solution all the HCl molecules react with water to form hydrogen ions:



Acetic acid is a **weak** acid, because when pure covalent acetic acid is dissolved in aqueous solution, only some of the acetic acid molecules actually react with water to form hydrogen ions:



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 required the testing of some acids and bases which were at the same concentration of approximately 1 mol/L. The concentration of each acid/base had to be maintained to obtain fair and far more accurate results as the dilution of the substances could affect results obtained from the pH probe/meter and indicators.

The **acid-base indicators** can be used to **determine the acidity of an unknown substance based on the colour** that the indicator changes when mixed with the acid/base. Because the **indicators rely on colour** to show the difference in concentration of hydrogen ions (pH), they will be **affected by colour of the substance itself**. This poses issues when mixing.

However, the **pH meters/probes are far more accurate than the colour indicators** on two levels. The **meters have about a 0.01 accuracy** on their display, whereas the **acid-base indicators only have up to 1pH of accuracy** at best and require the comparison with the pH colour chart.

Also, because the **probes do not rely on the colour** of a substance they will not be affected by the colour of the acid or base - it can be colourless, red or blue etc. Despite this, extreme care must be taken when using the pH meters/probes for the most accurate results - due to their sensitivity, they must be calibrated and rinsed with water before use.

2. plan and perform a first-hand investigation to measure the pH of identical concentrations of strong and weak acids

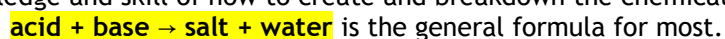
Refer to orange dot point 6 for specific examples. Testing citric, acetic and hydrochloric acid.

A pH meter/probe is required to test the pH level to a more accurate degree than the traditional colour reliant acid-base indicators such as litmus and bromothymol blue. The probe is calibrated and rinsed first and then dipped into 0.1 mol/L of solution to acquire the pH level. Record the data and test again, noting any minor changes.

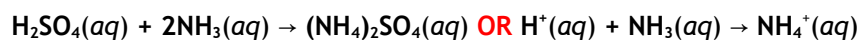
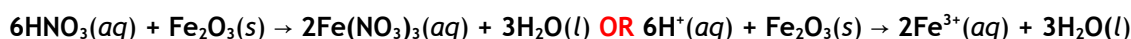
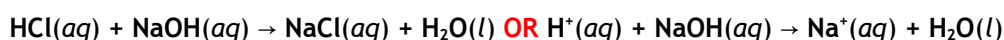
Errors or deviations in data could occur from the mixing of substances when testing for pH level. It is best to test the same acid or base three times to take the average, although if the tubes are clean and the probe rinsed each time, then errors should be avoided.

3. gather and process information from secondary sources to write ionic equations to represent the ionisation of acids

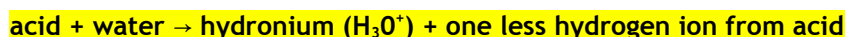
This requires the knowledge and skill of how to create and breakdown the chemical formulae.



Some equations require the addition of  $\text{H}^+$  ions or  $\text{H}_3\text{O}^+$  on the left of the reaction.



There is also the ionisation of some acids in water. In this case:



4. use available evidence to model the molecular nature of acids and simulate the ionisation of strong and weak acids

Using modelling kits and computer aided programs.

5. gather and process information from secondary sources to explain the use of acids as food additives

Acids are frequently added to food in order to:

- **Improve taste** - giving certain foods a sour or tartness taste, such as sweets or soft drinks.
- **Preserve food** - by increasing a foods acidity level, such as canned fruits or chutneys, because most bacteria have no survival rate in acidic environments.

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

Refer to orange dot point 2

Acid	Chemical formula	Naturally occurring place
Citric	$\text{C}_6\text{H}_8\text{O}_7$	Citrus fruits
Acetic	$\text{CH}_3\text{COOH}$	Vinegar
Hydrochloric	$\text{HCl}$	Glands in stomachs
Base		
Ammonia	$\text{NH}_3$	Stale urine and decomposing matter
Carbonates	$\text{XCO}_3$	Limestone, shells and coral
Metallic Oxides	$\text{Fe}_2\text{O}_3$	Extracted from respective metals

7. process information from secondary sources to calculate pH of strong acids given appropriate hydrogen ion concentrations

pH is given by:

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

We can calculate the pH level of a substance by simply putting this into the calculator, knowing the hydrogen ion concentration which would change significantly in very different environments depending on acidity or alkalinity.

4. Because of the prevalence and importance of acids, they have been used and studied for hundreds of years. Over time, the definitions of acid and base have been refined

1. outline the historical development of ideas about acids including those of:

— Lavoisier (1743 - 1798)

Lavoisier hypothesised that **all acids contained oxygen** - and in fact the word acid is derived from the Greek word meaning 'produce sour taste'. This was, however, proven wrong after the discovery that many substances containing oxygen were in fact basic.

— Davy (1778 - 1829)



Davy discovered that HCl (which we know to be acidic) did not contain oxygen and therefore proposed that **all acids contained replaceable hydrogen**. Also showing that when acid reacts with a base, a salt will be formed. → This could not explain strong/weak acids, however.

— Arrhenius (1859 - 1927)

Svante Arrhenius proposed that **acids were substances that ionised in solution to produce hydrogen ions**. Arrhenius also concluded that the degree of ionisation was an indication of the strength of the acid. (refer to orange dot point 3.4).

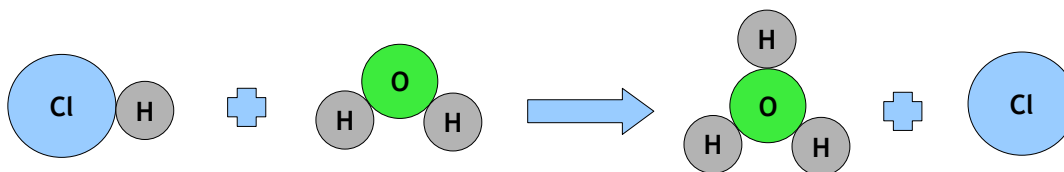
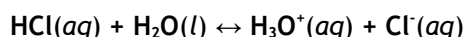
This definition is suitable but has limitations that:

- *cannot be applied outside aqueous solutions*
- *does not explain that amphiprotic substances can act as either a base or acid*

2. **outline the Brönsted-Lowry theory of acids and bases**

- An acid is a substance that, in solution, tends to give up protons (hydrogen ions): **An acid is a proton donor.**
- A **base** is a substance that tends to accept protons: **A base is a proton acceptor.**

According to the theory an acid/base reaction is one in which a proton is transferred from an acid to a base. In the Brönsted-Lowry theory of acids and bases the ionisation of HCl is represented by the equation:

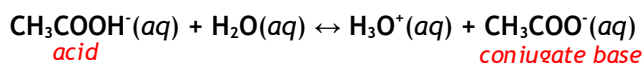


Thus water is effectively a base in the reaction.

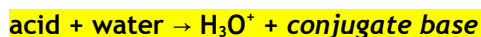
The Brönsted-Lowry definition increased our understanding further by showing that **acidity depends not just upon the structure of the substance itself but rather on its properties relative to those of the solvent or other reactant present in the solution.**

3. **describe the relationship between an acid and its conjugate base and a base and its conjugate acid**

An acid gives up a proton to form its conjugate base:

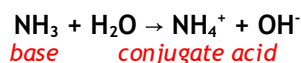


Or more generally:



The reaction product is called a conjugate *base* because in solution it tends to gain (accept) protons or hydrogen ions.

A base accepts a proton to form its conjugate acid:



Or more generally:



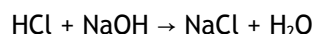
4. **identify a range of salts which form acidic, basic or neutral solutions and explain their acidic, neutral or basic nature**

Not all salts produced from an acid/base reaction are neutral pH in solution, rather this depends on the acidity and neutrality of the products or constituents.



### Weak Acid + Strong Base → Basic Salt

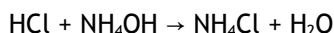
**Neutral Salt:** (sodium chloride)



**hydrochloric acid (strong acid) + sodium hydroxide (strong base) → sodium chloride + water**

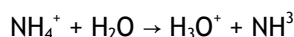
The NaCl ionises in water to create  $\text{Na}^+$  and  $\text{Cl}^-$  and there are no  $\text{H}^+$  or  $\text{OH}^-$  ions present. Thus sodium chloride (table salt) is a neutral salt.

**Acidic Salt:** (ammonium chloride)



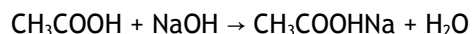
**hydrochloric acid (strong acid) + ammonium hydroxide (weak base) → ammonium chloride + water**

The  $\text{NH}_4\text{Cl}$  then goes on to ionise into  $\text{NH}_4^+$  and  $\text{Cl}^-$ . The  $\text{NH}_4^+$  will then go even further to ionise in water as follows:



Therefore the hydronium ion ( $\text{H}_3\text{O}^+$ ) is an acidic salt.

**Basic Salt:** (sodium acetate)



**acetic acid (weak acid) + sodium hydroxide (strong base) → sodium acetate + water**

The basic nature of the sodium acetate salt can be described as: when placed in water, the salt ionise to form an acetic acid ion, sodium ion and hydroxide ion. The production of the hydroxide ion indicates it's alkalinity.

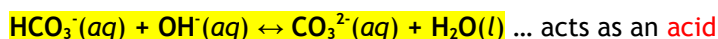
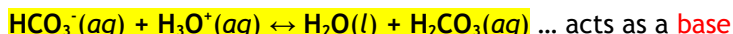
#### 5. identify conjugate acid/base pairs

This requires to identify the *active* part of the acid or base (as I call it) which will have ionised. e.g. hydrochloric acid's conjugate base is the chlorine ion.  $\text{HCl} \rightarrow \text{Cl}^-$

An acid and a substance that has one less proton than the acid make a conjugate acid/base pair. Similarly a substance which has one more proton than the base make a conjugate acid/base pair.

#### 6. identify amphiprotic substances and construct equations to describe their behaviour in acidic and basic solutions

A substance that can act as a **proton donor and proton acceptor** is known as an **amphiprotic - meaning both**. The hydrogen carbonate ion is amphiprotic, because in aqueous solution it reacts with water in two ways:



Depending on the environment which the **hydrogen carbonate** is placed in, it will react differently. If the carbonate is put in a alkaline solution, then it will act as an acid by donating a proton to the solution; however, the opposite occurs when placed in an acidic solution so that it acts as a base. When creating the two different formulae:

**acidic reaction:** include  $\text{H}_3\text{O}^+$  and then balance

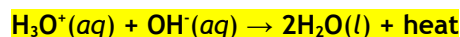
**basic reaction:** include  $\text{OH}^-$  and then balance

- The most obvious and classic amphiprotic substance is **water**. It has the ability to take a proton ( $\text{H}_3\text{O}^+$ ) away when mixed with an acid → or give up a proton ( $\text{OH}^-$ ) when mixed with a base.

**Note:** it is of interest to note that **amphiprotic substances can exist as either monoprotic, diprotic or triprotic** substances which have the ability to give up one, two or even three protons respectively. Diprotic and triprotic substances will lose their hydrogen components through a series of steps where they will reaction in solution (or as a salt) losing one hydrogen ion each time.

### 7. identify neutralisation as a proton transfer reaction which is exothermic

The basis of all neutralisation reactions is the reaction between an acid and a base which in turn produces water and heat:



Neutralisation is a **proton transfer** which is **exothermic** with a result of heat.

In all acid-base/neutralisation reactions there is always the production of a salt as well, however, the salt is not always the most important in reactions referring to proton transfer. We are interested in the salt when dealing with mass calculations however.

### 8. describe the correct technique for conducting titrations and preparation of standard solutions

**Titration** is used to **accurately measure the concentration of solutions** using practical measures and testing. Titration achieves the '*just efficient*' volume in which the concentration of solution A can be compared to solution B. To do so, we must know what the *equivalence point* is and when it occurs. Such that:

$$\frac{C_a V_a}{a} = \frac{C_b V_b}{b}$$

**Equivalence Point:** is a point in a chemical reaction at which the amounts of the two reactions are just sufficient to **cause complete consumption of both reactants**.

We detect the equivalence point of a reaction using an indicator such as Bromothymol blue or litmus etc.

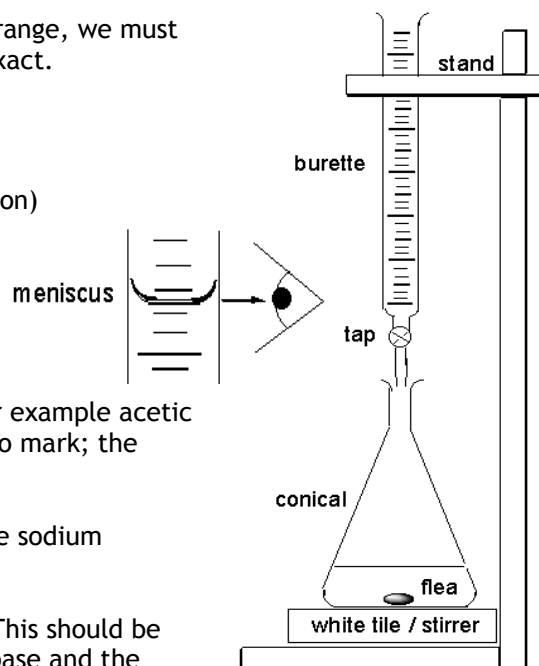
*Note:* Because indicators will change quite rapidly over a small pH range, we must take care when testing titrations as the equivalence point is very exact.

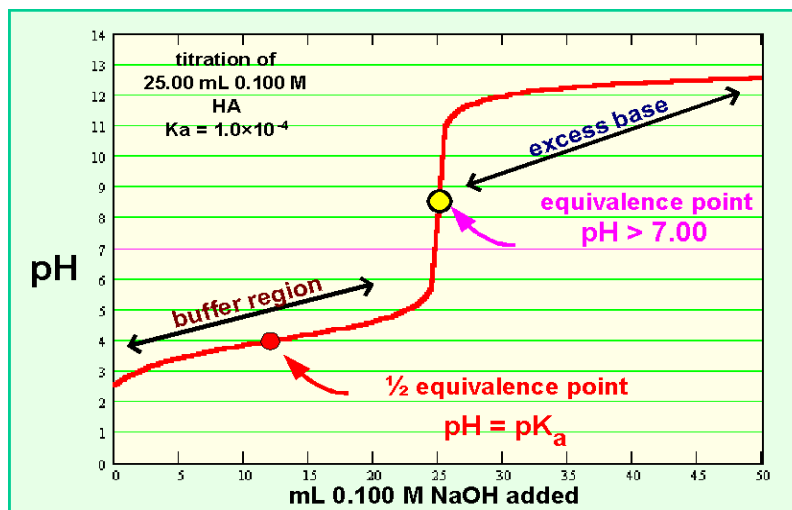
#### Materials:

- retort stand, clamp and white paper
- **conical flask** (used over a beaker for best swirling of solution)
- **burette**
- pipette [used to transfer the base into conical flask]
- **indicator, acid and base**

#### Method:

1. Fill a **burette** with a solution of known concentration (for example acetic acid) and adjust the solution level in the burette to the zero mark; the solution in the burette is called the **titrant**.
2. Place a solution of the sample to be analysed (for example sodium hydroxide) in a **conical flask** under the burette.
3. Add one or two drops of a suitable indicator to the flask. This should be determined using the evaluation of a *strong/weak* acid or base and the resultant salt formed.
4. Place a **white tile or card** under the flask to make it easier to detect the colour change in the indicator without overshooting the equivalence point.
5. Slowly run solution from the burette into the flask with **continuous swirling** until the indicator just changes colour. Drop solution in when near equivalence point.
6. Repeat the titration several times until a consistent volume of titrant is obtained (volumes differing by no more than about  $\pm 0.1$ ). Recording each.
7. Calculate the required concentration of the unknown using the basic principles of molarity calculations.



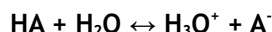


Before conducting a titration, a **primary standard** must be made to be titrated, known also as the titrate. The primary standard is created by **dissolving a known amount of a substance in a known volume water** and then calculating its *concentration* → this is done in a **volumetric flask**. A secondary standard is the concentration of the titrant which is calculated by finding the amount of titrate needed to neutralise the titrant.

#### 9. qualitatively describe the effect of buffers with reference to a specific example in a natural system

A **buffer** is a solution that **resists changes in pH** when an acid or a base is added. A buffer solution contains **comparable amounts of weak acid and its conjugate base** and which is therefore able to maintain approximately a constant pH.

There are many natural systems in life which require a constant or stable pH level to maintain and sustain life forms. For this reason, there are numerous effective and natural buffer systems in place to maintain this balance. The general equation for a buffer system is:



#### Fresh Water River Systems:

Are considered **buffer systems** only if there are comparable amounts of each weak acid and base. We know that natural rain which falls is a slightly acidic solution, very dilute amount of  $\text{H}_2\text{CO}_3$  (carbonic acid) which has been absorbed into the rain; however rain on its own is **not a buffer system**.

Natural rain which falls into these fresh water buffer systems must contain available resources and amounts of carbonate rocks. The buffer system that keeps an equilibrium buffer equation is as follows:



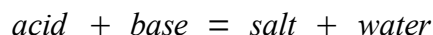
As comparable amounts of carbonic acid and carbonate are required in order for a fresh water river system to act as a buffer, **the absence of comparable amount will not permit the river system to act as a buffer**. *Effects of a river losing its ability to act as a buffer system include:*

- The river will **grow slightly acidic over time** as diluted concentrations of carbonic acid in rain fall into the river - pushing the equation to the left.
- Fish and other **aquatic life will not be able to exist** (in particular, sensitive trout) when the acidity varies too low so that the pH level is  $< 5.5$

If these effects are identified early enough, one would be able to add a base such as forms of carbonate or limestone to the natural river system to help the river return to its buffer system state. This is a similar way which pH can be detected in pools with pH strips.

#### 1. gather and process information from secondary sources to trace developments in understanding and describing acid/base reactions

**Arrhenius'** theory of acids and bases also posed the solution to when acids and bases would react with each, known as neutralisation. He proposed that when an acid reacts with a base, it is the **hydrogen** and **hydroxide** ions that react to form the neutral water:



Arrhenius' theory does solve and explain the many properties of acids and bases, however it is fallible

when considering the following examples:

1. **Metallic oxides and carbonates**: By experimentation, chemists know that **metallic oxides and carbonates neutralise acids**; however many carbonates are insoluble in water but will react with an acid to neutralise it. Carbonates and the metallic oxides also **do not contain the hydroxide ion** (property of a base) which must dissociate into water.



Arrhenius' theories could not explain such a phenomena.

2. **Varying acidity/basicity of salts**: We notice that salts composed of one part which is the same when comparing between two salts, some may be acidic while others are extremely basic. For example. A solution of sodium chloride is neutral while sodium sulfide is basic.

Once again, Arrhenius could not account for these instances and thus deemed his theories efficient and necessary, however, not sufficient to formulate a sustainable theory on a macro level.

2. choose equipment and perform a first-hand investigation to identify the pH of a range of salt solutions

3. 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

Tested using **HCl and NaOH** to create NaCl which is a neutral salt → and pH measurements were determined using titration methods (*refer to orange dot point 4.8*). The pH of the solution based on the concentration of  $\text{H}^+$  ions;  $\text{pH} = -\log[\text{H}_3\text{O}^+]$

4. perform a first-hand investigation to determine the concentration of a domestic acidic substance using computer-based technologies

5. 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

Neutralisation reactions can be used to minimise damage to infrastructure and humans when conducted correctly and in safe conditions to neutralise an acidic or basic spillage.

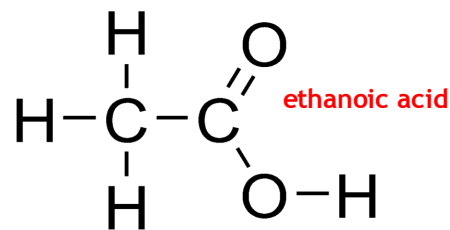
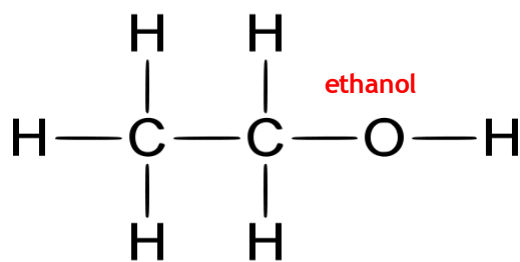
Spillages of strong acids and bases often occur in the lab and in such a situation a chemist is able to cease the corrosion of the acid or base by neutralising it (often with its conjugate acid/base). e.g. if there is a **strong acid spilt** on the floor, then a **weak base is required to neutralise the acid**; however **neutralisation reactions are exothermic and therefore produce heat which can be dangerous**. However, concentrated substances should never be used (in contrast with strong/weak) as this will only cause the situation to progress at a faster rate of reaction.

In large spillages where time is of the essence, the first measure taken to contain such a spillage of an acid/base across the roads would be to contain the spillage to prevent it from flowing into any nearby catchment/river systems or drains. Secondly, if an alkaline material has been spilt, the standard would be to use an amphiprotic substance such as **sodium hydrogen carbonate** or **water**. This is chosen because it can react with the unknown base or otherwise neutralise an unknown acid.

## 5. Esterification is a naturally occurring process which can be performed in the laboratory

1. describe the differences between the alkanol and alkanoic acid functional groups in carbon compounds  
**Alkanols** are compounds that contain the alcohol functional group (**−OH**).

The **alkanoic acid** family are **weak organic acids**. Contain the carboxylic acid functional group (**−COOH**). The alkanoic acids therefore contain a **double bonded oxygen atom** which allows for **two hydrogen bonds** instead of the one in the respective alkanol. The extra hydrogen bond results in higher boiling and melting temperatures.



On the left we observe the parent ethanol from which ethanoic acid can be formed. In each alkanol case, we observe the alkyl (-OH) group remains while two hydrogen atoms are replaced with one singularly, double-bonded oxygen atom.

Alkanoic acids have the general equation:  $C_nH_{2n+1}COOH$  (where  $n = 0, 1, 2, 3..$ )

2. 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

Esters are formed by reacting an alkanoic acid with an alkanol. The IUPAC nomenclature for naming such esters obeys the following rules.

1. Write down the **alkanol**, removing the “anol” and replace with “yl”
2. Write down the **alkanoic acid**, remove the “oic acid” and replace with “oate”
3. Place them together

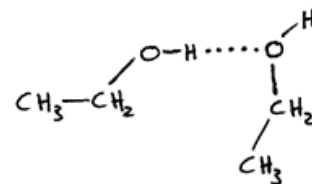
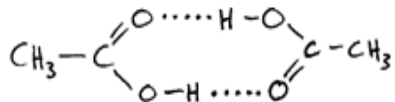
e.g. When *hexanoic acid* is reacted with *propanol*, we form: *propyl hexanoate*.

3. explain the difference in melting point and boiling point caused by straight-chained alkanoic acid and straight-chained primary alkanol structures

Alkanes, alkenes and alkynes (**hydrocarbons**) are not polar and do not have hydrogen bonding. They only have **weak dispersion forces**, this gives them **low boiling points**.

**Alkanols** have **1 centre of polarity** and form dipole bonding and **hydrogen bonding** between molecules and therefore have higher boiling points.

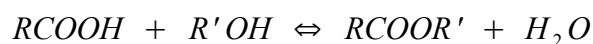
**Alkanoic acids** have **2 centres of polarity** and dipole bonding and **hydrogen bonding** and have even higher boiling points.



Considering the straight chain alkanoic acids and their parent alkanols, the *respective alkanoic acid will always have a higher melting/boiling point* due to the two centres of polarity and thus making the molecule harder to physically separate.

4. identify esterification as the reaction between an acid and an alkanol and describe, using equations, examples of esterification

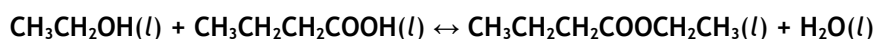
Esterification is the production of an ester due to the reaction between a parent alkanol and an alkanoic acid.



alkanoic acid + alkanol  $\leftrightarrow$  alkyl alkanoate + water

In the formation of an ester, the inactive part of the alkanoic acid (R) is taken along with the main part of the carboxylic group (the COO part) which is combined with the inactive group of the alkanol. The reaction to form an ester is an **equilibrium reaction**.

An example of the formation of an ester is the reaction between ethanol and butanoic acid to form ethyl butanoate. The reaction is as follows:



5. describe the purpose of using acid in esterification for catalysis

**Heat** is used in esterification as a catalyst to speed up the process so that the reaction will move to the



side in which will produce the least amount of heat - in this case, the reaction will continue to go to the right **and produce more the ester at a faster rate.**

**Concentrated Sulfuric acid** ( $\text{H}_2\text{SO}_4$ ) is used as a **dehydrating agent** by forming strong intermolecular bonds with the water molecule so that it can be removed from the system. Accordingly, with **Le Chatelier's principle**, the equilibrium will try to counter this adjustment to the reaction and therefore **the system will move to the right** to produce more water, and thus a higher yield of the ester. This is effectively a catalytic process since it speeds up the process.

#### 6. explain the need for refluxing during esterification

Since esterification is carried out at **high temperatures** the reactants (alkanols and alkanolic acids) will begin to evaporate due to their **volatility** and the heat of the reaction. **Refluxing** is a process of esterification in which **the vaporised reactants are condensed and cooled back from their gaseous form and return to the reaction vessel.**

Refluxing results in a **higher yield of the desired ester** and there is no waste of reactants.

#### 7. outline some examples of the occurrence, production and uses of esters

Esters are naturally occurring substances and can be found throughout nature and because of this they are extracted and produced so they can be utilised in other ways:

- **Artificial and Natural Flavourings:** The characteristic smells and tastes of fruits and flowers are due to the naturally occurring esters which inhabit them. **Butyl Butanoate** is the ester which is pineapple flavour. **Octyl acetate** is the ester for orange flavour.
- **Perfumes/Fragrances:** Some of the above smells can be extracted to be used in cosmetic perfumes and lotions. **Phenylmethyl acetate** is an ester with a *jasmine fragrance* that is **used in various cosmetics.**
- **Solvents and Coatings:** Many natural and synthetic esters are used in wood lacquers, thinners and a *variety of coatings*. Some low molecular weight esters have a high volatility and so are useful components of some solvents used in surface finishes. Those esters of higher molecular weight are less volatile and find use in latex paints. An example of such a solvent is **Ethyle Ethanoate** which is used as a **cleaning fluid, coatings and printing.**

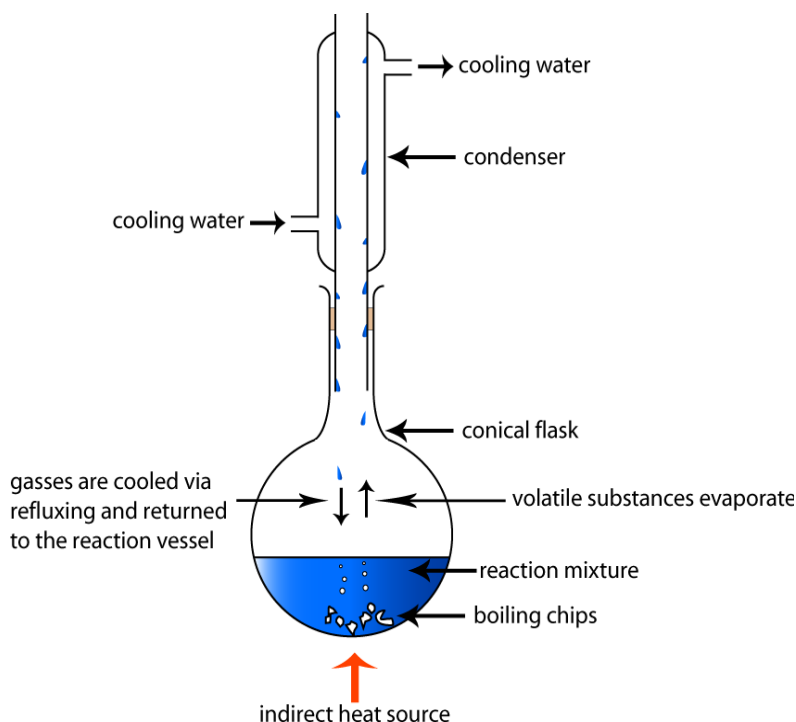
#### 1. identify data, plan, select equipment and perform a first-hand investigation to prepare an ester using reflux

The apparatus required involves measuring the known/required amounts of an **alkanol and an alkanolic acid** and placing them in a **conical flask**. On the neck of this conical flask, a cooling apparatus known as a **reflux condenser** is attached to ensure that the evaporated, volatile reactants are returned to the reacting reservoir which ensures the highest yield of ester.

Esterification must be **conducted under the use of a catalyst, heat**, and so an indirect heat source is used to heat the alkanol/alkanoic acid solution. An **indirect heat source may come from a beaker filled with water that is being boiled by a Bunsen burner**, so that the steam heats the conical bulb. Boiling chips can be used to speed up and catalyse the reaction process.

For further catalytic speeds, **concentrated sulfuric acid** can be added to the reacting mixture since it acts as a dehydrating agent. We must take precaution when using concentrated  $\text{H}_2\text{SO}_4$  because it is **extremely corrosive** and any contact with skin should be immediately rinsed under water and assistance sought.

**Safety glasses** must also be worn since we are dealing with highly volatile substances and also boiling a mixture. After several minutes of high temperature esterification we are able to assume that the reactants have reacted.



## 2. process information from secondary sources to identify and describe the uses of esters as flavours and perfumes in processed foods and cosmetics

Esters are used for flavouring and artificial smells in processed foods that can be purchased from the shelf in many stores. Esters are only used in foods for simply sensory means so that a **satisfactory taste and smell is achieved**. **Butyl Butanoate - Pineapple Flavour**

Esters which are used in cosmetics are used not only for fragrance purposes but can also act as emulsifiers and **hydrophilic** substances. Esters with **high molecular weights are used in emulsifiers** to prevent the oils and waters of the cosmetic product from separating. This is done to achieve the best customer result and satisfaction. **Phenylmethyl acetate - jasmine scented** *also used in emulsifiers*.