

**I. Indicators were identified with the observation that the colour of some flowers depend soil composition**

1) Classify some common substances as acidic, basic, or neutral

Acidic	Basic	Neutral
Vinegar	Cloudy ammonia	Water
Lemon juice	Washing soda	Salt
Aspirin	Antacid tablets	Sugar
Vitamin C	Oven cleaners	

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

pH	Colour Change				
	0	4	7	10	14
Indicator	Very acidic	Slightly acidic	Neutral	Slightly basic	Very acidic
Methyl orange	red →	yellow	yellow	yellow	yellow
Bromothymol blue	yellow	yellow	→	blue	blue
Litmus	red	red	→	blue	blue
Phenolphthalein	colourless	colourless	colourless	Colourless →	red

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

Indicators can be used to ascertain the pH of soils. First a non-reactive white powder such as barium sulfate is added to the sample of soil to aid in visualising the colour change. Then, universal indicator is added. The colour change can then be matched against a colour chart, and the pH of the soil determined.

Another use for indicators is to determine the pH of pool water. Pool water is ideally pH neutral, so litmus paper can be used to check if it is either too acidic or too basic. If the pool water turns red litmus blue, then it is alkaline. If it turns blue litmus red, then it is acidic.

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

See Prac Book (2.1)

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

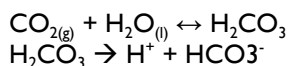
Indicator	pH Range	Colour Range
Methyl orange	3.1 → 4.4	Red → yellow
Methyl red	4.4 → 6.0	Pink → yellow
Bromothymol blue	6.2 → 7.6	Yellow → blue
Litmus	6.2 → 7.4	Red → blue
Phenol red	6.8 → 8.4	Yellow → red
Phenolphthalein	8.3 → 10.0	Colourless → red

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

**2. While we usually think of the air around us as neutral, the atmosphere naturally contains acidic oxides of carbon, nitrogen and sulfur. The concentrations of these acidic oxides have been increasing since the Industrial Revolution**

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

Carbon dioxide, sulfur dioxide and nitrogen dioxide all dissolve in water forming acidic solutions. Most non-metal oxides (except for CO, NO, and N<sub>2</sub>O, which are neutral) form an acidic solution with water. These substances act as acids when they ionise in water.



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

From the periodic table, it can be seen that:

- Most metal oxides are basic
- Most non-metal oxides are acidic

Five elements close to the borderline between metals and non-metals form oxides that are amphoteric. That is, they are neither exclusively acidic nor basic, but will react with either an acid or a base, depending on the substances that are reacted with them.

- 3) Define Le Chatelier's principle

If a system at equilibrium is disturbed, then the system adjusts itself so as to minimise the disturbance.

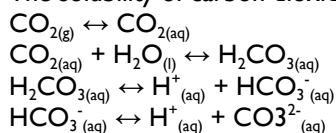
- 4) Identify factors which can affect the equilibrium in a reversible reaction

Factors which can affect the equilibrium in a reversible reaction are:

- Change in concentration of products or reactants
- Change in temperature
- Change in gas pressure

- 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

The solubility of carbon dioxide gas in water can be fully described using four equations



An equilibrium shift to the left will produce carbon dioxide gas. An equilibrium shift to the right will dissolve carbon dioxide gas.

Le Chatelier's principle predicts that:

Addition of acid (i.e. increase [H<sup>+</sup>])

Equilibrium will shift to oppose the change

i.e. lower [H<sup>+</sup>]

The reaction that reduces [H<sup>+</sup>] is the reverse reaction

∴ Equilibrium will shift to the left (carbon dioxide gas produced)

Addition of a base (i.e. reacts with and reduces [H<sup>+</sup>])

Equilibrium will shift to oppose the change

i.e. increase [H<sup>+</sup>]

The reaction that increases [H<sup>+</sup>] is the forward reaction

∴ Equilibrium will shift to the right

Addition of a soluble carbonate (i.e. increase [CO<sub>3</sub><sup>2-</sup>])

Equilibrium will shift to oppose the change

i.e. decrease  $[\text{CO}_3^{2-}]$

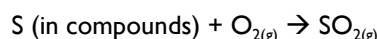
The reaction that decreases  $[\text{CO}_3^{2-}]$  is the reverse reaction

$\therefore$  Equilibrium will shift to the left

- 6) Identify natural and industrial sources of sulfur dioxide and oxides of nitrogen
- 7) Describe using equations, examples of chemical reactions which release sulfur dioxide and chemical reaction which release oxides of nitrogen

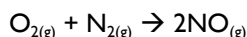
About two-thirds of the sulfur dioxide released to the atmosphere comes from natural sources such as geothermal hot springs and volcanoes. Since volcanic activity varies greatly, atmospheric concentrations of sulfur dioxide also vary quite widely.

The main human activities that release sulfur dioxide to the atmosphere are processing or burning fossil fuels and extracting metals from sulphide ores.

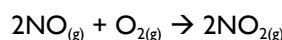


There are three common oxides of nitrogen, nitrous oxide ( $\text{N}_2\text{O}$ ), nitric oxide ( $\text{NO}$ ), and nitrogen dioxide,  $\text{NO}_2$ .  $\text{NO}_2$  is acidic, whilst the other two are neutral.

The major natural source of nitric oxide is lightning. At the very high level localised temperatures generated by lightning, atmospheric oxygen and nitrogen gases combine to form nitric oxide



Nitric oxide slowly reacts with oxygen to form nitrogen dioxide



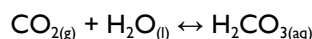
Industrial activity such as combustion in power stations and in car engines generates large amount of nitric oxide and nitrogen dioxide. At the high temperatures in engines and generators, oxygen and nitrogen combine to form nitric oxide, and again, nitric oxide is slowly converted to nitrogen dioxide.

- 8) Asses the evidence which indicates increases in atmospheric concentrations of oxides of sulfur and nitrogen

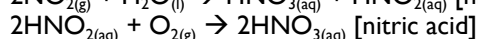
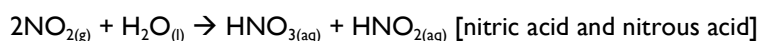
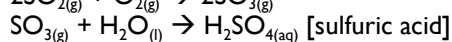
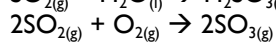
Since  $\text{SO}_2$  and  $\text{NO}_2$  are water-soluble and therefore washed out of the atmosphere by rain, there appears to be little significant build-up of their concentrations over the last 100 years (unlike  $\text{CO}_2$  which has seen a 30% rise over the last century). However, it is difficult to be sure about  $\text{SO}_2$  and  $\text{NO}_2$ , because there is a lack of data for periods before the 1950s. It has only been in the last few decades that we have been able to measure concentrations of these gases below 0.1 ppm with sufficient accuracy.

- 9) Calculate volumes of gases given masses of some substances in reaction, and calculate masses of substances given gaseous volumes, in reactions involving gases at  $0^\circ\text{C}$  and 101.3 kPa or  $25^\circ\text{C}$  and 101.3 kPa
- 10) Explain the formation and effects of acid rain

Distilled water in contact with the atmosphere is not pH neutral. It has a pH of roughly 5.5, due to the absorption of carbon dioxide forming acidic carbon acid.



In Australia unpolluted rainwater has a pH between 5 and 6. If the pH is below 5, an acidic substance, such as sulfur dioxide or nitrogen dioxide has dissolved in the water, which is consequently termed acid rain. In the northern hemisphere, pHs as low as 2 have been recorded. The source of the  $\text{SO}_2$  or  $\text{NO}_2$  could be thousands of kilometres from where the acid rain actually falls.



If the quantity of acid rain is greater than the capacity of an environment to neutralise it, then the following can occur:

- Soil pH can drop, making it difficult for plants to absorb sufficient mineral cations
- Soil chemistry can change, leading to death of important micro-organisms, and release of normally insoluble aluminium and mercury into soil water
- Protective waxes can be lost from leaves, causing extensive leaf damage
- Buildings constructed from carbonates, such as concrete, mortar, limestone and marble, can be gradually dissolved away
- Aquatic animals can die as water pH drops below 5
- Smog and acid rain can combine to form killer fog, as happened after the second world war in London, when many homes burnt sulfur dioxide releasing coal

- 11) Identify data, plan and perform a first-hand investigation to decarbonate soft drink and gather data to measure the mass changes involved and calculate the volume of gas released at 25°C and 101.3 kPa

See Prac Book (2.2)

- 12) Analyse information from secondary sources to summarise the industrial origins of the above gases and evaluate reasons for concern about their release into the environment

After the industrial revolution in the 1800s, there was a great increase in emissions of SO<sub>2</sub> to the atmosphere surrounding industrial cities, mainly from burning coal and extracting metals. Air quality significantly deteriorated until the 1950s, when pollution-related deaths started occurring. This led to the introduction of regulations to restrict emissions from factories. The legislation eventually resulted in air quality improvement.

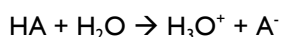
Serious pollution from NO<sub>x</sub> did not appear until the 20<sup>th</sup> century as electricity generation and motorcar usage expanded dramatically. Photochemical smog first became a problem in the 1960s in Los Angeles. This led to the introduction of emissions controls on motor vehicles. As a result improvements in air quality were measured for about two decades. However increasing city populations and increasing vehicle-kilometres being driven have to a large extent nullified further benefits of increasingly stringent emissions controls. In most major cities, air quality has remained largely stagnant over the past decade.

The annual average concentration of SO<sub>2</sub> and NO<sub>2</sub> in most large cities around the world is about 0.01 ppm for each gas. This is about 10 times the value for clean air, though a concentration of 0.01 ppm is not harmful. Of more concern is the number of days per year on which the concentration of one of these pollutants exceeds what health authorities regard as safe. In Sydney and Melbourne while the annual concentrations of these gases is suitably low, there is some concern about the number of days per year that safe levels are exceeded.

### 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 proton donor. When an acid molecule is in contact with water it can ionise, donating a proton to a water molecule.



A hydrogen atom, H, consists of one proton and one electron. A hydrogen ion, H<sup>+</sup>, is formed when the H atom loses its electron, leaving just a proton. A proton and a hydrogen ion are thus the same and be represented as H<sup>+</sup>.

When an acid molecule is placed in water, it can ionise, releasing a proton and forming a negative ion. The proton can attach itself to water forming H<sub>3</sub>O<sup>+</sup>.

- 2) Identify acids including acetic, citric, hydrochloric and sulfuric acid

- Acetic acid (ethanoic acid, vinegar) – CH<sub>3</sub>COOH
- Citric acid – C<sub>6</sub>H<sub>8</sub>O<sub>7</sub> (2-hydroxypropane-1,2,3-tricarboxylic acid)
- Hydrochloric acid – HCl
- Sulfuric acid – H<sub>2</sub>SO<sub>4</sub>

- 3) Describe the use of the pH in comparing acids and bases

The pH scale is used to compare the concentration of hydrogen ions  $[H^+]$  in solutions of acids and bases.

$$\begin{aligned} \text{pH} &= -\log[H^+] \\ \text{pOH} &= -\log[OH^-] \\ \text{pH} + \text{pOH} &= 14 \end{aligned}$$

A pH of 7 is neutral, where the concentration of  $H^+$  and  $OH^-$  is equal

pH < 7; acidic  
pH > 7; basic

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

A strong acid fully ionises into  $H^+$  ions and its conjugate base in solution.

e.g.  $HCl \rightarrow H^+ + Cl^-$

A weak acid does not fully ionise in solution.

e.g.  $CH_3COOH \leftrightarrow H^+ + CH_3COO^-$

A concentrated acid has a relatively high number of acid molecules per volume of solution.

e.g. 4M HCl

A dilute acid has a relatively low number of acid molecules per volume of solution

e.g. 0.01M HCl

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

'p' means  $-\log_{10}$   
 $\therefore \text{pH} = -\log_{10}[H^+]$

If a substance has a molarity of 1M, it has a pH of 1

If a substance has a molarity of 0.1M, it has a pH of 2

$\therefore$  a change in pH of 1, is a ten-fold change in concentration of hydrogen ions

- 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

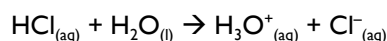
Hydrochloric acid is a strong acid with 100% ionisation in solution.

Citric acid and acetic acid are weak acids, with typically only 1% ionisation in solution.

Thus a similar concentration of the acid would yield different pHs.

- 7) Describe the difference between a strong and a weak acid in terms of equilibrium between the intact molecule and its ions

In a strong acid, such as hydrochloric acid, an equilibrium is formed during ionisation:



In the equilibrium of the strong acid, the equation completely lies on the right side (near 100% ionisation). The molecule of the strong acid completely ionises.

In a weak acid, such as acetic acid, an equilibrium is formed during ionisation:



In the equilibrium of the weak acid, the equation lies mostly on the left (partial ionisation). The molecule of the weak acid is in solution with few of its ions.

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

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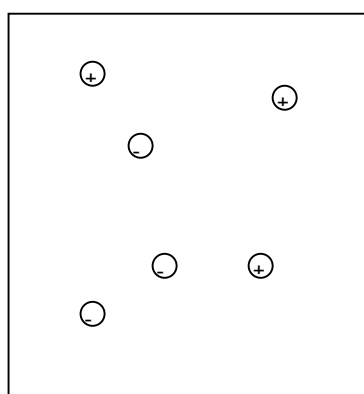
- 9) Plan and perform a first-hand investigation to measure the pH of identical concentrations of strong and weak acids

See Prac Book (2.3)

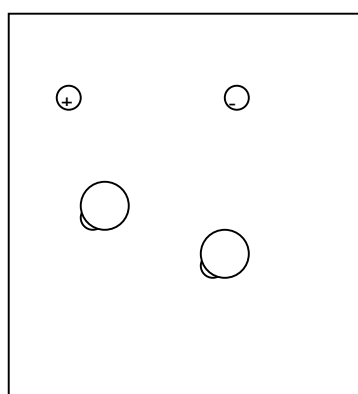
- 10) Gather and process information from secondary sources to write ionic equations to represent the ionisation of acids

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



Strong acid e.g. HCl



Weak acid e.g.  $\text{CH}_3\text{COOH}$

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

Sulfur dioxide can be added to foods as gaseous  $\text{SO}_2$ , as a solution in water, or as a sulfate, bisulfate or metabisulfite salt.

A food acid is any acidic substance not normally consumed as a food itself and not normally used as an ingredient in food, but is usually added to food to preserve it. That is, to prevent micro-organisms from decomposing food. The micro-organisms cannot survive in the acidic conditions caused by the food additives such as sulfur dioxide.

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

name	composition	acid/base	pH
Hydrochloric acid	HCl	Acid	~2
Citric acid	C <sub>6</sub> H <sub>8</sub> O <sub>7</sub>	Acid	~3
Sodium hydroxide	NaOH	Base	14
Sodium oxide	Na <sub>2</sub> O	Base	

- 14) Process information from secondary sources to calculate pH of strong acids given appropriate hydrogen ion concentration

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#### 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, Davy and Arrhenius

Antoine Lavoisier, in about 1780, proposed that acids were substances that contained oxygen: in fact the name 'oxygen' derives from Greek words meaning 'produces sour taste'. However this theory was soon disproved: many oxygen-containing substances were basic (such as sodium and calcium oxides) and some acidic substances (such as hydrochloric acid-called muriatic acid at that time) were distinctly acidic but contained no oxygen.

Humphry Davy in 1815 suggested that acids were substances that contained replaceable hydrogen-hydrogen that could be partly or totally replaced by metals. When acids reacted with metals they formed salts (such as ZnCl<sub>2</sub>, FeSO<sub>4</sub>, NaNO<sub>3</sub>). Bases were substances that reacted with acids to form salts and water. These definitions worked quite well for most of that century.

Svante Arrhenius in 1884 proposed that an acid was a substance which ionised in solution to produce hydrogen ions: acids were strong if they ionised almost completely (hydrochloric, nitric), and weak if they ionised only slightly (acetic, formic). Arrhenius defined a base as a substance that in solution produced hydroxide ions. This definition is somewhat narrow in that it excludes metallic oxides that are distinctly basic.

However there are some inadequacies with the Arrhenius definition of an acid. First, it does not give due recognition to the role of the solvent. Ionisation of an acid is not something the acid does in isolation: rather it is a reaction between the acid molecule and the solvent. Whether an acid is strong or weak depends not only upon the nature of the acid itself, but also upon the nature of the solvent it is dissolved in: hydrochloric acid in water is a strong acid, but when dissolved in diethyl ether it is quite weak. Second, acid-base reactions often occur in solvents in which the 'acid' is not ionised at all: hydrochloric acid, when dissolved in benzene, is not ionised yet it still reacts with ammonia to form ammonium chloride. Hydrochloric acid is still reacting with a base to form a salt, although the acid does not ionise in that solvent.

- 2) Outline the Bronsted-Lowry theory of acids and bases

In 1923 two chemists, Lowry from Britain and Bronsted from Denmark, independently proposed new definitions for acids and bases in terms of proton donors and acceptors.

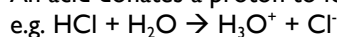
An acid is a substance, which in solution, tends to give up protons, and a base is a substance which tends to accept protons.

*An acid is a proton donor  
A base is a proton acceptor*

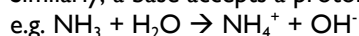
The Bronsted-Lowry concept relates acidity and basicity to the structure of both the substance and the solvent. Production of hydrogen ions is seen to be due not solely to the properties of the acid, but rather to the properties of the acid relative to those of the solvent.

3) Describe the relationship between an acid and its conjugate base, and a base and its conjugate acid

An acid donates a proton to form what is called its conjugate base

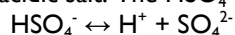


Similarly, a base accepts a proton to form what is called its conjugate acid

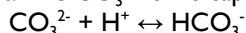


4) Identify a range of salts which form acidic, basic or neutral solutions and explain their acidic neutral or basic nature

Sodium hydrogen sulfate ( $\text{NaHSO}_4$ ) is an acidic salt. The  $\text{HSO}_4^-$  ion is capable of donating a proton.



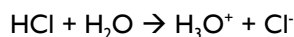
Sodium carbonate ( $\text{Na}_2\text{CO}_3$ ) is a basic salt. The  $\text{CO}_3^{2-}$  ion is capable of accepting a proton



Sodium chloride ( $\text{NaCl}$ ) is a neutral salt. Neither ion is capable of accepting nor donating an ion.

5) Identify conjugate acid/base pairs

Whenever an acid and a base react, they form their conjugates.



HCl – acid I

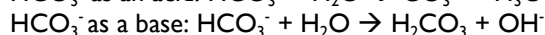
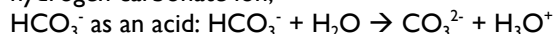
$\text{H}_2\text{O}$  – base I

$\text{H}_3\text{O}^+$  - conjugate acid of base I

$\text{Cl}^-$  - conjugate base of acid I

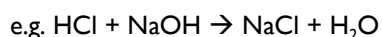
6) Identify amphiprotic substance and construct equations to describe their behaviour in acidic and basic solutions

A substance which can act as both a proton donor and a proton acceptor is said to be amphiprotic. It can thus act as both an acid or a base depending on the solution it is placed in. an example is the hydrogen carbonate ion,



7) Identify neutralisation as a proton transfer reaction which is exothermic

Neutralisation is a proton transfer reaction which is exothermic (!)



The proton is donated by the chloride ion to the hydroxide ion forming a water molecule.

8) Describe the correct technique for conducting titrations and preparations of standard solutions

A solution of accurately known concentration is called a standard solution.

For a chemical to be suitable to prepare as a standard solution, it must:

- Be a water soluble solid
- Have high purity - usually Analytical Reagent (A.R.) grade
- Have an accurately known formula
- Be stable in air, i.e. it does not lose or gain water or react with oxygen or carbon dioxide in air.

The solution is prepared by:

1. Accurately weighing a calculated amount of solid
2. Dissolving it in water
3. Transferring all of the dissolved solid to a volumetric flask
4. Adding water to the flask to prepare a fixed volume of solution.

The concentration is usually calculated in  $\text{mol L}^{-1}$ .

A standard solution can be reacted with a solution of unknown concentration using titration technique. One reactant in solution is slowly added to another reactant in solution until an end point is reached.

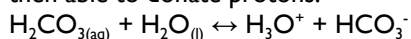
The end point of the titration is usually indicated by a change in colour of a small amount of indicator solution added to the mixture of reactants. For an acid-base titration an indicator is selected that changes colour at the pH of the salt solution formed at the point of neutralisation. This is known as the equivalence point.

At senior high school level equipment such as burettes, pipettes and volumetric flasks give readings to three significant figures. Calculations are carried out to three significant figures.

9) Qualitatively describe the effects of buffers with reference to a specific example in a natural system

A buffer solution contains comparable amounts of a weak acid and its conjugate base. Thus, it is able to maintain an approximately constant pH even when significant amounts of strong acid or strong base are added to it.

Since many processes in nature require carefully controlled pH, there are many natural buffer systems. The simplest one involves carbon dioxide gas and hydrogen carbonate in human blood. Carbon dioxide gas forms carbonic acid in human blood by combining with water to form dihydrogen carbonate, which is then able to donate protons.



If acid is added it would react with  $\text{HCO}_3^-$  forming  $\text{H}_2\text{CO}_3$ , pushing equilibrium to the right and reforming  $\text{HCO}_3^-$  and  $\text{H}_3\text{O}^+$  ions.

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

Originally an acid was perceived to be a substance which had a sour taste, and which reacted with certain metals (Zn and Fe).

Around 1780, Antoine Lavoisier proposed that acids were substances which contained oxygen. He justified this by citing non-metal oxides such as sulfur dioxide ( $\text{SO}_2$ ) were acidic. However, his theory was disproved by the existence of acids such as hydrochloric acid which do not contain oxygen.

In 1815, Humphry Davy suggested that acids contained replaceable hydrogen. That is, hydrogen that could easily be displaced by reaction with metals. For example, HCl reacts with zinc to produce zinc chloride and hydrogen gas. This theory was accepted for most of the 19<sup>th</sup> century,

In 1884, Svante Arrhenius proposed that an acid was a substance that produced hydrogen ion in water. He also defined bases as substances which produced hydroxide ions in water. However there are some flaws in Arrhenius' definition. The ionisation of an acid is something that happens in a solution, not in isolation. He does appropriately address the role of the solvent in ionisation.

In 1923, two scientists Bronsted and Lowry independently devised a new definition. They defined acids as proton donors and bases as proton acceptors. This is the definition currently used.

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

See Prac Book (2.4)

12) 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

See Prac Book (2.5)

13) Perform a first-hand investigation to determine the concentration of a domestic acidic substance

See Prac Book (2.6)

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

Neutralisation reactions are used for safety purposes in laboratories and industry since many acids and alkalis are very corrosive. It is important to neutralise spills of these substances quickly. In addition sewage authorities put strict limits on the pH of factory and laboratory effluent discharge.

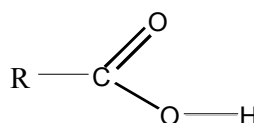
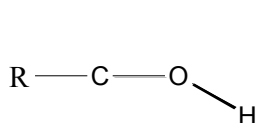
Sodium carbonate ( $\text{Na}_2\text{CO}_3$ ) is widely used to neutralise acid spills since it is a stable solid, which is easily stored. It is also cheaply available, and if too much is used, it is less danger than other bases.

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

- 1) Describe the differences between the alkanol and alkanolic acid functional groups in carbon compounds

Alkanols contain the functional group OH

Alkanolic acids contain the functional group COOH

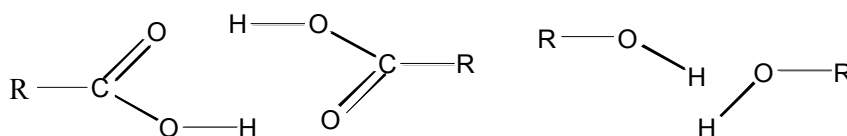


- 2) Identify the IUPAC nomenclature for describing the esters produced by reactions of straight-chained alkanolic acids from C1 to C8 and straight-chained primary alkanols from C1 to C8

The alkanol loses its ending '-anol', and 'thyl' is added. The alkanolic acid becomes an '-oate'  
e.g. methanol + ethanoic acid becomes methyl ethanoate

	Alkanol	Alkanolic acid
C <sub>1</sub>	Methyl	Methanoate
C <sub>2</sub>	Ethyl	Ethanoate
C <sub>3</sub>	Propyl	Propanoate
C <sub>4</sub>	Butyl	Butanoate
C <sub>5</sub>	Pentyl	Pentanoate
C <sub>6</sub>	Hexyl	Hexanoate
C <sub>7</sub>	Heptyl	Heptanoate
C <sub>8</sub>	Octyl	Octanoate

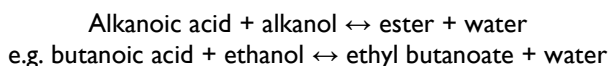
- 3) Explain the difference in melting point and boiling point caused by straight-chained alkanolic acid and straight-chained primary alkanol structures



Whilst the boiling point of an alkanol is higher than its complimentary alkane, due to hydrogen bonding between the alkanol molecules, there exist two hydrogen bonds between a pair of alkanolic acid molecules as demonstrated in the diagram above. Since hydrogen bonds are a relatively strong intermolecular force, more energy is required to break one two hydrogen bonds per pair of molecules than is required to break one bond per pair of molecules in an alkanol. Therefore an alkanolic acid has a higher boiling point than its complementary alkanol.

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

Esterification is the process where an ester is formed by an alkanol and an alkanolic acid. They are named as alkyl alkanoates, where the alkyl is the abbreviated form of the alkanol used and alkanoate the form of the alkanolic acid.



This acid catalysed, reversible reaction is called esterification.

5) Describe the purpose of using acid in esterification for catalysis

Only a few drops of concentrated acid needs to be added to the esterification mixture of alkanol and alkanolic acid to catalyse the reaction.

If concentrated sulfuric acid is added in large amounts, it can have a significant effect on the position of equilibrium. Concentrated sulfuric acid is a dehydrating agent, that is, it has a strong affinity for water. If a large amount of sulfuric acid is present, it will shift equilibrium position to the right, by absorbing water, and thus produce more ester.

6) Explain the need for refluxing during esterification

The esterification process is endothermic, and thus heating the reaction would push equilibrium to the right producing more esterification. However the products and reactants of esterification are highly volatile, meaning they easily escape as vapour. Refluxing is the process of heating a mixture in a vessel with a cooling condenser attached in order to prevent loss of any volatile reactants or products.

Thus, it allows the reaction to proceed at a higher temperature than would otherwise be possible. Performing the experiment in a closed system would lead to a dangerous build-up of gases.

7) Outline some examples of the occurrence, production and uses of esters

Esters generally have fruity and/or pleasant odours, and are therefore used extensively in artificial perfumes and food flavourings e.g. artificial flavours such as apple, banana ore grape can be produced with esters.

Flavour is essentially a crude combination of taste and odour, and it is the odour of esters that contributes to their flavours. There is now an extensive industry for developing and manufacturing synthetic flavours and perfumes. The first step is to identify the chemical constituents of the naturally occurring flavour, and then to synthesise it with esters. The artificial flavours are usually cheaper to produce than their natural extracts, and do not pose any known health risks.

8) Identify data, plan, select equipment and perform a firsthand investigation to prepare an ester using reflux

See Prac Book (2.7)

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

See above