

The Acidic Environment

- 1.1.1**
- (a) An indicator is a substance that changes colour if placed in an acid or a base. It is an organic molecule whose colour depends on the acidity of the solution in which it is dissolved.
 - (b) Various, e.g. litmus, methyl orange, universal indicator, bromothymol blue.
 - (c) Red cabbage leaves, many flower petals, e.g. geraniums and roses, lichens.
 - (d) Chop them finely (or grind with a mortar and pestle). Pour boiling water on them. Allow to sit for a few minutes. Drain off the liquid – this acts as an indicator.

- 1.1.2**
- (a) Various, e.g. red cabbage leaves.
 - (b) Various, e.g.
 Pour boiling water over chopped red cabbage leaves.
 Soak, then decant the coloured solution (the indicator).
 Record the colour of the solution.

Add 5 mL of this indicator to 3 identical test tubes containing:

- water
- dilute hydrochloric acid
- dilute sodium hydroxide solution.

Record the colour in each liquid.

- (c) Various, e.g.

Chemical used	Colour with red cabbage leaf water
Water	Purple
Hydrochloric acid	Pink
Sodium hydroxide	Green

- (d) Various, e.g. Red cabbage water would be very suitable for use as an indicator. It is a distinctly different colour in acidic, basic and neutral substances, so would be very useful in determining whether a substance is acidic, basic or neutral.
 - (e) If you (or other students) repeated the testing process, and obtained identical results, then you could say that your results were reliable. Without repetition you do not know if your results are reliable.
 - (f) pH probe with data logger and computer.
- 1.2.1** Various, e.g. you may have used a range of plant materials to make natural indicators, as well as some commercial indicators such as methyl orange, litmus and universal indicator. You may have added small quantities of each of these potential indicators to a range of substances including acids, bases and salts. You should be able to name the substances you used.
- 1.2.2** Various, e.g. Different indicators change colour at different pH values.
- 1.2.3** Various, e.g. Acids and bases are corrosive so avoid contact with the skin and eyes.
- 1.3.1**
- (a) No change.
 - (b) No change.
 - (c) No change.
 - (d) Red would stay red and blue would turn red.
 - (e) No. Both must be used and their colour changes compared.

- 1.3.2 (a) 3.1 to 4.4
- (b) No. Bromothymol blue changes colour closest to pH 7. However, if it is yellow, then the substance could be 6.0 or below. If it turns blue, then the substance is 7.6 or higher. There is no colour change at pH = 7.
- (c) Place a small sample of A, B and C into three separate test tubes.

Add a few drops of bromothymol blue to each test tube.

Repeat using phenolphthalein.

Identify each solution from the pattern of results as shown in the table.

pH of solution	Colour with bromothymol blue	Colour with phenolphthalein
5.5	yellow	colourless
8.0	blue	colourless
10.3	blue	red

- 1.3.3 (a) red
- (b) green
- (c) violet

1.4.1 Use indicators, a pH meter or pH probe and data logger.

1.4.2

Chemical	Formula	Acidity	Home use
Acetic (ethanoic) acid	CH_3COOH	Acid	In vinegar, to flavour and preserve food.
Ethanol	$\text{C}_2\text{H}_5\text{OH}$	Neutral	Alcoholic drinks and in methylated spirits.
Sodium hydrogen carbonate	NaHCO_3	Base	Bicarb soda used to make cakes rise and to safely neutralise acids.
Magnesium sulfate	MgSO_4	Slightly basic	Epsom salts — for constipation.
Hydrochloric acid	HCl	Acid	Cleaning cement off bricks. Maintain pool pH.
Sodium hydroxide	NaOH	Basic	In drain cleaners.





- 1.5.1 (a) The colour of the detergent would mask any colour change of the indicator.
- (b) Dilute the detergent until there is no noticeable colour and test the diluted solution; or use a pH meter or pH probe and data logger.

- 1.5.2 (a) Red rose petals, red geranium petals, red cabbage leaf.
- These are useful as they change colour in acid and base.

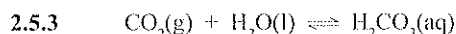
- (b) (i) pink
- (ii) green
- (iii) pink
- (iv) green

- 1.5.3 (a) 7.5 to 8.5
- (b) 4.5 to 6
- (c) red with methyl orange, yellow with bromothymol blue, colourless with phenolphthalein.

- 1.6.1 When testing the pH of substances such as soil, the colour of the soil can hide the indicator colour change. To prevent this, an insoluble, neutral white powder, such as barium sulfate can be added to the top layer of damp soil before adding the indicator. This acts as a white background so any colour change is readily visible.

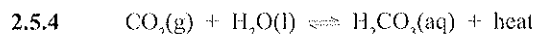
- 1.6.2 Various, e.g.:
- Checking the water in swimming pools – The pH of swimming pool water needs to be kept close to 7.4 so as not to cause skin and eye discomfort.
 - Testing aquarium water – Fish are sensitive to the pH of water so it must be maintained at a suitable level. Saltwater fish may need a pH of about 8.5.
 - Finding the end point of an acid-base reaction – Indicators are used to do this in a quantitative technique called a titration.
- 2.1.1 Oxides of non-metals, e.g. sulfur dioxide and carbon dioxide act as acids when dissolved in water.
- 2.1.2 sulfur dioxide + water \rightarrow sulfurous acid
 $\text{SO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{SO}_3(\text{aq})$
 carbon dioxide + water \rightarrow carbonic acid
 $\text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{CO}_3(\text{aq})$
- 2.1.3 Carbon monoxide (CO) and dinitrogen monoxide (N₂O)
- 2.2.1 Key
- | | | | |
|---|--------|---|------------|
|  | Basic |  | Amphoteric |
|  | Acidic |  | No oxide |
- 2.2.2 Acidic, basic.
- 2.2.3 (a) Add water and indicator.
 See if it reacts with a base to form water and a salt.
- (b) Sulfur dioxide + sodium hydroxide \rightarrow water + sodium sulfate
 $\text{SO}_2(\text{g}) + 2\text{NaOH}(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{Na}_2\text{SO}_3(\text{aq})$
- 2.2.4 (a) Magnesium oxide + water \rightarrow magnesium hydroxide
 $\text{MgO}(\text{s}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{Mg}(\text{OH})_2(\text{aq})$
- (b) Magnesium oxide + hydrochloric acid \rightarrow water + magnesium chloride
 $\text{MgO}(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{MgCl}_2(\text{aq})$
- 2.3.1 (a) closed
 (b) constant
 (c) constant
 (d) microscopic
 (e) backward
- 2.3.2 If a system is in equilibrium and it is disturbed or changed in some way, then the system adjusts itself to minimise the amount of change.
- 2.4.1 Changes in concentration, temperature. Pressure and volume changes can also affect equilibrium if one or more reactants or products is a gas.
- 2.4.2 A catalyst does not have any effect on an equilibrium position. It speeds up the reaction in both directions and thus makes the system reach equilibrium sooner.
- 2.5.1 Carbon dioxide is soluble in water; its solubility increases as temperature decreases. When carbon dioxide dissolves in water in a closed system, the following equilibrium can occur:
 $\text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_2\text{CO}_3(\text{aq})$

- 2.5.2 (a) $\text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_2\text{CO}_3(\text{aq})$
- If the **concentration of carbon dioxide** is increased, the equilibrium shifts to the right to use up the added carbon dioxide – more dissolves.
 - If **more carbonic acid is added** to the system, the equilibrium shifts to the left to use up the carbonic acid added – less carbon dioxide dissolves.
 - If **carbonic acid is removed**, decreasing its concentration, the equilibrium shifts to the right to make more carbonic acid and replace what was removed – more carbon dioxide dissolves.
- (b) In each case the equilibrium shifts to minimise any change to the system as in Le Châtelier's principle, e.g. when more of a chemical is added the equilibrium shifts to whichever direction will use up that added substance. This fits with Le Châtelier's principle which states that if a system is in equilibrium and it is disturbed or changed in some way, then the system adjusts itself to minimise the amount of change.



Because a gas is present (carbon dioxide), changing the pressure will also alter the equilibrium position. If we increase the pressure, the equilibrium will move to try and reduce the pressure again. It can only do this by using up the carbon dioxide so there are fewer particles. Thus the equilibrium moves to the right and more carbonic acid is formed.

If pressure is decreased, equilibrium moves left to increase the gas particles and thus minimise the change as per Le Châtelier's principle.



Carbon dioxide becomes more soluble as temperature decreases.

This reaction is exothermic.

If the solution is heated, the equilibrium position will change so as to use up the added heat as per Le Châtelier's principle. Equilibrium will move to the left, so less carbon dioxide will dissolve as the solution becomes warmer.

Conversely, in cold water, the equilibrium will lie to the right, so more carbon dioxide will dissolve.

- 2.6.1 (a) Quantity of a substance that contains 6.022×10^{23} particles (Avogadro's number).
 (b) The formula mass of a substance in grams.
 (c) The volume, in litres, occupied by 1 mole of any gas at a particular temperature and pressure.

2.6.2 22.71 L, 24.79 L

- 2.6.3 number of moles = $\frac{\text{mass}}{\text{molar mass}}$
 number of moles = $\frac{\text{volume}}{\text{molar volume}}$
 number of moles = concentration (mol/L) \times volume (L)
 number of moles = $\frac{\text{number of particles}}{6.022 \times 10^{23}}$

2.6.4

Mass of gas	Number of moles of gas	Volume (L) of gas at 0°C	Volume (L) of gas at 25°C
44.01 g of carbon dioxide	1.00 mole	22.71	24.79
44.01 g of sulfur dioxide	0.69 mole	15.60	17.08
220.05 g of carbon dioxide	5.00 moles of carbon dioxide	113.55	123.95
320.35 g of sulfur dioxide	5.00 moles of sulfur dioxide	113.55	123.95
450.73 g of CO ₂	10.24 moles of CO ₂	232.55	253.85

- 2.6.5 (a) $\text{CuS}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{SO}_2(\text{g}) + \text{Cu}(\text{s})$
 (b) 10.2×10^3 mol
 (c) 10.2×10^3 mol
 (d) 2.54×10^5 L

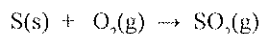
- 2.7.1**
- (a) Removing dissolved carbon dioxide from solution in the soft drink.
 - (b) Various, e.g.
Weigh a bottle of soft drink.
Open the cap and allow the gas to leave the bottle (you may have increased this effect by warming the soft drink so that the gas becomes less soluble).
Reweigh the bottle and lid.
Calculate the mass of the gas that has escaped (change in weight of the bottle).
Calculate the moles of carbon dioxide that have escaped and the volume this gas would occupy.
Compare the smaller change in mass of the soft drink with the large calculated volume of the same amount of gas released.
At the same time set up a control – same type and size of container containing the same volume of water as the volume of the soft drink.
 - (c) Control is needed to find out how much weight loss is due to evaporation of water rather than loss of carbon dioxide gas. Thus the loss of mass from the control bottle can be subtracted from the mass lost by the experimental bottle, providing more accurate results. A control improves the validity of the experimental design.
 - (d) Various, e.g. the soft drink and control would need be identical and to contain the same volume of liquid and be placed in the same area so they were at the same temperature and other physical factors such as any breeze would be identical for both containers. They would both be set up at the same time and left standing for the same period of time before reweighing. Controlling variables improves the validity of the experimental method.
 - (e) Various. You may not personally have repeated your experiment because of time constraints. However, there was probably more than one group in the class, so that should have allowed a comparison of results. Repeating experiments is important as it checks the reliability of experiment results. Discuss how closely your results agreed and make an assessment as to their reliability.
- 2.7.2**
- (a) 0.06 mol
 - (b) 1.41 L
- 2.7.3** Pure distilled water may have a pH of 7 when just distilled. With time the pH drops below 7 due to the formation of carbonic acid as carbon dioxide from the air dissolves in the distilled water.
- $$\text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_2\text{CO}_3(\text{aq}) + \text{heat}$$
- pH will be lower on a colder day as this is an exothermic reaction, so cooling pushes the equilibrium to the right and more acid is produced, lowering the pH (Le Châtelier's principle).
- 2.8.1**
- (a) Colourless, toxic gas with a pungent odour.
 - (b) Extremely irritating to the eyes and respiratory tract, causing lung damage and asthma.
 - (c) Sulfur dioxide is used as a food additive where it acts as a preservative, for example in wine, dried fruits and manufactured meat such as salami. It is also used in the manufacture of sulfuric acid.
 - (d) Release of sulfur dioxide into the atmosphere is one of the main causes of acid rain. Most of the sulfur dioxide released into the atmosphere returns to earth as sulfuric acid in acid rain or as sulfate particles.
- 2.8.2** Natural sources – Combustion of organic matter, e.g. bushfires, decomposition of organic matter and volcanoes.
- Industrial sources – Combustion of fossil fuels, especially in power plants and motor vehicles, metal smelters extracting lead, copper and zinc from sulfides, incineration of garbage, petroleum refineries and industries using sulfur dioxide for production of sulfuric acid, paper, food processing and sewage treatment.

2.8.3

Names and formula	Appearance and acidity	Natural sources	Industrial sources
Nitrogen dioxide, nitrogen(IV) oxide (NO_2)	Red-brown acidic gas Poisonous	Action of sunlight on NO and O_2	Combustion of fossil fuel in motor vehicles and power stations.
Dinitrogen monoxide, nitrogen(I) oxide, nitrous oxide (N_2O)	Colourless, neutral gas	Produced by soil bacteria	Manufactured as a fuel for combustion in racing cars and for use as a sedative/analgesic (called laughing gas).
Nitrogen monoxide, nitrogen (II) oxide, nitric oxide (NO)	Colourless, acidic gas	Produced by soil bacteria and lightning	Burning of biomass. Combustion of fuel in motor vehicles and power plants.

2.9.1 The main contributor to high sulfur dioxide levels is the combustion of fossil fuels in power plants and vehicles. Coal and petroleum products contain sulfur, which combines with oxygen in air to form sulfur dioxide.

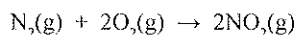
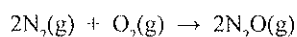
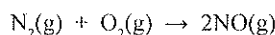
sulfur + oxygen \rightarrow sulfur dioxide



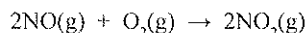
The smelting of metal sulfides such as copper sulfide also produces sulfur dioxide.



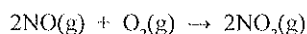
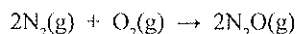
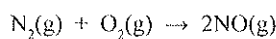
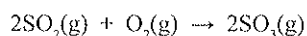
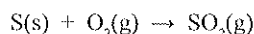
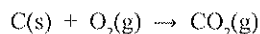
2.9.2 Nitrogen undergoes combustion to form nitrogen monoxide and dinitrogen monoxide. Fossil fuels and biomass contain nitrogen (in proteins), so they also burn to produce these oxides.



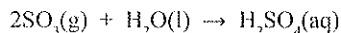
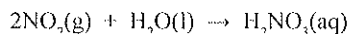
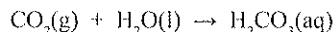
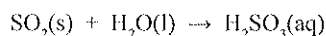
Nitrogen monoxide burns to produce nitrogen dioxide.



2.10.1 Acid rain forms when acidic oxides of non-metals dissolve in rain to form acids. These oxides are formed mainly by the combustion of fossil fuels, especially in motor vehicles and power stations. Combustion of garbage and biomass also contribute to the formation of these acidic oxides and industries such as metal sulfide smelters also produce them. The main acidic oxides involved in the formation of acid rain are carbon dioxide, sulfur dioxide and oxides of nitrogen. These form by the combustion of carbon, sulfur and nitrogen present in fossil fuels. Include equations, e.g.



These oxides form acids when they dissolve in rainwater, making the rain acidic, e.g.



2.10.2 Various.

Effects of acid rain	Explanation
Causes defoliation and stunts growth of plants and some sensitive ones, e.g. spruce trees, are destroyed.	Causes chemical reactions in soil, e.g. Aluminium salts, which are usually insoluble, become soluble in acidic water, releasing toxic aluminium ions into the soil where they are absorbed by plants. Sulfuric acid ionises, forming hydrogen ions, hydrogen sulfate ions and sulfate ions in water. The sulfate ions attract calcium and magnesium ions that occur naturally in soil and are essential for plant growth. The sulfate ions leach (draw out) the calcium and magnesium ions from the soil, leaving it low in these essential minerals. This affects growth.
Animal life is affected.	Acid rain lowers the pH of lakes and streams, killing fish eggs and other aquatic life such as crayfish which cannot tolerate a low pH. In many cases the death of fish, or their inability to reproduce, has been caused by the presence of aluminium released from the surrounding soil. Aluminium also causes the gills of fish to clog with mucus.
Corrosion and decay of metals, buildings, stone structures, paint and statues.	Acid rain dissolves the calcium carbonate in concrete, limestone and marble. $\text{CaCO}_3(\text{s}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) + \text{CaSO}_4(\text{s})$ It also corrodes metals and other materials. $\text{Fe}(\text{s}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{H}_2(\text{g}) + \text{FeSO}_4(\text{s})$
Sulfate particles have affected peoples' health.	Sulfate particles suspended in air are very fine and can be inhaled. They have contributed to the rising incidence of chronic respiratory diseases in humans, for example lung cancer, asthma, bronchitis and emphysema. In the eastern United States sulfate particles make up about 25% of the particles inhaled.

2.11.1 Various – describe and explain the evidence, then make an assessment.

In your answer you should discuss evidence from a number of areas such as:

- The analysis of Antarctic ice core samples by the CSIRO and the Australian Antarctic Division.
- Increasing damage to buildings, forests, and aquatic organisms.
- Increase in acid rain, which is mainly caused by acidic oxides of nitrogen and sulfur dissolved in rain water, e.g.
$$\text{SO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{SO}_3(\text{aq})$$
- Increase in acidity of rivers and lakes.
- Higher atmospheric concentrations of sulfur dioxide and nitrogen oxides in industrial areas than in non-industrial areas.

You should also discuss difficulties with obtaining evidence, e.g.

- It is difficult to obtain accurate evidence about changes in the atmospheric concentration of sulfur and nitrogen oxides because these gases are only present in such small concentrations (about 0.001 ppm), and instruments capable of measuring these small concentrations have only been available since the 1970s.
- Also sulfur dioxide and nitrogen dioxide form sulfate and nitrate ions respectively and these are soluble in water, so they circulate in the biosphere and hydrosphere, but are not precipitated out in forms we can readily study. The validity of atmospheric measurements of sulfur and nitrogen oxides can be questioned when they are so readily soluble.

Then make an assessment, e.g.

Although we do not have accurate measures of atmospheric oxides of sulfur and nitrogen taken over a long period of time, there is enough indirect evidence to conclude that significant increases in atmospheric concentrations of oxides of sulfur and nitrogen have indeed taken place, especially since industrialisation.

- 2.12.1**
- Combustion of fossil fuels, especially in power plants and motor vehicles, metal smelters extracting lead, copper and zinc from sulfides, incineration of garbage, petroleum refineries and industries using sulfur dioxide for production of sulfuric acid, paper, food processing and sewage treatment.
 - Nitrogen dioxide – Combustion of fossil fuel in motor vehicles and power stations.
Dinitrogen monoxide – Manufactured as a fuel for racing cars.
Nitrogen monoxide – Burning of biomass, combustion of fuel in motor vehicles and power stations.

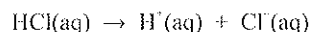
2.12.2 Various. In your answer you should discuss (describe, explain and include equations where possible):

- Effects of sulfur dioxide and nitrogen oxides on human health.
- Effects of sulfur dioxide and nitrogen oxides on the environment.
- The formation of acid rain.
- Effects of acid rain on human health and on the environment.
- An evaluation – is there reason for concern or not? Should we be very concerned or only slightly concerned?

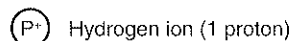
- 3.1.1 (a) Various, for example you may have set up a series of beakers, each containing solutions of laboratory chemicals (e.g. HCl, H₂SO₄, NaCl, Na₂CO₃, NaHCO₃, NH₄OH) or household substances (e.g. orange juice, detergent, drain cleaner, tap water, rainwater, liquid soap, alcohol). The probe would be placed into each solution in turn to measure the pH, washing it after testing each solution to avoid contamination.
- (b) pH meter or probe is more accurate, it provides an actual pH reading. Indicators only give a range within which the pH falls.

3.2.1 Acids ionise in water to release hydrogen ions. Each hydrogen ion consists of one proton.

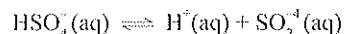
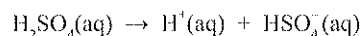
For example, hydrochloric acid ionises in water to form hydrogen ions and chloride ions.



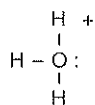
3.2.2



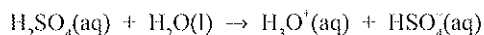
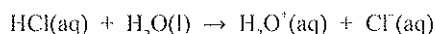
3.2.3 Sulfuric acid can release one or two hydrogen ions (protons). It is described as diprotic (two hydrogen ions).



3.2.4 (a) Diagram of hydronium ion:



- (b) The hydrogen ions (protons), released by an acid ionising, do not usually exist alone. They attach themselves to a water molecule, forming an hydronium ion (H₃O⁺). This is sometimes called a hydrated hydrogen ion, e.g.



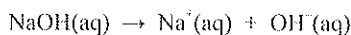
3.3.1

Name	Formula	Found in/used for
Acetic (ethanoic) acid	CH ₃ COOH	Vinegar.
Citric acid (2-hydroxypropane-1,2,3-tricarboxylic acid)	$\begin{array}{c} \text{CH}_2\text{COOH} \\ \\ \text{COHCOOH} \\ \\ \text{CH}_2\text{COOH} \end{array}$	Citric fruit, e.g. oranges, lemons.
Hydrochloric acid	HCl	Stomach acid. Cleaning bricks.
Sulfuric acid	H ₂ SO ₄	Acid rain, batteries.

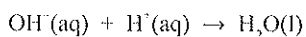
3.4.1

Common substance	Acid present	Chemical composition
Citric juice, e.g. oranges, and lemons	Citric acid, ascorbic acid	Organic compounds
Vinegar	Acetic acid	Organic compound (CH ₃ COOH) and water
Stomach acid	Hydrochloric acid	HCl
Sour milk, yoghurt	Lactic acid	CH ₃ CHOHCOOH
Ant stings	Methanoic acid	HCOOH

- 3.4.2** A base will ionise in water to form hydroxide ions. The hydroxide ions will accept a proton to form water, e.g. sodium hydroxide ionises to form hydroxide ions.



These hydroxide ions accept a proton(hydrogen ion) to form water.



- 3.4.3** A base is a substance that will accept a proton (hydrogen ion). An alkali is a base that is soluble in water.

3.4.4

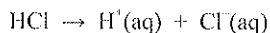
Common substance	Base present	Natural or synthetic	Chemical composition
Limewater	Calcium hydroxide	Synthetic	Ca(OH)_2
Bicarbonate of soda	Sodium hydrogen carbonate	Natural	NaHCO_3
Lime	Calcium oxide	Synthetic	CaO
Drain and oven cleaner	Sodium hydroxide	Synthetic	NaOH
Household cleaners (with ammonia)	Ammonium hydroxide	Synthetic	NH_4OH

- 3.5.1** The pH (potential hydrogen) scale is a logarithmic scale that indicates the concentration of hydrogen ions in a solution. A low pH indicates a high concentration of hydrogen ions and thus an acidic substance. A high pH indicates a low concentration of hydrogen ions and thus a basic substance.

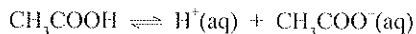
3.5.2

Substance	Approximate pH
0.1 mol L ⁻¹ HCl	1
Stomach acid (HCl)	2
Lemon juice	2
Vinegar	3
Soft drink (contains carbonic acid)	3
Wine	5
Urine	6
Pure water	7
Human blood	7.4
Sea water	8
Toothpaste	9-10
Detergent	10-11
Oven and drain cleaner	11-14

- 3.6.1** A strong acid is one that ionises completely, donating protons freely. Examples are HCl, H₂SO₄, HNO₃, HBr, HI.

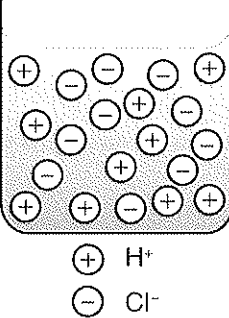
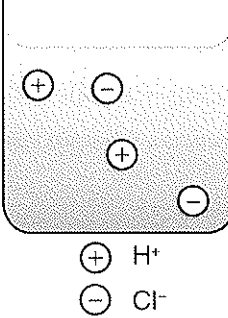
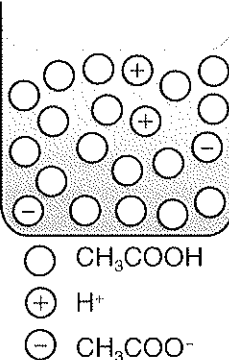
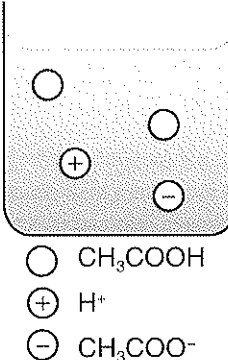


A weak acid is one that does not ionise completely; it is not a good proton donor. Examples are H₂CO₃, CH₃COOH, H₂SO₃, H₂S.



- 3.6.2** A concentrated acid contains a large number of moles of the substance per litre.

A dilute acid contains a small number of moles of the substance per litre.

	Concentrated solution	Dilute solution
Strong acid, e.g. HCl		
Weak acid, e.g. CH ₃ COOH		

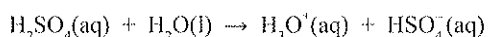
3.7.1 (a) Various, e.g. you may have used 0.1 mol L⁻¹ solutions of ethanoic (acetic), carbonic acid, hydrochloric and sulfuric acids.

(b) Various, e.g. you may have used indicators such as methyl orange or a pH meter or probe and data logger.

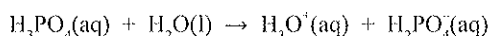
(c) Various, e.g. you could rank the acids in order of acidity according to your measurements – the lower the pH, the greater the [H⁺] concentration – and based on these results you should be able to draw conclusions as to the degree of ionisation, and thus the strength, of the acids you used.

3.8.1

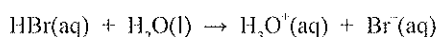
(a) $\text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{H}^+(\text{aq}) + \text{HSO}_4^-(\text{aq})$ or



(b) $\text{H}_3\text{PO}_4(\text{aq}) \rightarrow \text{H}^+(\text{aq}) + \text{H}_2\text{PO}_4^-(\text{aq})$ or



(c) $\text{HBr}(\text{aq}) \rightarrow \text{H}^+(\text{aq}) + \text{Br}^-(\text{aq})$ or



(d) $\text{CH}_3\text{COOH}(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{CH}_3\text{COO}^-(\text{aq})$ or



3.9.1

Pure acids such as HCl occur as molecules; there are no ions present. When the acid reacts with water it forms ions and is then classified as an ionic substance.

3.9.2

(a) Various – perhaps you used diagrams or ball-and-stick models or molecular model kits. You could use plasticine, smarties, marbles or any other similar objects with objects such as toothpicks to show the bonds.

(b) Various. You may have made models of molecules and removed ‘ions’, or you may have used diagrams or a computer simulation.

3.10.1

(a) [H⁺] indicates the concentration of hydrogen ions, measured in moles per litre (mol L⁻¹).

(b) (i) 1

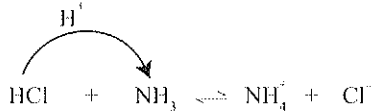
(ii) 2

(iii) 3

(c) (i) 10

(ii) 100

- 3.10.2 (a) $\text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{OH}^-(\text{aq})$
 (b) 10^{-7} moles per litre
 (c) In pure water or a solution of an acid, a base or a salt, water ionises, producing small quantities of hydrogen ions and hydroxide ions.
 $\text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{OH}^-(\text{aq})$
 At any given temperature, the product of the $[\text{H}^+]$ and the $[\text{OH}^-]$ is constant. At 25°C its value is always 10^{-14} .
 $[\text{H}^+] \times [\text{OH}^-] = 10^{-14}$
 The ionic product (or ionisation constant) of water is 10^{-14} , and the symbol used is K_w .
 (d) In any aqueous solution, the concentration of hydrogen ions or hydroxide ions can be calculated if you know the concentration of one of them.
- 3.10.3 (a) 10^{-8}
 (b) 10^{-10}
 (c) (i) 10^{-7}
 (ii) 7
 (iii) 7
- 3.10.4 (a) 7
 (b) 11
 (c) 11
- 3.11.1 Measure pH with a pH meter or a probe.
 Compare their ability to conduct an electric current.
- 3.11.2 Strong acids ionise completely, weak acids are only partially ionised. Thus there are more ions present in strong acids than weak acids of the same concentration. Ions are needed for a charge to travel through a solution. The more ions, the better the solution can conduct the electric current.
- 3.11.3 (a) Hydrochloric acid – it is a strong acid, it completely ionises forming the maximum number of ions. The other two are weak acids, they do not completely ionise. As they contain fewer ions, they are less able to conduct electricity.
 (b) Hydrochloric acid – strongest acid, ionises most, most hydrogen ions present, thus the most acidic and lowest pH.
- 3.12.1 (a) A **strong acid** is one that ionises completely, donating protons freely, e.g. Hydrochloric acid is a strong acid so the equilibrium lies completely to the right; all molecules ionise.
 $\text{HCl} \rightarrow \text{H}^+(\text{aq}) + \text{Cl}^-(\text{aq})$
 Examples of strong acids are HCl , H_2SO_4 , HNO_3 , HBr , HI .
 A **weak acid** is one that does not ionise completely; it is not a good proton donor, e.g.
 $\text{CH}_3\text{COOH} \rightleftharpoons \text{H}^+(\text{aq}) + \text{CH}_3\text{COO}^-(\text{aq})$
 Acetic acid is a weak acid; the equilibrium lies to the left. Only a small percentage (about 1%) of the molecules ionise; the rest remain as molecules of acetic acid.
 Examples of weak acids are H_2CO_3 , $\text{C}_2\text{H}_5\text{COOH}$, H_2SO_3 , H_2S .
 (b) $\text{pH} = -\log_{10}[\text{H}^+]$ and as strong acids ionise completely and weak acids only ionise partially, the hydrogen ion concentration will be greater in a strong acid than in a weak acid of the same concentration.
- 3.13.1 (a) 1
 (b) 0.7
 (c) 2.1
 (d) 13

- 3.13.2 (a) 1.3
(b) 4.1
(c) 2.7
- 3.14.1 (a) They are all acids.
(b) Acids prevent the growth of micro-organisms, such as *Clostridium botulinum*, which are sensitive to acids and cannot grow in an acidic environment.
(c) People may be allergic to the additive and also some additives, e.g. sulfur dioxide, cause asthma in susceptible people.
- 4.1.1 (a) Various, e.g.
www.dbhs.wvusd.k12.ca.us/AcidBase/Acidbase.html
(b) Information is organised under headings and this helps you focus on the relevant aspects of the information and assists in the future analysis of the material.
- 4.2.1 Oxygen, oxygen, oxygen, oxygen
- 4.2.2 Hydrochloric, oxygen, hydrogen, hydrogen, oxygen, hydrogen
- 4.2.3 Hydrogen ions, negative, hydrogen, positive, hydroxide, negative
- 4.2.4 Lavoisier based his theory on observations of experiments. He noticed that non-metals underwent combustion to produce oxides and these oxides dissolved in water to form acid solutions. This led him to think that acids contain oxygen.
Davy based his theory on observations of the products obtained from electrolysis and the displacement of hydrogen from acids by metals.
Arrhenius' theory is based on ideas (concepts) about the types of particles in acids and bases.
- 4.2.5 Arrhenius' definition of acids and bases is useful, but it is limited because:
- it applies only to aqueous solutions
 - it only accounts for substances which already have hydrogen or hydroxide ions in their structure, and does not explain why some salts act as acids and some as bases
 - it cannot explain how some substances (described as amphoteric), e.g. zinc oxide and aluminium oxide and the hydrogen carbonate ion, can act as both an acid and a base.
- 4.3.1 Johannes Brønsted – a Danish chemist. Thomas Lowry – an English chemist.
In 1923, Brønsted and Lowry both independently proposed the same new theory about acids and bases. This came to be called the Brønsted-Lowry theory.
- 4.3.2 (a) A substance that donates a proton (hydrogen ion) to a base.
(b) A substance that accepts a proton from an acid.
(c) A reaction in which a proton is transferred from an acid to a base.
- 4.3.3 A proton is transferred from the hydrogen chloride molecule to the ammonia molecule, so this reaction is an acid-base reaction according to the Brønsted-Lowry theory.
- 
- Arrhenius would not have classified this as an acid-base reaction as there is no water present.
- 4.3.4 The Brønsted-Lowry theory applies to non-aqueous as well as aqueous solutions; it explains why some salts act as acids and some as bases even when they do not have hydrogen or hydroxide ions in their structure; and it can explain how some substances can act as both an acid and a base.

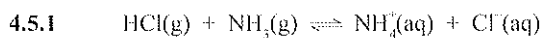
4.3.5

Model	Definition of acid	Definition of base
Arrhenius	Hydrogen ion producer in water	Hydroxide ion producer in water
Brønsted-Lowry	Hydrogen ion (proton) donor	Hydrogen ion (proton) acceptor

4.4.1 In the Brønsted-Lowry theory every acid has a conjugate base, which has one proton less than the acid. Together this acid and base are called a conjugate pair. For example, HCl is an acid, Cl^- is its conjugate base and these two make up a conjugate pair.

4.4.2

Conjugate acid	Conjugate base
H_3O^+	H_2O
HCl	Cl^-
CH_3COOH	CH_3COO^-
H_2SO_4	HSO_4^-
HSO_4^-	SO_4^{2-}
H_2O	OH^-

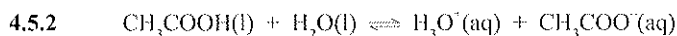


acid 1 base 2 acid 2 base 1

HCl is an acid, its conjugate base is Cl^- and these are a conjugate pair.

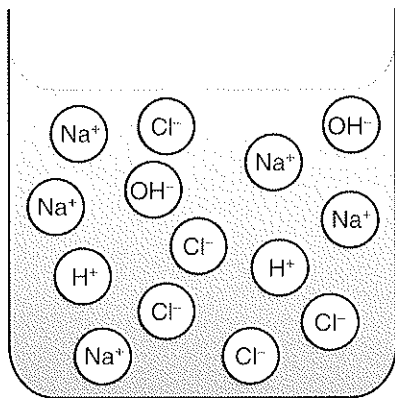
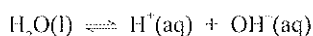
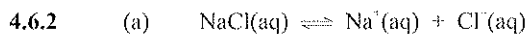
NH_4^+ is an acid, its conjugate base is NH_3 and these two are a conjugate pair.

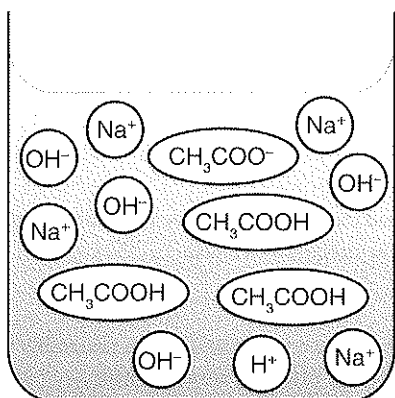
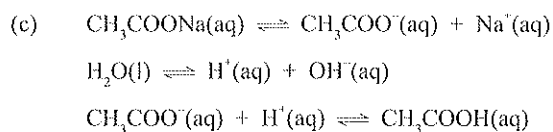
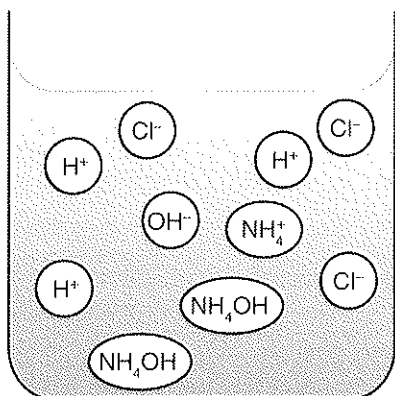
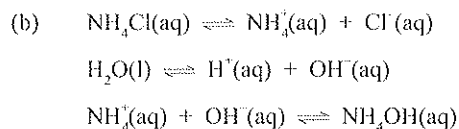
Thus there are two conjugate pairs involved in this reaction.



Conjugate pairs are (CH_3COOH and CH_3COO^-), also (H_3O^+ and H_2O).

- 4.6.1 (a) Various, e.g. you may have used indicators that detect a broad range of pH, such as universal or cabbage leaf indicators or pH meters or probes. Describe how you used the method you chose.
- (b) Various, e.g. ammonium chloride, NH_4Cl , $\text{pH} = 10$, basic.
- (c) Only if they were all the same concentration. If they were the same concentration, then remember that the solution with lowest pH, is the strongest acid and the solution with highest pH is the strongest base.





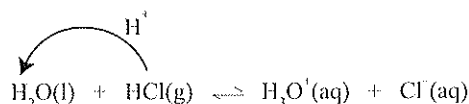
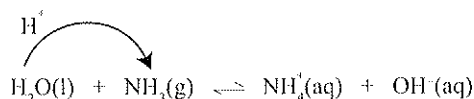
- 4.6.3 (a) Various, e.g. These models help to summarise our knowledge, to visualise and understand the concepts involved and to apply these ideas by making predictions about the acidity of other solutions.
- (b) Various, e.g. These models are simplifications, we must be careful that we realise that the chemical species involved are not drawn to scale and they do not show any information about the numbers of ions and molecules present in the solutions. They also give no idea about the movements of particles or the forces between particles.

4.7.1

Name of salt	Acidic/basic/neutral
Ammonium sulfate	Acidic
Lithium chloride	Neutral
Ammonium carbonate	Neutral
Calcium acetate	Basic
Ammonium nitrate	Acidic
Calcium carbonate	Basic
Sodium sulfate	Neutral

- 4.7.2 (a) Basic. CO_3^{2-} ion reacts with hydrogen ions in water, forming HCO_3^- or H_2CO_3 . These only slightly ionise, forming very few H^+ ions. K^+ and OH^- ions remain as ions. There are not enough hydrogen ions available to neutralise the free OH^- ions, so the salt is basic.
- (b) Neutral. Salt of a strong acid and a strong base. All exist as ions so the hydrogen and hydroxide ions neutralise each other.
- (c) Acidic. Cl^- ions are theoretically attracted to H^+ ions from the water, but HCl is a strong acid and ionises completely, forming lots of free hydrogen ions. NH_4^+ ions join with the OH^- ions from water to form NH_4OH , which is a weak base and so only slightly ionises. There are more free H^+ than OH^- ions so the salt is acidic.

- 4.8.1 (a) An amphiprotic substance is one that can act as an acid or a base, giving and receiving protons.



In the first reaction, water is acting as an acid, giving up a proton. In the second reaction, water is acting as a base, accepting a proton. Water is amphiprotic, it can act as an acid or a base by donating or accepting protons.

- (b) Amphoteric is the more general term – it refers to any substance that can act as an acid or a base, including amphiprotic substances which do so by being able to donate or receiving a proton. All amphiprotic substances are amphoteric. However, some substances are amphoteric but not amphiprotic – they can act as acid or base – but they cannot both donate and accept a proton, e.g. oxides of zinc and aluminium.

- 4.8.2 HSO_4^- ion acting as a base:

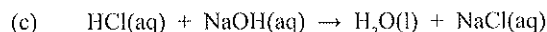


HSO_4^- acting as an acid:

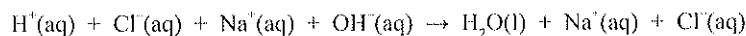


- 4.9.1 (a) A chemical reaction between an acid and a base to form water and a salt.

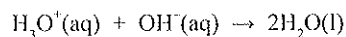
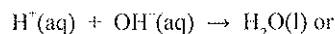
- (b) Heat is released, the container becomes warm. This indicates that neutralisation reactions are exothermic.



- (d) Full ionic equation:



- (e) Net ionic equation:



- 4.9.2 (a) $\text{H}_3\text{O}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O} + \text{H}_2\text{O}$

acid 1 base 2 acid 2 base 1

- (b) A proton is transferred from a hydronium ion to a hydroxide ion forming two molecules of water.

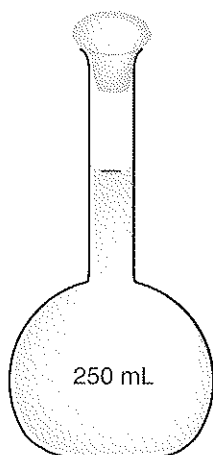
- 4.10.1**
- A procedure used to find experimentally the concentration in moles per litre (molarity) of a solution. It is an accurate **volumetric analysis** technique. In a titration, a solution of known concentration (the standard solution) is added to a solution of unknown concentration until the reaction between them is complete. The unknown concentration can then be calculated.
 - Quantitative analysis involves finding accurately the composition of a solution by measuring. Volumetric means that volumes are measured in a technique such as a titration.
 - The reactant solution, in a titration, whose concentration is accurately known.
 - The point when the reaction is complete. At this point, the reactants and products are present in the stoichiometric ratio shown by the balanced equation.
 - When the indicator changes colour. If the indicator is chosen correctly, these two points (the equivalence and end points) will coincide, the indicator changing colour when the reaction is complete.

- 4.10.2**
- Primary standard is a solution that is made by dissolving an accurately measured mass of a solute in a small amount of the solvent and making the volume up to a measured volume.

A secondary standard is a solution whose concentration is determined by titration against a primary standard.

- Can be obtained in a pure form (common acids such as HCl and H_2SO_4 would not be suitable as their concentration varies from batch to batch.)
 - Has a known chemical formula.
 - Is stable and does not change when exposed to air. (Sodium hydroxide would not be suitable as it absorbs water from air and reacts with carbon dioxide in the air).
 - Is soluble.
- The most commonly used primary standard is anhydrous sodium carbonate (Na_2CO_3). This makes a basic solution, suitable for analysing acids. A suitable chemical to use as an acidic primary standard would be hydrated oxalic acid ($\text{H}_2\text{C}_2\text{O}_4 \cdot 2\text{H}_2\text{O}$).

4.10.3



Calculate the amount of chemical needed to make the required volume of solution

Weigh out a mass of chemical exactly equal to the calculated mass

Dissolve solid in a small amount of pure water in beaker

Transfer this into a volumetric flask that holds the required volume

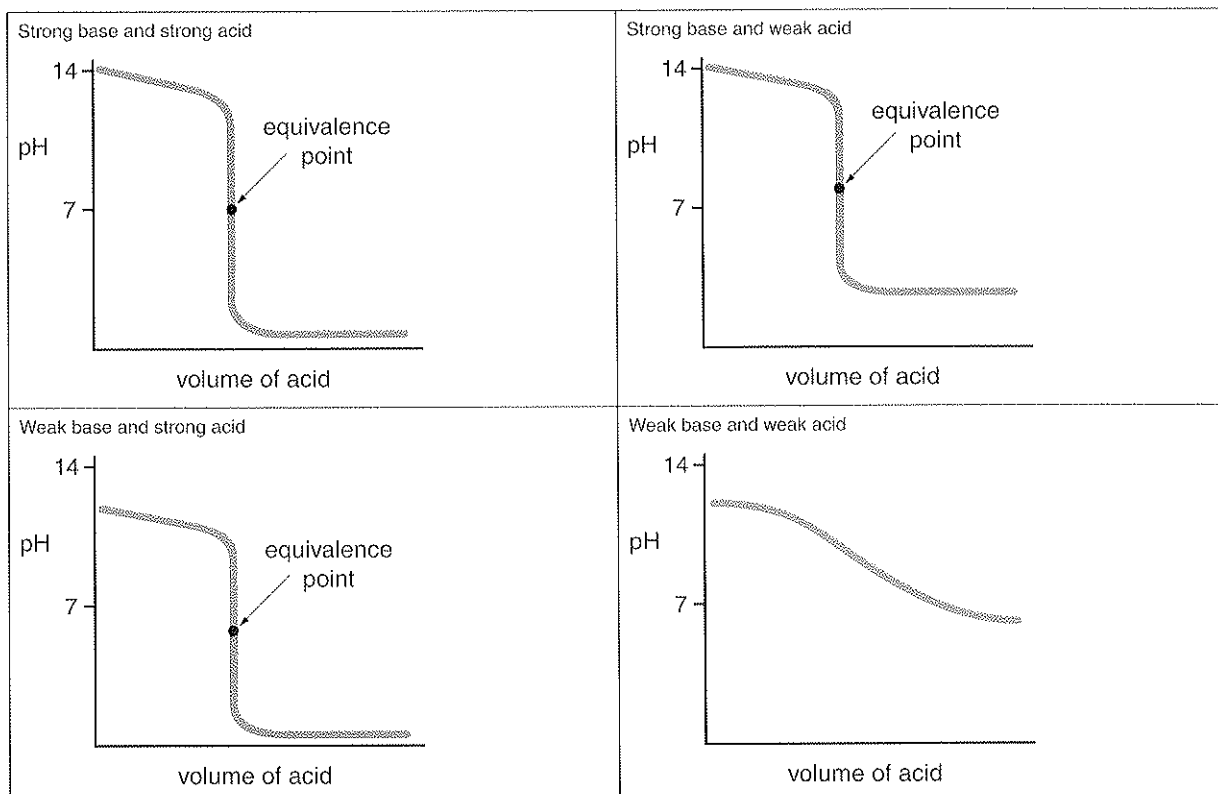
Rinse the container into the volumetric flask so as to transfer all the chemical

Add water to the graduation mark to produce the exact volume of solution

Stopper, invert and shake to mix well

Label your solution

4.10.4



4.10.5

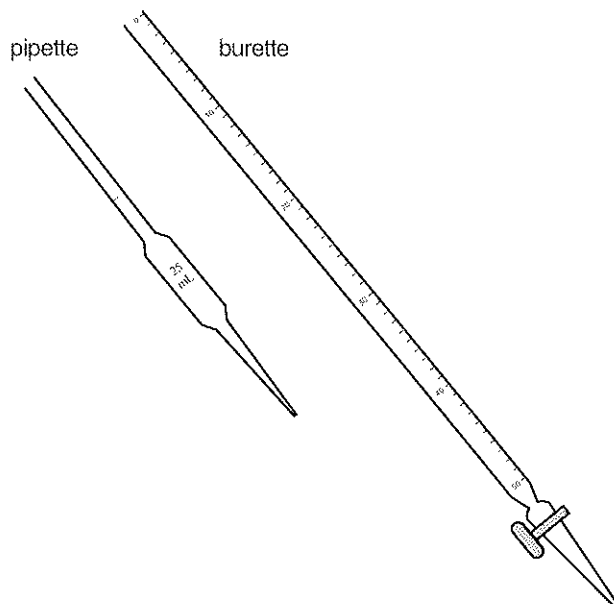
Acid	Strong or weak	Base	Strong or weak
HCl	Strong	Mg(OH) ₂	Strong
H ₂ SO ₄	Strong	NH ₄ OH	Weak
H ₂ CO ₃	Weak	NaOH	Strong
CH ₃ COOH	Weak	KOH	Strong

4.10.6

Acid	Base	Suitable indicator	Justification
HCl	KOH	Bromothymol blue	Strong acid/strong base so equivalence point approx = 7. Bromothymol blue changes colour around pH = 7.
H ₂ SO ₄	NH ₄ OH	Methyl orange	Strong acid/weak base so equivalence point < 7. Methyl orange changes colour at pH < 7.
H ₂ CO ₃	NaOH	Phenolphthalein	Weak acid/strong base so equivalence point > 7. Phenolphthalein changes colour at pH > 7.
CH ₃ COOH	Ca(OH) ₂	Phenolphthalein	Weak acid/strong base so equivalence point > 7. Phenolphthalein changes colour at pH > 7.

4.10.7 Remember that every diagram MUST be labelled.

(a)



(b)

	Pipette	Burette
Function	Accurately measures a fixed volume of solution to provide a fixed number of moles of one reactant.	Allows you to measure the exact volume of a reactant needed to reach the equivalence point.
Procedure for washing before a titration	Before use, the pipette must be washed with distilled water and then with the solution to be used in it.	Burette should be washed with distilled water and then with the solution to be used in it.
Volume measured	An aliquot.	A titre.

- 4.10.8** (a) Rinsing with the solution would change the number of moles of the chemical being added due to the film of the solution left on the glass.

A film of water left on the glass does not matter as water will not change the number of moles of the solution added to the flask.

- (b) Four times. The first titration should be a quick one to establish the approximate volume needed from the burette. You should then obtain three accurate titres, within 0.1 mL of each other, and calculate the average of these three.

- 4.10.9** (a) Ensure all equipment is clean and rinsed correctly.

Add one solution to the burette.

Use the pipette to measure a volume of the other solution.

Transfer this to a conical flask.

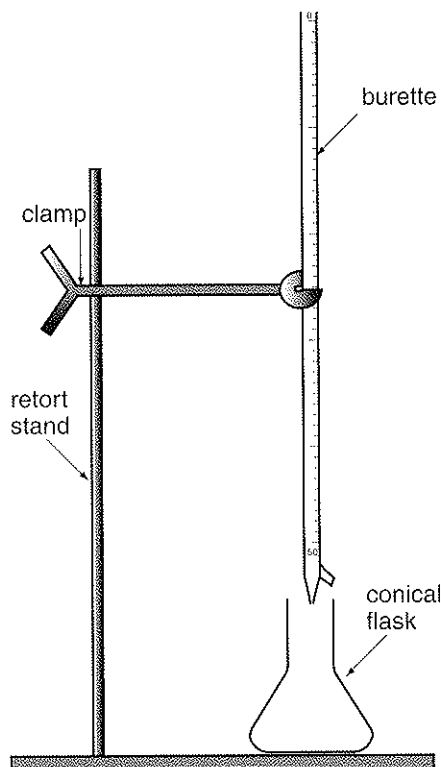
Add a few drops of a suitable indicator to the flask.

Do a rough titration – slowly add the solution from the burette into the conical flask, swirling constantly, until the indicator changes colour.

Repeat carefully until you have three readings within 0.1 mL of each other.

Perform calculations.

(b)



4.11.1 Various. Your list should include the following:

- A standard solution of anhydrous sodium carbonate was prepared and titrated against dilute hydrochloric acid.
- The hydrochloric acid was used as a secondary standard to titrate a dilute solution of sodium hydroxide.
- This sodium hydroxide solution was then used as a secondary standard to titrate a diluted solution of vinegar.

4.11.2 Sodium hydroxide pellets are deliquescent – they absorb water from the air; also atmospheric carbon dioxide will dissolve in sodium hydroxide solution. Thus it is not suitable for a primary standard, you cannot prepare the pure solution of definite concentration required for a primary standard.

4.11.3 (a) $\text{NaOH(aq)} + \text{HCl(aq)} \rightarrow \text{NaCl(aq)} + \text{H}_2\text{O(l)}$
10.0 mL

(b) $2\text{NaOH(aq)} + \text{H}_2\text{SO}_4\text{(aq)} \rightarrow \text{Na}_2\text{SO}_4\text{(aq)} + 2\text{H}_2\text{O(l)}$
20.0 mL

4.11.4 (a) (i) 105.99 g
(ii) 0.03 moles
(iii) 0.12 mol L^{-1}
(full value for further calculations is 0.122265308)

(b) (i) 3.07 moles
(ii) 6.13 moles
(iii) 0.5 mol L^{-1}

4.11.5 Graph starts off fairly level, sloping down slightly, indicating a slow drop in pH while the first 24 mL of HCl is added and the sodium hydroxide is being neutralised.

During the addition of the next 1.5 mL there is a sudden drop in pH from 11.31 to 3.00. The graph then levels out, curving down slightly as excess acid is added. The sharp drop in pH indicates the equivalence point – the reaction between the acid and base is complete.

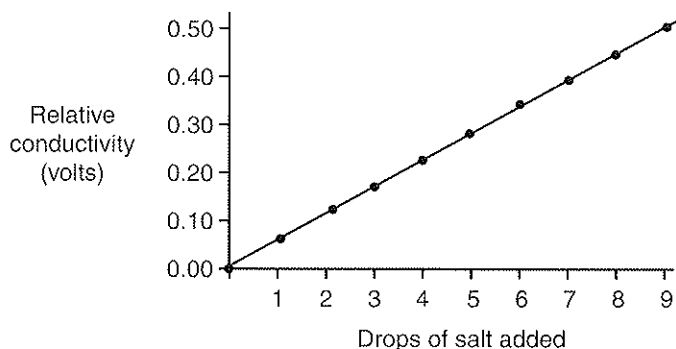
The middle of this sharp drop is at $\text{pH} \approx 7$, indicating a reaction between a strong base and a strong acid and the formation of a neutral salt (sodium chloride).

- 4.12.1 (a) Various, e.g. vinegar (contains ethanoic acid).
- (b) a solution of sodium hydroxide whose concentration you have already determined (a secondary standard).
- (c) The sodium hydroxide would be a secondary standard as its concentration was determined by titration against another solution of known concentration. Sodium hydroxide is a base so it would be a suitable substance to titrate against vinegar which is acidic.
- (d) Describe how you diluted the vinegar (or other commercial substance) to a concentration suitable for titrating against your standard.
- Describe the titration – a diagram would be useful here. Remember to state that you performed a rough titration and then three accurate titrations.
- (e) Calculate the moles of standard used. Use the equation to find the moles of the commercial product reacting. Calculate the concentration of the solution titrated. Correct for the dilution to calculate the concentration of the commercial product.
- (f) Various, e.g. if the commercial solution is coloured (e.g. orange juice) you would have to dilute it until the colour would not affect the colour of the indicator.

4.12.2 Various, e.g. you may have used a pH probe to detect changes in pH as the solutions were mixed. This could be connected to a data logger to record the results and a computer used to graph the results.

Alternatively, you may have used a pH meter or conductivity meter to detect pH changes which you recorded and graphed.

4.12.3 (a)



- (b) Relative conductivity increases as concentration increases.
- (c) The same direct relationship should be found, however as MgCl_2 dissociates to produce three ions (instead of two for NaCl) the conductivity values should all be higher.

4.13.1 Various, e.g.

Identify	Describe	Explain
Take care when diluting an acid.	When diluting an acid, always add acid to water (never add water to acid).	Dilution of acids is an exothermic reaction, a great deal of heat can be released. If water is being added to acid, this heat may make the acid boil and spit, spattering the person with concentrated acid which is corrosive and thus can burn skin and benches. If acid is being added to water, the water will boil and spit rather than the acid so this procedure is safer.
When pouring acids, hold the bottle with your hand over the label.	This prevents acid being dripped onto the label.	Acid is corrosive so if it drips onto the label, the label will be damaged and may become difficult to read. A bottle containing a chemical that cannot be identified is dangerous.
Cover up when using acids, especially concentrated acids.	Acids are corrosive and may get on skin or in eyes.	Acids may splash into eyes or be spilled on skin or clothing. Acids are corrosive so they can burn the skin and eyes. Always wear protective clothing, including goggles when using acids to prevent acids coming into contact with skin or eyes.

- 4.13.2** Neutralisation reactions (acid + base to form water) do replace corrosive acids and bases with neutral water and salts that may be neutral or slightly acidic or basic, e.g.



However, there are problems associated with using a neutralisation reaction to treat spills:

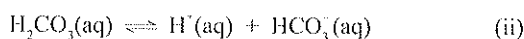
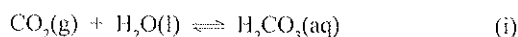
- The neutralisation of acids is an exothermic process, so you should not add hydroxide to neutralise acid spilt on the skin. The neutralisation reactions will cause even more heat to be released and intensify the burn. Instead it is better to use copious quantities of running water.
- To neutralise spills of concentrated acid on floors and benches the addition of a strong base will produce much heat which can intensify the damage. Also the base itself is corrosive and causes further damage. It is better to add the weak base sodium hydrogen carbonate (NaHCO_3) and then wash with a large volume of water. Being a weak base, this will produce heat slowly as it neutralises the acid and the addition of water will have a cooling effect as well as diluting the corrosive acid.

Thus neutralisation reactions should not be used on skin, but may be useful for spills on benches or floors, provided the weak base sodium hydrogen carbonate is used rather than a strong base.

- 4.14.1** (a) A buffer is a solution that resists rapid changes in pH when an acid or base is added.

A buffer is usually made by mixing together a weak acid and the salt of that acid. A buffer can also be made using a weak base and the salt of that base. Examples of buffers are acetic acid (CH_3COOH) and sodium acetate (CH_3COONa), carbonic acid (H_2CO_3) and sodium bicarbonate (NaHCO_3) or ammonia (NH_3) and ammonium chloride (NH_4Cl).

- (b) Blood is a buffered solution, containing carbonic acid and sodium bicarbonate. This allows its pH to stay constant at 7.4. Carbon dioxide dissolves in our blood, forming carbonic acid, a weak acid that ionises to form the hydrogen carbonate ion.



If more CO_2 dissolves in blood, more H^+ ions will form. This would push the pH down, but the equilibrium moves to the left, minimising the change.

If the pH is tending to increase (not acidic enough) more carbonic acid molecules ionise and the equilibrium moves to the right to produce more hydrogen ions, again minimising the change.

Because carbonic acid is a weak acid, any increase or decrease in the amount of acid present only produces a very small change in hydrogen ion concentration and thus a very small change in pH.

- (c) In living things, the rate of chemical reactions is controlled by biological catalysts called enzymes. Each enzyme acts at a specific pH, so a slight change in pH can prevent an enzyme working and thus greatly affect the chemistry of living cells. A high hydrogen ion concentration slows down energy production in cells and interferes with the functioning of calcium ions. pH must be controlled in living things and this control is achieved by buffering systems.

- 4.14.2** (a) In both graphs the pH starts at the same point (about 3), it increases as sodium hydroxide is added and eventually levels off at around pH 10 to 11.
- (b) The graph for citric acid shows a much steeper rise in pH than for the buffered lemonade. The buffered Lemon-fizz lemonade neutralised almost three times as much sodium hydroxide as the unbuffered citric acid.
- (c) The unbuffered citric acid showed a fast change in pH as sodium hydroxide was added whereas the buffered lemonade changed pH slowly with the addition of hydroxide, as less base is needed to neutralise the unbuffered citric acid.

5.1.1

	Alkanols	Alkanoic acids
Functional group	Hydroxyl group -OH	Carboxyl group – COOH $\begin{array}{c} \text{O}^{\delta-} \\ \parallel \\ -\text{C} \\ \\ \text{O} - \text{H}^{\delta+} \end{array}$
General formula	ROH where R is a saturated carbon chain with general formula $\text{C}_n\text{H}_{2n+1}$	RCOOH where R is a saturated carbon chain with general formula $\text{C}_n\text{H}_{2n+1}$
Polarity	Polar due to OH group.	More polar than alkanols due to COOH group.
Other	Soluble in water. Higher melting and boiling points than similar alkanes, alkenes or alkynes due to hydrogen bonding.	Soluble in water. Higher melting and boiling points than similar alkanols. Alkanoic acids have more hydrogen bonds than similar sized alkanols.

5.1.2

Name	Structural formula	Molecular formula
Methanoic acid	$\begin{array}{c} \text{O} \\ \parallel \\ \text{H} - \text{C} \\ \\ \text{OH} \end{array}$	HCOOH
Ethanoic acid	$\begin{array}{c} \text{H} \\ \\ \text{H} - \text{C} - \text{C} \\ \quad \parallel \\ \text{H} \quad \text{OH} \end{array}$	CH_3COOH
Propanoic acid	$\begin{array}{c} \text{H} \quad \text{H} \\ \quad \\ \text{H} - \text{C} - \text{C} - \text{C} \\ \quad \quad \parallel \\ \text{H} \quad \text{H} \quad \text{OH} \end{array}$	$\text{C}_2\text{H}_5\text{COOH}$
Butanoic acid	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \\ \text{H} - \text{C} - \text{C} - \text{C} - \text{C} \\ \quad \quad \quad \parallel \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{OH} \end{array}$	$\text{C}_3\text{H}_7\text{COOH}$
Pentanoic acid	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \quad \\ \text{H} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} \\ \quad \quad \quad \quad \parallel \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{OH} \end{array}$	$\text{C}_4\text{H}_9\text{COOH}$
Hexanoic acid	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \quad \quad \\ \text{H} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} \\ \quad \quad \quad \quad \quad \parallel \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{OH} \end{array}$	$\text{C}_5\text{H}_{11}\text{COOH}$
Heptanoic acid	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \quad \quad \quad \\ \text{H} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} \\ \quad \quad \quad \quad \quad \quad \parallel \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{OH} \end{array}$	$\text{C}_6\text{H}_{13}\text{COOH}$
Octanoic acid	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \quad \quad \quad \quad \\ \text{H} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} \\ \quad \quad \quad \quad \quad \quad \quad \parallel \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{OH} \end{array}$	$\text{C}_7\text{H}_{15}\text{COOH}$

5.1.3

- (a) 2-hexanol
 (b) 2-pentanol
 (c) methanoic acid
 (d) propanoic acid

5.2.1 For all four graphs, as the number of carbon atoms in the molecule increases, the boiling point gets higher.

Alkanoic acids have higher boiling points than alkanes, alkenes or alkanols with the same number of carbon atoms.

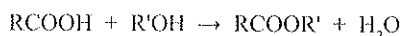
5.2.2 Boiling points are higher for alkanoic acids than for alkanols of similar size.

Alkanols and alkanoic acids both have dispersion forces and they also both have dipole forces, due to their polar bonding, and hydrogen bonds between molecules. However, each alkanol molecule only has one oxygen atom, whereas each alkanoic acid molecule has two oxygen atoms, so the acids are more polar and they form more hydrogen bonds between molecules.

Thus alkanoic acids have the strongest forces between molecules and their boiling points are the highest. More energy is needed to break the forces between the molecules.

5.3.1 Esterification is a reaction between an alkanoic acid and an alkanol.

alkanoic acid + alkanol \rightarrow ester + water



5.3.2 A molecule of water condenses out.



Notice that the bridging oxygen atom in the ester comes from the alkanol so the oxygen atom of the alkanol ends up in the ester, not in the water.

- 5.3.3 (a) propyl ethanoate + water
(b) methyl butanoate + water
(c) ethyl propanoate + water
(d) octyl methanoate + water

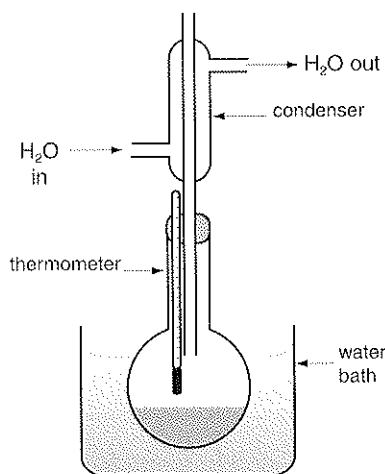
5.4.1 Organic, alkanol, acid, liquids, volatile

5.4.2

Alkanol	Alkanoic acid	Ester
Ethanol	Methanoic acid	Ethyl methanoate
Propanol	Methanoic acid	Propyl methanoate
Butanol	Methanoic acid	Butyl methanoate
Pentanol	Propanoic acid	Pentyl propanoate
Octanol	Ethanoic acid	Octyl ethanoate

Name of homologous series	Functional group	General formula	Name and formula of an example
Alkene	Double C=C bond	C_nH_{2n}	Propene $\begin{array}{c} \text{H} & \text{H} \\ & & \\ \text{H} & - \text{C} = & \text{C} - & \text{C} - \text{H} \\ & & & \\ & & & \text{H} \end{array}$
Alkane	Single C-C bonds	C_nH_{2n+2}	Propane C_3H_8
Alkanol	Hydroxyl (OH) group	$C_nH_{2n+1}OH$	Propanol C_3H_7OH
Alkanoic acid	Carboxyl (COOH) group	RCOOH Where R is an alkyl chain with formula C_nH_{2n+1}	Propanoic acid C_2H_5COOH
Ester	$\begin{array}{c} \text{O} \\ // \\ \text{R} - \text{C} \\ \backslash \\ \text{O} - \text{R}' \end{array}$	RCOOR' where R and R' are carbon chains	Butyl propanoate $\begin{array}{c} \text{O} \\ // \\ \text{C}_2\text{H}_5\text{C} \\ \backslash \\ \text{O} - \text{C}_4\text{H}_9 \end{array}$

- 5.5.1 (a) The use of a catalyst in a chemical reaction to change the rate of the reaction.
- (b) Concentrated sulfuric acid is used as a catalyst because it absorbs water. Esterification is an equilibrium reaction and removing one of the products (water) moves the reaction towards the products (to replace the removed product – i.e. Châtelier's principle) thus increasing the yield of ester.
- 5.6.1 (a) Refluxing is the process of heating a reaction mixture in a container with a cooling condenser attached vertically. This is necessary to prevent the loss of the volatile reactants and products to the environment. It allows the reaction to be carried out at a higher temperature than would otherwise be possible.



- (b) The water enters at the bottom, rises to fill the jacket and leaves through the upper outlet. This ensures that the jacket remains filled with circulating water for efficient cooling.

- 5.7.1**
- (a) Various, e.g. Reactants = ethanol and ethanoic acid. Products = ethyl ethanoate and water.
 - (b) Prevents loss of volatile reactants and products, allowing the reaction to take place at higher temperatures and thus increasing the rate of reaction.
 - (c) Reactants and products are volatile and flammable, so the reaction vessel must be heated with a water bath rather than with a naked flame. No stopper on reaction vessel. Good ventilation and use fume cupboard to avoid inhaling vapours. Wear safety goggles.
 - (d) Reflux – condenser is connected vertically.
Distillation – condenser is connected horizontally.

- 5.8.1**
- (a) ethyl ethanoate $\text{CH}_3\text{COOC}_2\text{H}_5$
 - (b) pentyl ethanoate $\text{CH}_3\text{COOC}_5\text{H}_{11}$ banana flavour

5.9.1 Various, e.g.

Name	Formula	Use
Ethyl butanoate	$\text{C}_4\text{H}_7\text{COO C}_2\text{H}_5$	Pineapple ester
Pentyl ethanoate	$\text{CH}_3\text{COOC}_5\text{H}_{11}$	Banana ester

Notes

This image shows a single sheet of white paper with horizontal ruling lines. The lines are evenly spaced and run across the width of the page. There are no margins, text, or other markings on the paper.