

Topic 2
Focus 1

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

- Classify common substances as acidic, basic or neutral.

An **acid** is a substance which in solution produces **hydrogen ions, H⁺**, or more strictly **hydronium ions, H₃O⁺**.

A **base** is a substance which either contains the oxide O²⁻ or the **hydroxide ion, OH⁻**. That, or a substance that when in solution produces the **hydroxide ion**.

A soluble base is called an **alkali**.

A **neutral solution** is one in which the **hydrogen ion concentration** and **hydroxide ion concentration** is **equal**.

Properties of:

Acids:

- Sour taste
- Sting or burn the skin
- Conduct electricity in solution
- Turns blue litmus red. (**B**lue in **A**cid goes **R**ed – **BAR**)

Bases:

- Bitter taste
- Slippery or soapy feel
- Conduct electricity in solution
- Turns red litmus blue.

Reactions:

Acid + Base → Salt + Water

Acid + Metal → Salt + Hydrogen gas

Acid + Carbonate → Salt + Water + Carbon dioxide gas

Common Examples:

Acids	Bases	Neutral
Lemon juice	Milk of Magnesia	Distilled water
Vinegar	Ammonia cleanser	Sugar
Carbonated soft drinks	Caustic soda solution	Pure alcohol
Citrus juices	Washing soda	Sodium Chloride
Formic acid	Antacid tablets	

- Identify that indicators such as

Indicators are substances that change colour in solution depending on the **alkalinity/acidity** of that solution.

<p>litmus, phenolphthalein, methyl orange and bromothymol blue can be used to determine the acidic or basic nature of a material over a range identified by a change in indicator colour.</p>	<p>Bromothymol Blue:</p> <table border="1" data-bbox="552 226 1415 304"> <thead> <tr> <th>H. Acidic</th> <th>S. Acidic</th> <th>Neutral</th> <th>S. Basic</th> <th>H. Basic</th> </tr> </thead> <tbody> <tr> <td>Yellow</td> <td>Yellow</td> <td>→</td> <td>Blue</td> <td>Blue</td> </tr> </tbody> </table> <p>Litmus:</p> <table border="1" data-bbox="552 378 1415 456"> <thead> <tr> <th>H. Acidic</th> <th>S. Acidic</th> <th>Neutral</th> <th>S. Basic</th> <th>H. Basic</th> </tr> </thead> <tbody> <tr> <td>Red</td> <td>Red</td> <td>→</td> <td>Blue</td> <td>Blue</td> </tr> </tbody> </table> <p>Methyl Orange:</p> <table border="1" data-bbox="552 530 1415 609"> <thead> <tr> <th>H. Acidic</th> <th>S. Acidic</th> <th>Neutral</th> <th>S. Basic</th> <th>H. Basic</th> </tr> </thead> <tbody> <tr> <td>Red →</td> <td>Yellow</td> <td>Yellow</td> <td>Yellow</td> <td>Yellow</td> </tr> </tbody> </table> <p>Phenolphthalein:</p> <table border="1" data-bbox="552 683 1415 761"> <thead> <tr> <th>H. Acidic</th> <th>S. Acidic</th> <th>Neutral</th> <th>S. Basic</th> <th>H. Basic</th> </tr> </thead> <tbody> <tr> <td>C'less</td> <td>C'less</td> <td>C'less</td> <td>C'less</td> <td>→ Red</td> </tr> </tbody> </table>	H. Acidic	S. Acidic	Neutral	S. Basic	H. Basic	Yellow	Yellow	→	Blue	Blue	H. Acidic	S. Acidic	Neutral	S. Basic	H. Basic	Red	Red	→	Blue	Blue	H. Acidic	S. Acidic	Neutral	S. Basic	H. Basic	Red →	Yellow	Yellow	Yellow	Yellow	H. Acidic	S. Acidic	Neutral	S. Basic	H. Basic	C'less	C'less	C'less	C'less	→ Red
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<ul style="list-style-type: none"> Identify and describe some everyday uses of indicators including the testing of soil acidity/basicity. 	<p>Everyday Uses:</p> <ul style="list-style-type: none"> Monitoring Pool Water <ul style="list-style-type: none"> Acidity levels need to be carefully controlled to avoid skin & eye irritation. Regularly tested to determine pH. Soil Industry <ul style="list-style-type: none"> Includes both industrial and domestic use. Soil is moistened with water and indicator added. A neutral powder, barium sulfate, is placed on the soil so the colour change can be observed. Certain plants require certain levels of acidity/basicity for them to grow. Chemical Analysis: <ul style="list-style-type: none"> Used in titrations to determine when an acid has neutralised a base. Monitoring industrial wastes 																																								
<ul style="list-style-type: none"> Perform a first-hand investigation to prepare and test a natural indicator. 	<p><u>Experiment:</u> <u>Aim:</u> To prepare an indicator solution from red cabbage and test the indicator on a range of substances. <u>Method:</u></p> <ol style="list-style-type: none"> Put on safety glasses. Shredded cabbage leaves using scissors and placed in 800ml beaker. Added distilled water to beaker until water level just covered shredded cabbage leaves. Slowly boiled the water until the water turned dark reddish-purple and the leaves lost most of their colour. Allowed beaker to cool. Poured off liquid from 800ml beaker into clean 250ml beaker. This became the red cabbage indicator (RCI). Placed approximately 2ml of a known acid (HCl) into a test tube and added a few drops of RCI. Recorded colour change. This became reference for other substances. Placed approximately 2ml of a known base (NaOH) into a 																																								

	<p>test tube and added a few drops of RCl. Recorded colour change. This became the reference colour for bases.</p> <ol style="list-style-type: none"> Filled two test tubes with 2ml of other common substances. Added RCl to one test tube and recorded colour change. Added Universal Indicator to other test tube and recorded colour change. Determined whether the substance was an acid, base or neutral solution. Repeated steps 6 & 7 for all the test substances ensuring that solid substances were first dissolved into a solution with distilled water. <p><u>Results:</u> The pH of each solution could be determined.</p> <p><u>Conclusion:</u> Red cabbage becomes an effective indicator which remains purple in neutral solutions, turns green or yellow when basic or turns red when acidic.</p>																																								
<ul style="list-style-type: none"> Identify data and choose resources to gather information about the colour changes of a range of indicators. 	<p>The colour changes of various indicators are shown below:</p> <p>Bromothymol Blue:</p> <table border="1" data-bbox="552 927 1423 1003"> <thead> <tr> <th>H. Acidic</th> <th>S. Acidic</th> <th>Neutral</th> <th>S. Basic</th> <th>H. Basic</th> </tr> </thead> <tbody> <tr> <td>Yellow</td> <td>Yellow</td> <td>→</td> <td>Blue</td> <td>Blue</td> </tr> </tbody> </table> <p>Litmus:</p> <table border="1" data-bbox="552 1077 1423 1153"> <thead> <tr> <th>H. Acidic</th> <th>S. Acidic</th> <th>Neutral</th> <th>S. Basic</th> <th>H. Basic</th> </tr> </thead> <tbody> <tr> <td>Red</td> <td>Red</td> <td>→</td> <td>Blue</td> <td>Blue</td> </tr> </tbody> </table> <p>Methyl Orange:</p> <table border="1" data-bbox="552 1227 1423 1303"> <thead> <tr> <th>H. Acidic</th> <th>S. Acidic</th> <th>Neutral</th> <th>S. Basic</th> <th>H. Basic</th> </tr> </thead> <tbody> <tr> <td>Red</td> <td>→ Yellow</td> <td>Yellow</td> <td>Yellow</td> <td>Yellow</td> </tr> </tbody> </table> <p>Phenolphthalein:</p> <table border="1" data-bbox="552 1377 1423 1453"> <thead> <tr> <th>H. Acidic</th> <th>S. Acidic</th> <th>Neutral</th> <th>S. Basic</th> <th>H. Basic</th> </tr> </thead> <tbody> <tr> <td>C'less</td> <td>C'less</td> <td>C'less</td> <td>C'less</td> <td>→ Red</td> </tr> </tbody> </table>	H. Acidic	S. Acidic	Neutral	S. Basic	H. Basic	Yellow	Yellow	→	Blue	Blue	H. Acidic	S. Acidic	Neutral	S. Basic	H. Basic	Red	Red	→	Blue	Blue	H. Acidic	S. Acidic	Neutral	S. Basic	H. Basic	Red	→ Yellow	Yellow	Yellow	Yellow	H. Acidic	S. Acidic	Neutral	S. Basic	H. Basic	C'less	C'less	C'less	C'less	→ Red
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<ul style="list-style-type: none"> Solve problems by applying information about the colour changes of indicators to classify some household substances as acidic, neutral or basic. 	<p>Use the information provided (such as the tables above) to determine what the acidity/basicity of a solution is by aligning the colour change of that indicator to the appropriate determination of acidity/basicity.</p>																																								
<p>Focus 2 While we usually think of the air around us as neutral, the atmosphere naturally</p>																																									

<p>contains acidic oxides of carbon, nitrogen and sulfur. The concentrations of these acidic oxides have been increasing since the Industrial Revolution.</p>	
<ul style="list-style-type: none"> Identify oxides of non-metals which act as acids & describe the conditions under which they act as acids. 	<p>Non-metal oxides</p> <ul style="list-style-type: none"> Are generally acidic oxides. (Exceptions: CO, NO and N₂O) <p>An acidic oxide is one that either:</p> <ul style="list-style-type: none"> Reacts with water to form an acid, or Reacts with bases to form salts <p>Common Examples:</p> $\text{CO}_2(g) + \text{H}_2\text{O}(l) \longrightarrow \text{H}_2\text{CO}_3(aq) \quad (\text{Carbonic Acid})$ $\text{P}_2\text{O}_5(s) + 3\text{H}_2\text{O}(l) \longrightarrow 2\text{H}_3\text{PO}_4(aq) \quad (\text{Phosphoric Acid})$ <p>Amphoteric oxides react with acids to forms salts but also react with alkalis.</p>
<ul style="list-style-type: none"> Analyse the position of these non-metals in the Periodic table and outline the relationship between the position of elements on the Periodic table and the acidity/basicity of oxides. 	<p>General Rules:</p> <ul style="list-style-type: none"> Acidic oxides are generally oxides of non-metals. Basic oxides are oxides of metals. They are ionic compounds. As you go down a group, the more basic the oxide becomes. Group 1 Oxides – Strongly basic Group 2 Oxides – Basic Transition metals – Generally basic Groups 3-6 – Trend down the group (Acidic → Amphoteric → Basic) Group 7 Oxides – Strongly acidic Group 8 – No oxides.
<ul style="list-style-type: none"> Define Le Chatelier's principle. 	<p>Definition:</p> <p>If a system is at equilibrium and a change is made that upsets the equilibrium, then the system will alter in such a way as to counteract the change and establish a new equilibrium.</p> <p>A system is at equilibrium if the ratio of products to reactants remain constant (but not necessarily equal).</p>
<ul style="list-style-type: none"> Identify factors which can affect the equilibrium in a reversible reaction. 	<p>Factors affecting Equilibrium:</p> <ul style="list-style-type: none"> Concentration – Increasing the concentration of a substance favours the direction which will use up more of that substance. Pressure – Increasing the pressure (when gases are involved) will favour the direction that will decrease the

	<p>pressure on the system and use up some of the gases.</p> <ul style="list-style-type: none"> • Temperature – Increasing the temperature favours the endothermic reaction. Decreasing the temperature favours the exothermic reaction.
<ul style="list-style-type: none"> • Describe the solubility of carbon dioxide in water under various conditions as an equilibrium process and explain in terms of Le Chatelier's Principle. 	<p>The reaction between carbon dioxide and water is:</p> $\text{CO}_2(g) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_2\text{CO}_3(aq) \quad (\text{Carbonic Acid})$ <p>Changing Temperatures:</p> <ul style="list-style-type: none"> • Increasing temperature shifts equilibrium left. • Decreasing temperature shifts equilibrium right. <p>Changing Concentrations:</p> <ul style="list-style-type: none"> • Adding an acid (H⁺ ions) shifts equilibrium left. • Adding a base shifts equilibrium right. <p>Changing Pressure:</p> <ul style="list-style-type: none"> • Increasing pressure shifts equilibrium right. • Decreasing pressure shifts equilibrium left. <p>All these changes occur because Le Chatelier's principle states that when a change is made to a system in equilibrium, the system will alter to balance the change.</p>
<ul style="list-style-type: none"> • Identify natural and industrial sources of sulfur dioxide and oxides of nitrogen. • Describe, using equations, examples of chemical reactions which release sulfur dioxide and chemical reactions which release oxides of nitrogen. 	<p>Sulfur Dioxide:</p> <p><u>Natural Causes:</u></p> <ul style="list-style-type: none"> • Bacterial decomposition of organic matter produces hydrogen sulfide (H₂S) which oxides to form sulfur dioxide. $2\text{H}_2\text{S}(g) + 3\text{O}_2(g) \longrightarrow 2\text{SO}_2(g) + 2\text{H}_2\text{O}(l)$ <ul style="list-style-type: none"> • Combustion of organic matter in bushfires. • Geological activity – Volcanoes and hot springs. <p><u>Industrial Causes:</u></p> <ul style="list-style-type: none"> • Major producer is the combustion of fossil fuels containing sulfur. These sulfur impurities usually occur as FeS₂. $4\text{FeS}_2(s) + 11\text{O}_2(g) \longrightarrow 2\text{Fe}_2\text{O}_3(s) + 8\text{SO}_2(g)$ <ul style="list-style-type: none"> • The smelting of sulfide ores also produces sulfur dioxide. $2\text{ZnS}(s) + 3\text{O}_2(g) \longrightarrow \text{ZnO}(s) + 2\text{SO}_2(g)$ <p>Nitrogen Oxides:</p> <p><u>Types of Nitrogen Oxides:</u></p> <ul style="list-style-type: none"> • NO₂ – Nitrogen dioxide • NO – Nitric oxide • N₂O – Nitrous oxide <p><u>Natural Sources:</u></p> <ul style="list-style-type: none"> • Nitric oxide is formed through the action of lightning on oxygen and nitrogen gases. $\text{N}_2(g) + \text{O}_2(g) \longrightarrow 2\text{NO}(g)$ <ul style="list-style-type: none"> • Nitrogen dioxide is formed slowly through the action of sunlight on nitric oxide.

	$2\text{NO}(g) + \text{O}_2(g) \longrightarrow 2\text{NO}_2(g)$ <ul style="list-style-type: none"> Soil bacteria produce both Nitric oxide and nitrous oxide. <p><u>Industrial Sources:</u></p> <ul style="list-style-type: none"> The major source is high temperature combustion of fossil fuels in both internal combustion engines and coal fired power plants. $\text{N}_2(g) + \text{O}_2(g) \xrightarrow{\text{High Temp}} 2\text{NO}$ $\text{NO}_2(g) + \text{H}_2\text{O} \longrightarrow \text{H}_2\text{NO}_3$
<ul style="list-style-type: none"> Assess the evidence which indicates increases in atmospheric concentration of oxides of sulfur and nitrogen. 	<p>After the industrial revolution of the early 1800's there was a great increase in the emissions of sulfur dioxide. This continued to rise until the 1950's and 1960's, when deaths were recorded as a result of the air pollution. This was when stringent regulations were brought in.</p> <p>Reasons for Increase in Sulfur oxides:</p> <ul style="list-style-type: none"> Power station combustion of fossil fuels. Copper smelting plants. <p>Reasons for Increase in oxides of nitrogen:</p> <ul style="list-style-type: none"> The increase in the use of motor vehicles. Photochemical smog. <p>Evidence behind the figures:</p> <ul style="list-style-type: none"> Only in the last 30 years has the technology become available for testing the actual levels in the atmosphere so not a lot of historical data is known. Sulfur dioxide and nitrogen dioxide form ions that dissolve in water making them difficult to measure.
<ul style="list-style-type: none"> Calculate volumes of gases given masses of some substances in reactions and calculate masses of some substances given gaseous volumes, in reactions involving gases at 0° & 100kPa or 25° & 100kPa. 	<p>There is a law that states that one mole of two different gases will take up exactly the same volume if they are at the same temperature and pressure.</p> <p>To find the volume, multiply the moles of a gas by the appropriate value for L mol⁻¹ depending on the temperature and pressure.</p> <p>The values are:</p> <ul style="list-style-type: none"> 22.71 L mol⁻¹ at 0° and 100kPa. 24.79 L mol⁻¹ at 25° and 100kPa.
<ul style="list-style-type: none"> Explain the formation and effects of acid rain. 	<p>Acid rain is defined as rain with a pH < 5. It is usually the result of rain dissolving oxides of sulfur and nitrogen.</p> <p>Formation: Formation of nitrous and nitric acid: $2\text{NO}_2(g) + \text{H}_2\text{O}(l) \longrightarrow \text{HNO}_2 + \text{HNO}_3$ Formation of sulfurous acid from sulfur dioxide.</p>

	<p> $\text{SO}_2(g) + \text{H}_2\text{O}(l) \longrightarrow \text{H}_2\text{SO}_3(aq)$ Formation of sulfuric acid from sulfur trioxide. $\text{SO}_3(g) + \text{H}_2\text{O}(l) \longrightarrow \text{H}_2\text{SO}_4(aq)$ </p> <p>Effects:</p> <ul style="list-style-type: none"> • Increased weathering and erosion <ul style="list-style-type: none"> ○ Damage to statues and buildings ○ Rocks crumbling away ○ Limestone, dolomite and marble damaged • Change in pH of soil <ul style="list-style-type: none"> ○ Kills plants (including crops) ○ Depletes soil of important elements (eg: Mg, K, Ca) ○ Difficult for plants to absorb nutrients • Damage to aquatic ecosystems <ul style="list-style-type: none"> ○ High acidity prevents fish and aquatic life from thriving ○ Continued acidity will kill aquatic life. • Structural metals such as steel can be attacked.
<ul style="list-style-type: none"> • 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° and 100kPa. 	<p><u>Experiment:</u></p> <p><u>Aim:</u> To determine the amount of carbon dioxide in a bottle of soft drink.</p> <p><u>Equipment:</u></p> <ul style="list-style-type: none"> • 2 Schweppes® Dry Ginger Ales 315ml. • Beam balance • Water <p><u>Method:</u></p> <ol style="list-style-type: none"> 1. Record the mass of one of the unopened bottles. This bottle will now be known as bottle #1. 2. Open the cap of bottle #2. 3. Drink contents of bottle #2. 4. Fill bottle #2 with water to the same level as bottle #1. 5. Record mass of bottle #2 and water. 6. Open the cap of bottle #1. 7. Leave both bottles with their caps off for five hours. 8. Re-record mass both bottles. 9. Calculate loss of mass for both bottles. 10. Allow for changes in bottle #1's mass due to evaporation by the figures taken from bottle #2. 11. Calculate litres of gas lost. <p><u>Results:</u> Mass lost = 2.3 g $\text{Moles of CO}_2 = \frac{2.3}{44} = 0.052272727$ $\text{Volume of Gas Lost} = 0.052272727 \times 24.79 = 1.29584 \text{ L}$</p> <p><u>Conclusion:</u> A bottle of soft drink contains approximately 1.3 litres of carbon dioxide gas.</p>
<ul style="list-style-type: none"> • Analyse information from 	<p>Sulfur Dioxide: Is a chemical compound made up of sulfur and oxygen. It is</p>

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

made commercially by roasting sulfur or sulfides in air. It is also frequently produced as a by product of smelting sulfide ores such as zinc sulfide. Sulfur dioxide is used in the production of other chemicals, as it is used in the manufacturing of sulfuric acid and sodium sulfite, which is used to make paper. It is also used as a food preservative. Sulfur dioxide emitted into the atmosphere by industrial processes is ultimately converted into dilute sulfurous acid, returning to the earth as acid rain.

Oxides of Nitrogen:

They enter the atmosphere either through human or natural activities. The high temperatures in internal combustion engines or in coal fired power stations lead to a reaction between nitrogen and oxygen. The products are often then released into the atmosphere. Nitric oxide in the air leads to photochemical smog, which can cause breathing difficulties.

Focus 3

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

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

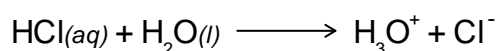
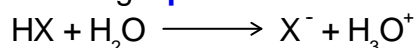
Acids are defined as **proton donors**.

When an acid molecule comes in contact with water it **ionises**, donating a **proton** to water.

A **hydrogen atom** consists of one **proton** and one **electron**. So when a **hydrogen atom** loses an **electron**, it becomes a **hydrogen ion**. From this, a **hydrogen ion** and a **proton** are the same thing.

Acids in Water:

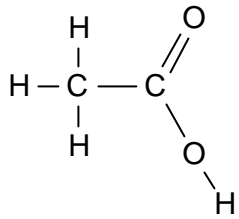
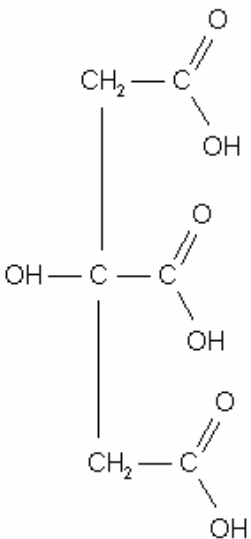
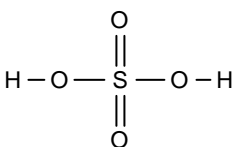
When an **acid molecule** is placed in water, it **ionises**, releasing a **proton**.



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

Common Acids:

Name(s)	Formula / Structure	Found in / Uses:
Acetic Acid Ethanoic Acid	CH ₃ COOH	<ul style="list-style-type: none"> Vinegar is 4% acetic acid. Forms when wine naturally ferments.

			
	<p>Citric Acid</p> <p>2-Hydroxypropane 1,2,3- tricarboxylic Acid</p>	$C_6H_8O_7$ 	<ul style="list-style-type: none"> • Citrus fruit • Formed during cellular respiration in body cells.
	<p>Hydrochloric Acid</p>	HCl $H - Cl$	<ul style="list-style-type: none"> • Stomach acid • High acidity activates enzymes in the stomach catalysing the breakdown of proteins.
	<p>Sulfuric Acid</p>	H_2SO_4 	<ul style="list-style-type: none"> • Cooking copper and iron sulfate minerals. • Sulfur dioxide emissions from volcanoes.
<ul style="list-style-type: none"> • Describe the use of the pH scale in comparing acids and bases. 	<p>The pH scale is used to compare the concentration of hydrogen ions and hydroxide ions in solutions of acids and bases.</p> <p>pH stands for: potential Hydrogen</p>		

	<ul style="list-style-type: none"> In an acidic solution, $[H^+] > 10^{-7} \text{ mol L}^{-1}$ and $\text{pH} < 7$. In a basic solution, $[H^+] < 10^{-7} \text{ mol L}^{-1}$ and $\text{pH} > 7$. In a neutral solution, $[H^+] = 10^{-7} \text{ mol L}^{-1}$ and $\text{pH} = 7$.
<ul style="list-style-type: none"> Describe acids and their solutions with the appropriate use of the terms strong, weak, concentrated and dilute. 	<p>A concentrated solution is one in which the total concentration of solute species is high.</p> <p>A dilute solution is one in which the total concentration of solute species is low.</p> <p>A strong acid is one in which all the acid present in solution has ionised to form hydrogen ions. There are no neutral acid molecules left.</p> <p>A weak acid is one in which only some of the acid molecules present in the solution have ionised to form hydrogen ions.</p>
<ul style="list-style-type: none"> Identify pH as $-\log_{10}[H^+]$ and explain that a change in pH of one means a ten fold change in $[H^+]$. 	<p>When the pH of a solution changes by one, the hydrogen ion concentration changes by tenfold.</p> <ul style="list-style-type: none"> If a solution's pH increases by one, the hydrogen ion concentration is divided by 10. If a solution's pH decreases by one, the hydrogen ion concentration is multiplied by 10. <p>The Formula: $\text{pH} = -\log_{10} [H^+]$</p>
<ul style="list-style-type: none"> Compare the relative strengths of equal concentrations of citric, acetic and hydrochloric acid and explain in terms of degree of ionisation of their molecules. 	<p>Hydrochloric Acid:</p> <ul style="list-style-type: none"> HCl Hydrochloric acid is a strong acid. Almost 100% ionisation. $\text{HCl} + \text{H}_2\text{O} \longrightarrow \text{Cl}^- + \text{H}_3\text{O}^+$ <p>Citric Acid:</p> <ul style="list-style-type: none"> $\text{C}_6\text{H}_8\text{O}_7$ 2-Hydroxypropane 1,2,3 Tricarboxylic Acid Citric acid is a weak acid. It only undertakes around 1% ionisation. $\text{C}_6\text{H}_8\text{O}_7 + \text{H}_2\text{O} \longleftrightarrow \text{C}_6\text{H}_7\text{O}_7^- + \text{H}_3\text{O}^+$ $\text{C}_6\text{H}_7\text{O}_7^- + \text{H}_2\text{O} \longleftrightarrow \text{C}_6\text{H}_6\text{O}_7^{2-} + \text{H}_3\text{O}^+$ $\text{C}_6\text{H}_6\text{O}_7^{2-} + \text{H}_2\text{O} \longleftrightarrow \text{C}_6\text{H}_5\text{O}_7^{3-} + \text{H}_3\text{O}^+$ <p>Acetic Acid:</p> <ul style="list-style-type: none"> CH_3COOH Acetic acid is a weak acid. It only undertakes in around 1% ionisation. $\text{CH}_3\text{COOH} + \text{H}_2\text{O} \longleftrightarrow \text{CH}_3\text{COO}^- + \text{H}_3\text{O}^+$
<ul style="list-style-type: none"> Describe the difference between 	<p>A strong acid completely ionises in water. The single pointed arrow, \rightarrow, is used to show this.</p>

<p>a strong acid and a weak acid in terms of an equilibrium between the intact molecule and its ions.</p>	<p>A weak acid does not completely ionise in water and thus the acid is in an equilibrium with its ions. This is shown by the double pointed arrow, \leftrightarrow.</p>																																																																
<ul style="list-style-type: none"> Solve problems and perform a first hand investigation to use pH probes and indicators to distinguish between acidic, basic and neutral chemicals. 	<p><u>Experiment:</u> <u>Aim:</u> To measure and record the pH of various household chemicals using a pH probe and universal indicator. <u>Method:</u></p> <ol style="list-style-type: none"> Put on safety glasses. Pour approximately 2-3 cm of a chemical solution into a beaker. Use pH probe to measure pH. Record value. Add a few drops of universal indicator. Use colour chart to determine pH. Record colour & pH. (In cases of opaque or dark solutions use universal indicator paper instead of solution) Determine hydrogen ion concentration. Repeat steps 2 to 5 for each chemical and distilled water (control). <p><u>Results:</u></p> <table border="1" data-bbox="603 1039 1374 1682"> <thead> <tr> <th>Substance</th> <th>Probe pH</th> <th>UI pH</th> <th>Acid/Base/Neutral</th> </tr> </thead> <tbody> <tr> <td>Distilled Water</td> <td>7.3</td> <td>7</td> <td>Neutral</td> </tr> <tr> <td>H₂SO₄</td> <td>1.1</td> <td>= 3</td> <td>Acid</td> </tr> <tr> <td>MgSO₄</td> <td>6.7</td> <td>7</td> <td>Neutral</td> </tr> <tr> <td>Vinegar</td> <td>3.2</td> <td>= 3</td> <td>Acid</td> </tr> <tr> <td>HCl</td> <td>2.3</td> <td>= 3</td> <td>Acid</td> </tr> <tr> <td>Soy Sauce</td> <td>5.4</td> <td>5</td> <td>Acid</td> </tr> <tr> <td>Hand Lotion</td> <td>7.3</td> <td>7</td> <td>Neutral</td> </tr> <tr> <td>Lab glass cleaner</td> <td>7.9</td> <td>8</td> <td>Base</td> </tr> <tr> <td>NaOH</td> <td>11.1</td> <td>= 11</td> <td>Base</td> </tr> <tr> <td>Cream of tartar</td> <td>3.9</td> <td>4</td> <td>Acid</td> </tr> <tr> <td>Washing Powder</td> <td>10.8</td> <td>= 11</td> <td>Base</td> </tr> <tr> <td>Salvital</td> <td>3.9</td> <td>4</td> <td>Acid</td> </tr> <tr> <td>Limewater</td> <td>11.4</td> <td>= 11</td> <td>Base</td> </tr> <tr> <td>Seltzer</td> <td>5.8</td> <td>5</td> <td>Acid</td> </tr> <tr> <td>NaCO₃</td> <td>10.2</td> <td>10</td> <td>Base</td> </tr> </tbody> </table> <p><u>Conclusion:</u> Once the pH was determined, the solution could be identified as acidic, basic or neutral. pH < 7: Acidic pH = 7: Neutral pH > 7: Basic</p>	Substance	Probe pH	UI pH	Acid/Base/Neutral	Distilled Water	7.3	7	Neutral	H ₂ SO ₄	1.1	= 3	Acid	MgSO ₄	6.7	7	Neutral	Vinegar	3.2	= 3	Acid	HCl	2.3	= 3	Acid	Soy Sauce	5.4	5	Acid	Hand Lotion	7.3	7	Neutral	Lab glass cleaner	7.9	8	Base	NaOH	11.1	= 11	Base	Cream of tartar	3.9	4	Acid	Washing Powder	10.8	= 11	Base	Salvital	3.9	4	Acid	Limewater	11.4	= 11	Base	Seltzer	5.8	5	Acid	NaCO ₃	10.2	10	Base
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<ul style="list-style-type: none"> Plan and perform a first hand investigation to 	<p><u>Experiment:</u> <u>Aim:</u> To compare the pH of strong and weak acids of the same</p>																																																																

measure the pH of identical concentrations of strong and weak acids.

concentration.

Method:

1. Calculate the molar mass of citric acid.
2. Determine the mass of citric acid needed to make a 0.1M solution of citric acid.
3. Make the 0.1M solution of citric acid.
4. Pour approximately 25ml of citric acid and hydrochloric acid into two separate beakers.
5. Use the pH probe to record the pH of each solution. Compare the results.

Results:

Substance	pH
Citric Acid	2.1
Hydrochloric Acid	1.1

Calculations:

Molar Mass of Citric Acid: 192.12532 grams

\searrow 0.1 moles of Citric Acid: 19.212532 grams

Converting this figure from per litre to 25ml:

Mass of Citric Acid powder needed to make

25ml of 0.1M Citric Acid solution = 0.4803 grams

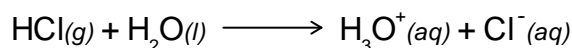
Conclusion:

The pH was lower for the strong acid while the weak acid registered a higher pH.

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

Example:

The reaction of hydrochloric acid with water.



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

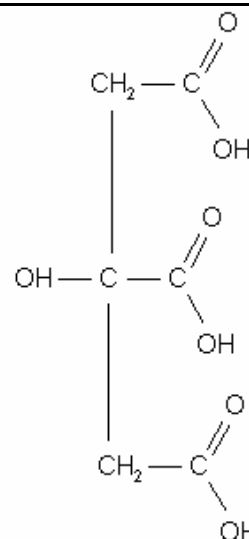
Activity:

Use molecular model kits.

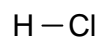
1. Make 4 HCl molecules, 4 citric acid molecules and eight water molecules.
2. First remove the hydrogen from HCl and attach it to the water molecule to model the hydronium ion. Do this 4 times. This will simulate complete ionisation.
3. Remove one hydrogen from citric acid and attach to water. Do this only once and it will model incomplete ionisation.

The Solutions:

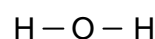
Citric Acid



Hydrochloric Acid



Water



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

Additives and preservatives are used to maintain product consistency and quality, improve or maintain nutritional value, maintain palatability and wholesomeness, provide leavening, control pH, enhance flavour, or provide colour.

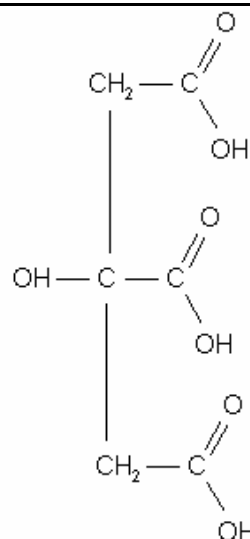
Additive	Uses	How does it work?
Sulfur Dioxide	<ul style="list-style-type: none"> • Dried fruit • Wine • Processed potatoes 	Sulfiting agents prevent discoloration. They also destroy vitamin B-1 and, most important, can cause severe reactions, especially in asthmatics.
Lactic Acid	<ul style="list-style-type: none"> • Spanish olives • Cheese • Frozen desserts • Carbonated beverages 	This safe acid occurs in almost all living organisms. It inhibits spoilage in Spanish-type olives, balances the acidity in cheese-making, and adds tartness to frozen desserts, carbonated fruit-flavoured drinks, and other foods.
Acetic Acid	<ul style="list-style-type: none"> • Thickening and sterilising agent • Prevents sugar from 	CMC is made by reacting cellulose with a derivative of acetic acid. Studies indicate it is safe.

		crystallizing <ul style="list-style-type: none"> • Ice cream • Beer • Candy 	
	Ascorbic Acid	<ul style="list-style-type: none"> • Antioxidant • Nutrient colour stabilizer • Cereals • Fruit drinks • Cured meats 	Ascorbic acid helps maintain the red colour of cured meat and prevents the formation of nitrosamines, which promote cancer. It helps prevent loss of colour and flavour by reacting with unwanted oxygen. Sodium ascorbate is a more soluble form of ascorbic acid.
	Propionic Acid	•	
	Citric Acid	<ul style="list-style-type: none"> • Flavouring • Ice cream • Sherbet • Fruit drink • Candy • Carbonated beverages • Instant potatoes 	Citric acid is versatile, widely used, cheap, and safe. It is an important metabolite in virtually all living organisms and is especially abundant naturally in citrus fruits and berries. It is used as a strong acid, a tart flavouring, and an antioxidant. Sodium citrate, also safe, is a buffer that controls the acidity of gelatin desserts, jam, ice cream, candy, and other foods.

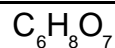
- Identify from data, gather and process information from secondary sources to identify examples of naturally occurring acids and bases and their chemical composition.

Citric Acid:

Structure:

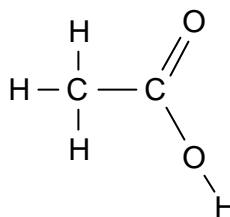


Formula:

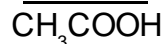


Acetic Acid:

Structure:



Formula:



- **Process information from secondary sources to calculate the pH of strong acids given appropriate hydrogen ion concentrations.**

Substitute known values in to equation to find pH.

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

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

- **Outline the**

Antoine Lavoisier (1790):

<p>historical development of ideas about acids including those of</p> <ul style="list-style-type: none"> ○ Lavoisier ○ Davy ○ Arrhenius 	<ul style="list-style-type: none"> • An acid was a substance that must contain oxygen. • This was later disproved as acids such as HCl does not include oxygen. <p>Humphry Davy (1815):</p> <ul style="list-style-type: none"> • An acid was a substance that contained hydrogen. These hydrogen atoms could be replaced by reaction with a metal. <p>Svante Arrhenius (1884):</p> <ul style="list-style-type: none"> • An acid was a substance that produced hydrogen ions when dissolved in water. • If they were strong acids, they completely ionised. • If they were weak acids, they only partially ionised. • A base was a substance that produced hydroxide ions when dissolved in water.
<ul style="list-style-type: none"> • Outline the Bronsted-Lowry theory of acids and bases. 	<p>In 1923, Johannes Bronsted and Thomas Lowry independently realised that acid-base reactions involved proton transfer.</p> <p>The Theory: An acid is a substance that, in solution, tends to give up protons, and a base is a substance that tends to accept protons.</p> <ul style="list-style-type: none"> • An acid is a proton donor • A base is a proton acceptor
<ul style="list-style-type: none"> • Describe the relationship between an acid and its conjugate base and a base and its conjugate acid. 	<p>When an acid donates a proton, it forms what is called its conjugate base.</p> <ul style="list-style-type: none"> • If the acid is strong, the conjugate base will be weak. • If the acid is weak, the conjugate base will be strong. <p>When a base accepts a proton, it forms its conjugate acid.</p> <ul style="list-style-type: none"> • If the base is strong, the conjugate acid will be weak. • If the base is weak, the conjugate acid will be strong.
<ul style="list-style-type: none"> • Identify a range of salts, which form acidic, basic or neutral solutions and explain their acidic, basic or neutral nature. 	<p>There are four types of reactions when dealing with the hydrolysis of salts.</p> <p>Weak Acid and a Strong Base: In a salt formed from a weak acid and a strong base, the anion is a weak base and so the salt in aqueous solution has pH greater than 7.</p> <p>Strong Acid and a Weak Base: In a salt formed from a weak base and a strong acid, the cation is a weak acid and so the salt in aqueous solution has a pH less than 7.</p> <p>Strong Acid and a Strong Base: The salt produced has a pH close to 7.0. This is because the conjugate base of a strong acid does not react with water. Similarly, the conjugate acid of a strong base does not react with water.</p>

	<p>Weak Acid and a Weak Base: The salt produced has a pH close to 7. This is because the anion and cation react with water to equal extents and therefore approximately cancel each other out.</p> <p>Rules to Remember: Weak Acid + Strong Base \longrightarrow Basic Salt + Water Strong Acid + Weak Base \longrightarrow Acidic Salt + Water Strong Acid + Strong Base \longrightarrow Neutral Salt + Water Weak Acid + Weak Base \longrightarrow Neutral Salt + Water</p>
<ul style="list-style-type: none"> Identify conjugate acid/base pairs. 	<p>Examples: $\text{HCl}_{(Acid)} + \text{H}_2\text{O} \longrightarrow \text{Cl}^-_{(Base)} + \text{H}_3\text{O}^+$ $\text{HCl} + \text{H}_2\text{O}_{(Base)} \longrightarrow \text{Cl}^- + \text{H}_3\text{O}^+_{(Acid)}$</p>
<ul style="list-style-type: none"> Identify amphiprotic substances and construct equations to describe their behaviour in acidic and basic solutions. 	<p>A substance that can act as a proton donor (acid) and also as a proton acceptor (base) is called an amphiprotic substance.</p> <p>Water is an example of an amphiprotic molecule: $\text{H}_2\text{O} \longrightarrow \text{H}^+ + \text{OH}^-$ $\text{H}_2\text{O} + \text{H}^+ \longrightarrow \text{H}_3\text{O}^+$</p> <p>Hydrogen Carbonate is another example: $\text{HCO}_3^- \longrightarrow \text{H}^+ + \text{CO}_3^{2-}$ $\text{H}^+ + \text{HCO}_3^- \longrightarrow \text{H}_2\text{CO}_3$</p>
<ul style="list-style-type: none"> Identify neutralisation as a proton transfer reaction, which is exothermic. 	<p>All neutralisation reactions involve proton transfer. Therefore, they are regarded as proton transfer reactions.</p> <p>Neutralisation reactions are exothermic.</p>
<ul style="list-style-type: none"> Describe the correct technique for conducting standard titration and preparation of standard solution. 	<p>Preparation of Glassware:</p> <ul style="list-style-type: none"> Wash the pipette with distilled water, and then rinse with the solution that will be placed in there. Wash the burette with distilled water, and then rinse with the solution that will be placed in there. <p>Preparation of Standard Solution:</p> <ol style="list-style-type: none"> Accurately measure and record the mass of a solid. Dissolving completely in distilled water. Transferring the entire dissolved solid to a volumetric flask. Adding water to the flask until an exact known volume is reached. Calculate the molarity of that solution. <p>Titration Procedure:</p> <ol style="list-style-type: none"> Use a pipette to measure a volume of one solution. Transfer this to a conical flask. Add a few drops of a suitable indicator.

	<p>4. Add the second solution slowly from a burette into the conical flask, swirling continuously, until the indicator changes colour. Record the amount of second solution used.</p> <p>5. Repeat until three readings are obtained within 0.1ml of each other.</p> <p>6. Perform calculations.</p> <p>Definitions: Titration – Procedure used to experimentally find the concentration of a solution. Equivalence point – Point when reaction is complete. End point – Point when indicator changes colour. Pipette – Titration equipment which measures a fixed volume of solution to provide fixed number of moles of one reactant. Titre – Volume of solution delivered from the burette. Titrant – Solution of known concentration and composition. Aliquot – Fixed volume of solution delivered by pipette.</p>
<ul style="list-style-type: none"> • Qualitatively describe the effect of buffers with reference to a specific example in a natural system. 	<p>A buffer is a solution that maintains a fairly constant pH when small amounts of acids/bases are added.</p> <p>If an acid is added to a buffer, the hydrogen ions are removed: $\text{H}^+ + \text{HCO}_3^- \longrightarrow \text{H}_2\text{CO}_3$</p> <p>If a base is added, hydroxide ions are removed. $\text{OH}^- + \text{HCO}_3^- \longrightarrow \text{H}_2\text{O} + \text{CO}_3^{2-}$</p> <p>Natural System: Hydrogen carbonate is a natural buffer in our bloodstream that helps to keep our blood at a pH of about 7.4.</p>
<ul style="list-style-type: none"> • Gather and process information from secondary sources to trace developments in understanding and describing acid/base reactions. 	<p>Antoine Lavoisier (1790):</p> <ul style="list-style-type: none"> • An acid was a substance that must contain oxygen. • This was later disproved as acids such as HCl does not include oxygen. <p>Humphry Davy (1815):</p> <ul style="list-style-type: none"> • An acid was a substance that contained hydrogen. These hydrogen atoms could be replaced by reaction with a metal. • Helped to classify substances but didn't interpret its properties. <p>Svante Arrhenius (1884):</p> <ul style="list-style-type: none"> • An acid was a substance that produced hydrogen ions when dissolved in water. • If they were strong acids, they completely ionised. • If they were weak acids, they only partially ionised. • A base was a substance that produced hydroxide ions when dissolved in water.

Bronsted-Lowry (1923):

- An **acid** is a substance that, in solution, tends to **give up protons**, and a **base** is a substance that tends to **accept protons**.
 - An **acid** is a **proton donor**
 - A **base** is a **proton acceptor**
- This concept provided a sound basis for the quantitative treatment of acid-base equilibria and pH calculations.

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

Experiment:Aim:

To determine the pH of various salt solutions.

Method:

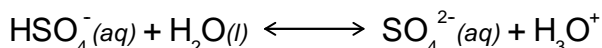
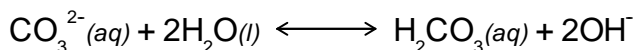
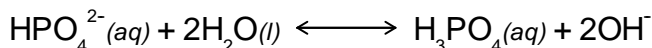
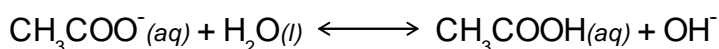
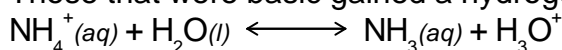
1. Place a small amount of each salt solution into separate beakers.
2. Use a pH probe to record the pH of each solution.
3. Use Universal Indicator solution and chart to determine pH of each solution.

Results:

Salt	pH probe	UI pH
NH ₄ Cl	5.9	5
KNO ₃	6.4	7
CH ₃ COONa	8.1	8
Na ₂ HPO ₄	8.2	8
Na ₂ CO ₃	9.8	10
NaCl	6.5	7
NaHSO ₄	3.2	= 3

Conclusion:

The salt solutions that were acidic lost a hydrogen atom. Those that were basic gained a hydrogen atom.



- **Perform a first-hand investigation and solve problems using titrations and including the preparation of standard solutions, and use available**

Experiment:Aim:

To prepare a primary standard solution of sodium carbonate and use it to determine the concentration of a hydrochloric acid solution.

Method:

1. Put approximately 2g of sodium carbonate into a clean and tared beaker. Record the exact mass of the sodium

evidence to quantitatively and qualitatively describe the reaction between selected acids and bases.

- carbonate.
- Add about 30ml of distilled water to the beaker and stir. Place the funnel into a clean volumetric flask and transfer the sodium carbonate solution into the flask. Use a wash bottle filled with distilled water to carefully wash all the sodium carbonate down the funnel and into the flask. Swirl the beaker so as to dissolve the sodium carbonate.
 - Rinse the beaker with distilled water and transfer wash water into the flask.
 - Add distilled water to the flask until it is about two-thirds full.
 - Fit the stopper and shake until all the sodium carbonate has dissolved.
 - Remove the stopper and fill the flask with distilled water until it is perfectly on the 250ml mark.
 - Rinse the burette with distilled water and then with the HCl solution to be used and discard the rinsings.
 - Set up the burette using a retort stand and burette clamp.
 - Ensure the stopcock is closed and fill the burette with the HCl solution. Record the starting volume.
 - Rinse the pipette with the sodium carbonate solution and discard the rinsings.
 - Fill the pipette with Na_2CO_3 solution exactly to the 25ml mark and place in a clean 250ml conical flask. Add 2-3 drops of indicator.
 - Fill a small beaker with approximately 25ml of the Na_2CO_3 solution and add a few drops of indicator.
 - Place the conical flask under the burette and the small beaker next to it to act as a guide.
 - Run the HCl into the flask swirling continuously until the colour changes from that of the guide. This is the end point.
 - Record the end volume.
 - Calculate the volume of HCl used and record.
 - Refill the burette and repeat steps 9 to 16 four more times.

Results:

Mass of Na_2CO_3 used: 2.01 g

	Trial 1	Trial 2	Trial 3	Trial 4	Trial 5
Volume Na_2CO_3 used	25 ml	25 ml	25 ml	25 ml	25 ml
Initial burette reading	2.7 ml	3.6 ml	0.4 ml	2.9 ml	2.2 ml
Final burette reading	46.7 ml	20.4 ml	35.0 ml	36.4 ml	25.6 ml
Volume of HCl used	44.0 ml	36.8 ml	34.6 ml	33.5 ml	33.4 ml

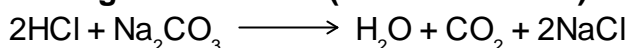
Calculations:

The results from Trial 1 were discarded as they were extremely inconsistent with the rest of the tests and it was the first attempt.

Using:

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

Average of HCl used (Trial 2 to Trial 5): 34.575 ml



$$\begin{aligned} \text{Number of Moles} &= \frac{2.01}{105.99} \\ &= 0.018964053 \text{ moles} \end{aligned}$$

$$\begin{aligned} \text{Molarity of Solution} &= \frac{\text{Moles}}{\text{Volume}} \\ &= \frac{0.018964053}{0.250} \\ &= 0.075856212 \text{ mol L}^{-1} \end{aligned}$$

$$\begin{aligned} \text{Number of moles used} &= \text{Molarity} \times \text{Volume} \\ &= 0.075856212 \times 0.025 \\ &= 1.896405321 \times 10^{-3} \text{ Moles} \end{aligned}$$

$$\begin{aligned} \text{Molarity of HCl} &= \frac{2 \times (1.896405321 \times 10^{-3})}{0.034575} \\ &= 0.1097 \text{ Mol L}^{-1} \end{aligned}$$

Conclusion:

The molarity of the Hydrochloric Acid solution was 0.1097 mol L⁻¹.

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

Experiment:

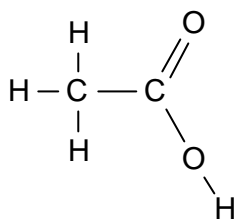
Aim:

To determine the concentration of acetic acid using a pH probe data logger.

Method:

1. Prepare standard solution and set up experiment the same way as for a regular titration.
2. Instead of stirring the solution when the titrant is being added, use a magnetic stirrer.
3. Instead of recording starting volume and ending volume, perform a flow rate analysis to determine the millilitres per second flow rate.
4. Use the data logger to record the pH in a graph and start the recording.
5. Once the reaction is complete, calculate the time taken for the reaction to reach its end point and multiply that by the calculated flow rate. This will give you the volume of titrant needed to reach the equivalence point.
6. From then on, do all the calculations the same way as

	<p>done for a standard titration.</p> <p><u>Conclusion:</u> The calculations will still return a concentration in mol L⁻¹.</p>
<ul style="list-style-type: none"> 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. 	<p>Neutralisation reactions are widely used for safety in laboratories and factories where acids or bases are used. Any spills can be neutralised by a suitable acid/base.</p> <p>When selecting an appropriate neutralising reagent, it is important to consider the following:</p> <ul style="list-style-type: none"> The speed of the neutralisation reaction The need for a reagent that will not have harmful if excess is added. The safety of handling and storing. The cost The possibility of using the same substance for acid and base spills. Hence, an amphoteric substance. <p>An example of a common neutralising reagent is sodium carbonate.</p>
<p>Focus 5 Esterification is a naturally occurring process which can be performed in the laboratory.</p>	
<ul style="list-style-type: none"> Describe the difference between the alkanol and alkanolic acid functional groups in carbon compounds. 	<p>Alkanols</p> <ul style="list-style-type: none"> An alkanol is a compound which contains an OH group attached to a carbon atom that only has carbon or hydrogen atoms attached to it, regardless of what else is in the molecule. Alkanols contain the hydroxy functional group (OH) High melting and boiling points Polar molecules An example is ethanol $ \begin{array}{c} \text{H} \\ \\ \text{H} - \text{C} - \text{O} - \text{H} \\ \\ \text{H} \end{array} $ <p>Alkanolic Acids:</p> <ul style="list-style-type: none"> Alkanolic acids contain the carboxylic functional group (COOH) Polar molecules An example is ethanoic acid



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

Rules for Naming Esters:

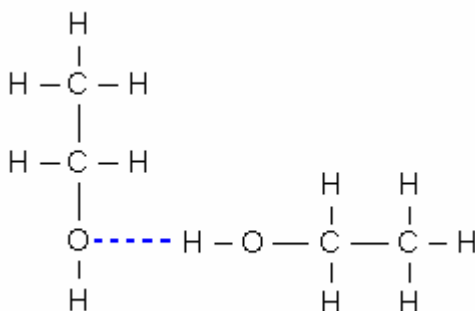
- The first part of the name comes from the **alkanol**. This is named the as an **alkyl group**.
- The second part of the name comes from the **alkanoic acid**. When naming the **ester**, the suffix '-oic acid' is replaced by '**anoate**.'

Carbons	First Part	Second Part
1	Methyl	Methanoate
2	Ethyl	Ethanoate
3	Propyl	Propanoate
4	Butyl	Butanoate
5	Pentyl	Pentanoate
6	Hexyl	Hexanoate
7	Heptyl	Heptanoate
8	Octyl	Octanoate

- Explain the difference in melting point and boiling point caused by straight-chained alkanolic acids and straight-chained primary alkanol structures.

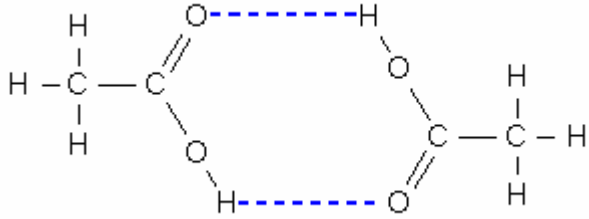

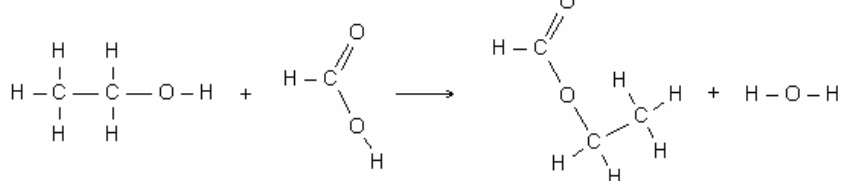
Straight-Chained Alkanols:

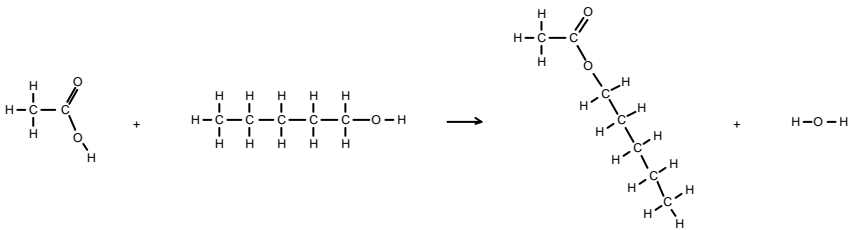
- Have high melting and boiling points due to hydrogen bonding between the O in one molecule and the H of a hydroxide group in a nearby molecule.
- Strong intermolecular forces.
- Are of polar nature.



Alkanoic Acids:

- Have even higher melting and boiling points than alkanols due to the even stronger polar nature.
- Are of polar nature
- Have stronger intermolecular forces.
- Involves two pairs of hydrogen bonding.

	
<ul style="list-style-type: none"> Identify esterification as the reaction between an acid and an alkanol and describe, using equations, examples of esterification. 	<p>The reaction between an alkanol and an alkanoic acid is called esterification.</p> $R_1COOH + OHR_2 \longrightarrow R_1COOR_2 + H_2O$  <p>Example:</p> $C_2H_5OH + HCOOH \longrightarrow HCOOC_2H_5 + H_2O$ 
<ul style="list-style-type: none"> Describe the purpose of using an acid in esterification for catalysis. 	<p>Esterification is a relatively slow process at room temperature and does not go to completion, it comes to equilibrium.</p> <p>To combat this, a catalyst, concentrated sulfuric acid, is used to absorb the product water and move the equilibrium to the right.</p>
<ul style="list-style-type: none"> Explain the need for refluxing during esterification. 	<p>Refluxing is the process of heating a reaction mixture in a vessel with a cooling condenser attached in order to prevent loss of any volatile reactant or product.</p> <p>This stops any alcohol being lost due to evaporation.</p>
<ul style="list-style-type: none"> Outline some examples of the occurrence, production and uses of esters. 	<p>Esters are compounds formed when alkanoic acids react with alkanols, or more generally, when carboxylic acids combine with alcohols.</p> <p>Esters have pleasant, fruity odours and occur widely in nature as perfumes and flavouring agents.</p> <p>Some common ones are listed below.</p>

	<ul style="list-style-type: none"> • Methyl butanoate and iso-pentyl pentanoate (Apple) • Iso-pentyl ethanoate (Banana) • Octyl thanoate (Orange) • Pentyl ethanoate (Pear) <p>Uses of Esters:</p> <ul style="list-style-type: none"> • Food flavourings • Nail polish remover • Plasticisers
<ul style="list-style-type: none"> • Identify data, plan and perform a first hand investigation to prepare and ester using reflux. 	<p><u>Experiment:</u> <u>Aim:</u> To prepare an ester using reflux. <u>Method:</u></p> <ol style="list-style-type: none"> 1. Place 10ml of 1-Pentanol, 12ml of glacial acetic acid and 1ml of concentrated sulfuric acid in the 50ml spherical flask. 2. Add a few boiling chips and assemble the refluxing apparatus. Clamp the flask and condenser securely to the retort stand. 3. Filled a large beaker with boiling water and used it as water bath to stop flammable substances falling into the flame of the Bunsen burner. 4. Connected the tubing to the tap and condenser, and turned on the water until a uniform flow is achieved. 5. Heat the mixture over a steady flame for approximately 40 minutes and allow cooling for 5 minutes. Turn off the water. Leave overnight. 6. The following day, reheat the mixture for a further ten minutes and allow to cool again. 7. Carefully remove the flask and pour the contents into a separating funnel containing 10-15ml of water. 8. Put the stopper on the funnel and shake. 9. Allow the layers to separate, drain off and discard the lower aqueous layer. 10. Add approximately 15ml of 1 mol L⁻¹ sodium carbonate solution. Shake, drain and discard the lower layer. 11. Smell the ester and describe. <p><u>Results:</u> The ester smelt of banana. The ester, 1-Pentyl ethanoate, was produced.</p> 
<ul style="list-style-type: none"> • Process information from 	<p>Esters are the basis of many commercial food flavourings. We can only taste sweet, sour, salty, or bitter flavours with</p>

secondary sources to identify and describe the uses of esters as flavours and perfumes in processed foods and cosmetics.

our tongues. The special nuances we associate with certain foods come from their complex aromas. These aromas are the result of **esters**, which can occur naturally, or can be food additives.

Ester	Uses/Properties
Propyl hydroxybenzoate	<ul style="list-style-type: none"> • The ester is soluble in water making it a useful ester. • Used in preservatives • Used in toothpaste • Skin Basics Sorbolene Cream • 0.8% content in Lice cream
Butyl phthalate	<ul style="list-style-type: none"> • Used in oils that are part of manometre gauges to measure pressure.
Octyl Methoxycinnamate	<ul style="list-style-type: none"> • Used in daily UV lotions • Used in sunscreens • Absorbs UV radiation by becoming absorbed by the skin, compared to other sunscreens which only have to sit on top to work.
Triethyl citrate	<ul style="list-style-type: none"> • Used as a direct food additive for: <ul style="list-style-type: none"> ○ Flavouring agent ○ Solvent vehicle ○ Surface active agent ○ Whipping agent • Deodorant agent • Citrus oils • Is not found naturally
Octyl gallate	<ul style="list-style-type: none"> • Used as an antioxidant • Used in: <ul style="list-style-type: none"> ○ Cosmetics ○ Margarine & Peanut butter ○ Pharmaceuticals • Added to fats and oils to stop going off.