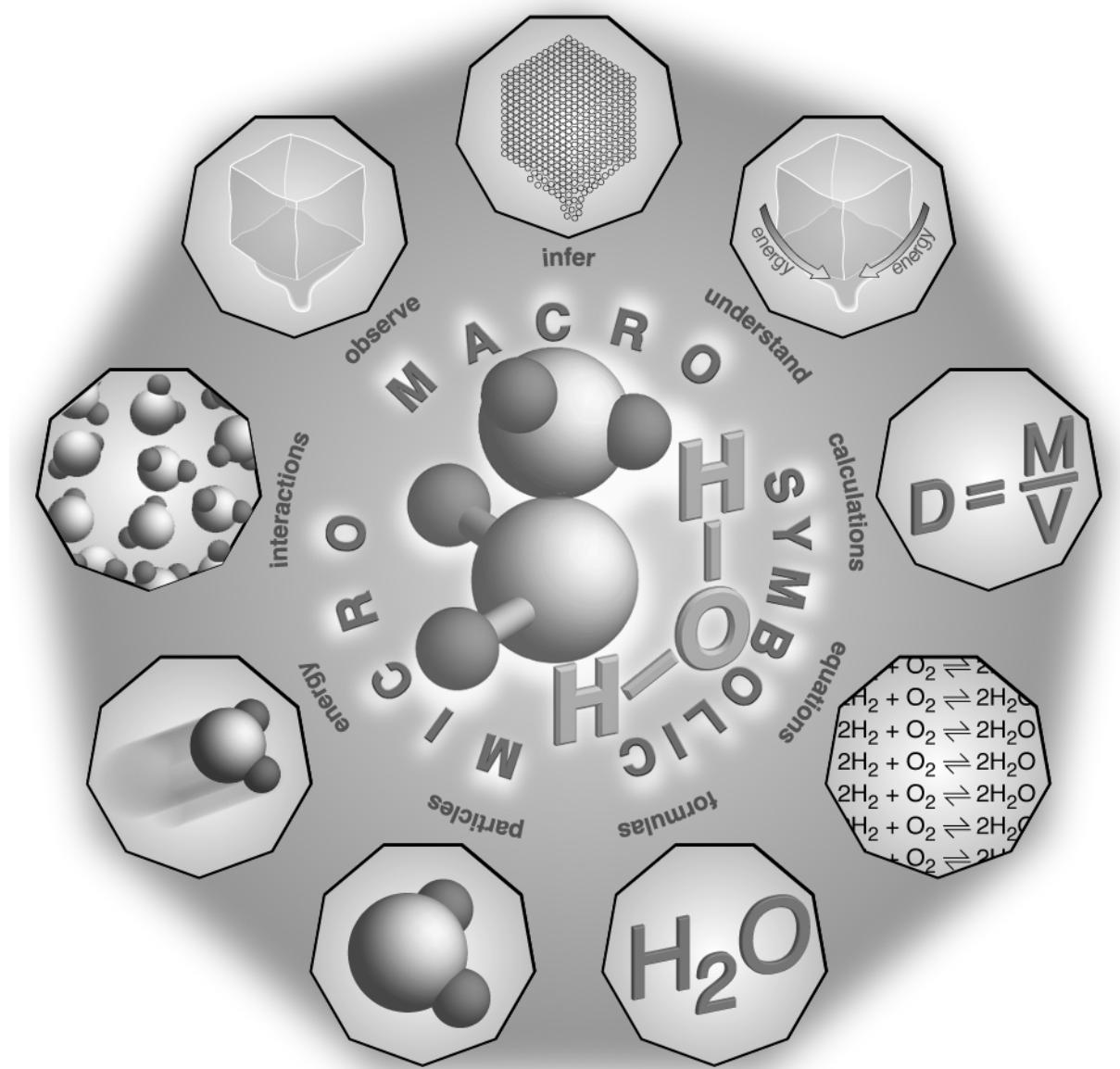




## The acidic environment



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infer  
understand



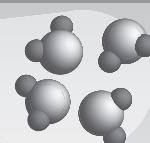
**SYMBOLIC**

$H_2O$  formulas  
equations  
calculations



**MICRO**

particles  
energy  
interactions



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**MACRO**observe  
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## Module overview

The extracts below are from the *Chemistry Stage 6 Syllabus*© Board of Studies NSW, originally issued 1999. The most up-to-date version is to be found at [http://www.boardofstudies.nsw.edu.au/syllabus99/syllabus2000\\_list.html](http://www.boardofstudies.nsw.edu.au/syllabus99/syllabus2000_list.html)

‘Acidic environments exist everywhere. The human body has a slightly acidic skin surface to assist in disease control and the digestive system operates in both acidic and alkaline environments to assist the breakdown of the biopolymers constituting food. Indeed, the metabolic processes of organisms are well adapted to life in acidic or alkaline environments’.

The oxidation of organic compounds which make up living things produces organic acids. Thus the surface of human skin where organic substances are oxidised is slightly acidic whereas the blood bathing the skin is close to neutral. More extreme acidity and basicity is found in the human digestive system. The stomach contents are very acid to favour the initial breakdown of protein. In the upper intestines, slightly basic conditions favour breakdown of carbohydrates and fats and further breakdown of smaller lengths of protein molecules. All living things have buffer chemicals to resist changes in alkalinity (basicity) and acidity.

‘Many industries use acidic and basic compounds for a wide range of purposes and these compounds are found in daily use within the home. Because of this, an awareness of the properties of acids and bases is important for safe handling of materials. Currently, concerns exist about the increased release of acidic and basic substances into the environment and the impact of these substances on the environment and the organisms within those environments’.

Many foods are acidic while most water soluble cleaning products are basic (alkaline). An awareness of how to handle acidic and basic substances and safe use of products that are hazardous to eyes and skin are important. The release of acidic gases from internal combustion engines into the atmosphere, acidic solutions from mining, basic solutions from mineral processing and acidic and basic substances from industrial processes affect life. Some populations may die out from this exposure causing a dramatic change in the balance of organisms in particular environments.

# Outcomes

The main course outcomes to which this *The acidic environment* module contributes are:

A student:

- H1 evaluates how major advances in scientific understanding and technology have changed the direction or nature of scientific thinking
- H2 analyses the ways in which models, theories and laws in chemistry have been tested and validated
- H6 explains reactions between elements and compounds in terms of atomic structures and periodicity
- H7 describes the chemical basis of energy transformations in chemical reactions
- H8 assesses the range of factors which influence the type and rate of chemical reactions
- H9 describes and predicts reactions involving carbon compounds
- H10 analyses stoichiometric relationships
- H11 justifies the appropriateness of a particular investigation plan
- H12 evaluates ways in which accuracy and reliability could be improved in investigations
- H13 uses terminology and reporting styles appropriately and successfully to communicate information and understanding
- H14 assesses the validity of conclusions from gathered data and information
- H15 explains why an investigation is best undertaken individually or by a team
- H16 demonstrates and justifies positive values about and attitude towards both the living and non-living components of the environment, ethical behaviour and a desire for critical evaluation of the consequences of the applications of science.

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# Indicative time

This module is designed to take a minimum of thirty hours. There are a number of practical activities. Organising materials and equipment for carrying out all these activities could take additional time but in doing so, you will better understand the type of work that chemists do.

# Resources

Materials and equipment you need to carry out activities are listed below. Access to a computer and the worldwide web are important for the study of modern chemistry. An important skill to develop in chemistry is planning ahead and thinking things through before carrying out the action. Make sure the resources you need are available when you start an activity. For details see the activity in the appropriate part.

## For Part 1 you will require:

- a few leaves of a red cabbage (sometimes called purple cabbage because of its purple colour)
- a small saucepan to be heated or a dish that is microwave safe
- water
- a means of heating the saucepan or a microwave oven
- two teaspoons and a tablespoon
- white vinegar (5% w/w acetic acid CH<sub>3</sub>COOH)
- baking soda or sodium hydrogen carbonate NaHCO<sub>3</sub>
- three clear glass containers such as glasses/beakers/test tubes
- red cabbage solution
- at least six different household substances (try the kitchen, laundry, cleaning solutions, bathroom, hardware chemicals that are water soluble). Read the label and heed the label of every substance that you are to use. Do not use the substance if it is marked CAUSTIC or POISON or S6, S7 or S8. (See the *Appendix* in Part one for an explanation of the S1 to S8 poisons schedule.)
- protective eyewear – either glasses or safety goggles
- old stainless steel knives or spoons for adding chemicals and stirring
- disposable paper towel for drying washed knives or spoons
- clear glass or plastic containers to hold red cabbage solution and added substances

- five different coloured flower petals. Include white flowers in your collection. Try to find out the names of the flowers.
- 15 test tubes or 5 hot water resistant containers and 10 other transparent containers
- container of hot water to hold 5 test tubes or 5 hot water resistant containers
- white vinegar
- solution of baking soda (sodium hydrogen carbonate  $\text{NaHCO}_3$ ) or washing soda (sodium carbonate  $\text{Na}_2\text{CO}_3$ )
- graphite pencil
- glove or small towel
- red cabbage solution or universal indicator paper
- bleach solution.

**For Part 2 you will require:**

- coloured pencils
- a means of weighing equipment to at least the nearest gram
- an unopened small bottle (or can) of soda water with liquid level as low as possible – 250 mL or 300 mL size is adequate
- either a source of dry heat such as an electric hotplate or a saucepan in which the soda water can be gently warmed and a dry towel and a thermometer (the warming method)  
OR 1g of table salt per 50 mL of soda water (the salting method)
- For Exercise 2.2 you will need computer access to the worldwide web [www](http://www).

**For Part 3 you will require:**

- pH meter probe/system
- indicator such as universal indicator paper or red cabbage solution
- a variety of household chemicals already in water solution or able to dissolve in water.

**For Part 4 you will require:**

- pH meter or broad range indicator such as universal indicator paper or red cabbage solution (the red cabbage range of colours can be seen at [www.lmpc.edu.au/science](http://www.lmpc.edu.au/science))
- a range of salt solutions containing 5 % w/w salt

Use salts such as:

- sodium chloride, NaCl (table salt)
- sodium hydrogen carbonate, NaHCO<sub>3</sub> (baking soda)
- sodium carbonate, Na<sub>2</sub>CO<sub>3</sub> (washing soda)
- sodium hydrogen sulfate, NaHSO<sub>4</sub> (toilet powder cleaner)
- magnesium sulfate, MgSO<sub>4</sub> (Epsom salt)
- two volumetric flasks (250 mL and 500 mL) OR two plastic storage bottles from supermarket used to store acidic or basic liquids.
- 25 mL bulb pipette OR 3 mL syringe (without needle) from a pharmacy
- 50 mL burette and stand and clamp for holding the burette OR 5 to 10 mL syringe (without needle) from a pharmacy
- 250 mL conical flask OR a small transparent glass or plastic container such as a 40 mL conical medicine measure as receiving flask
- three small beakers (50 – 250 mL) or three small open transparent glass or plastic containers for holding solutions
- phenolphthalein indicator solution (can be made using laxatives containing phenolphthalein: crush laxative, swirl with methylated spirits, decant liquid and add an equal volume of water) OR red cabbage solution
- white vinegar
- sodium hydroxide OR a spray can of caustic oven cleaner that contains sodium hydroxide from a supermarket and a plastic bucket
- safety goggles or glasses
- 3 intact aspirin (acetylsalicylic acid) tablets; do not use soluble aspirin tablets which are acetylsalicylate salts.
- 100 mL of white vinegar (undiluted) of known concentration calculated in the previous activity
- sodium hydroxide (pellets, flakes or granules)
- takeaway polystyrene cup and plastic lid with hole for a straw
- 0–110°C thermometer
- container for measuring out 100 mL of liquid

**For Part 5 you will require:**

EITHER THE FOLLOWING IF PREPARING ESTERS IN A LAB

- a selection of alcohols such as ethyl, propyl, butyl and pentyl
- acetic (ethanoic) acid and butyric (butanoic) acid
- concentrated sulfuric acid in a dropper bottle
- test tubes to hold the mixtures of alcohol, acid and conc. H<sub>2</sub>SO<sub>4</sub>
- a means of heating the test tubes such as a container of heated water on an electric hotplate

OR THE FOLLOWING IF PREPARING AN ESTER USING  
HOUSEHOLD CHEMICALS

- 10 aspirin (not soluble aspirin) tablets; each tablet has a mass of about 0.3 g aspirin and contains starch binder and cellulose filler
- 15 mL methylated spirits
- 10 mL concentrated hydrochloric acid; often sold as muriatic acid at hardware stores for cleaning mortar off bricks or lowering the pH of swimming pools
- 1 g of boric (boracic) acid B(OH)<sub>3</sub>; a solid acid sold in pharmacies and used in eyebaths, wound ointments and ant poisons
- two plastic spoons
- means of measuring or estimating liquid volumes to 0.5 mL
- filter and filter papers
- paper towel or filter papers for drying product
- 3 test tubes and 3 small (100 to 250 mL) beakers
- a source of hot water and a container in which test tubes can be placed
- thermometer (0 to 100°C)
- crushed ice or small iceblocks

# Icons

The following icons are used within this module. The meaning of each icon is written beside it.



The hand icon means there is an activity for you to do.  
It may be an experiment or you may make something.



You need to use a computer for this activity.



Discuss ideas with someone else. You could speak with family or friends or anyone else who is available.  
Perhaps you could telephone someone?



There is a safety issue that you need to consider.



There are suggested answers for the following questions  
at the end of the part.



There is an exercise at the end of the part for you to complete.

# Glossary

The following words, listed here with their meanings, are found in the learning material in this module. They appear bolded the first time they occur in the learning material.

acidic	having the properties of an acid
acidity	extent of acid properties
alkali	water soluble base; hydroxides of alkali metal (Group I) or alkaline earth metal (Group II) and ammonia are called alkalis
alkaline	basic; having the properties of a solution of an alkali
amphiprotic	ionic or molecular species which can act both as an acid by donating protons and a base by accepting protons
amphoteric	chemical which shows both acidic and basic properties
analytical reagent (AR) grade	Very high grade chemical with very low impurity levels
Arrhenius acid	substance that produces hydrogen ions in water
Arrhenius base	substance that produces hydroxide ions in water
basic	having the properties of a base
basicity	extent of basic properties
bromothymol blue	indicator used for end point of titrations of a strong acid with a strong base
Bronsted-Lowry acid	proton donor
Bronsted-Lowry base	proton acceptor
buffer	mixture of a weak acid and its conjugate base able to resist change in pH
calibration	checking of the accuracy of a measuring instrument
carbonated water	solution of carbon dioxide in water; soda water
carboxylic acid	acid containing the –COOH functional group
caustic	able to burn, corrode or destroy human tissue

concentrated solution	solution containing a large amount of solute in the solvent
condensation reaction	reaction between two molecules in which a small molecule such as water condenses out
condenser	equipment to change gas or vapour to liquid
conjugate acid	acid formed by addition of $H^+$ to a base
conjugate base	base formed by removal of $H^+$ from an acid
decarbonate	removal of carbon dioxide from a solution or solid
degas	to free from gas
delocalised electrons	electrons that move between atoms in a molecule or structure eg. the benzene ring of six carbon atoms has six delocalised electrons
denature	to deprive something of its nature; proteins can be denatured by physical or chemical changes that change the protein structure
dilute	solution containing a small amount of solute in the solvent
diprotic	acid able to release up to two protons per molecule
end point	point in a titration at which an indicator changes colour
equilibrium	situation in a reversible reaction in a closed system where the rate of forward reaction equals the rate of reverse reaction
equivalence point	point in a titration at which the reactants have completely reacted together in the ratio of mole quantities given in the balanced equation for the reaction
ester	molecule which contains the $-COOC$ functional group
ester hydrolysis	reaction of an ester with water to form a carboxylic acid and alkanol
esterification	reaction forming an ester from a carboxylic acid and an alkanol
flavour	taste
fragrance	odour
functional group	atom or group of atoms that react in a characteristic way

hazardous	having the potential to harm human health
homogeneous	uniformly mixed; each small part of the mixture has the same composition
hydronium ion	ion formed by hydrogen ion in water; represented by the formula $\text{H}_3\text{O}^+$
hydroxy	$-\text{OH}$ functional group
indicator	substance that changes (usually in colour) when there is a change in the chemical nature of a system; an acid-base indicator changes colour according to pH
ionisation	formation of an ion or ions as when an acid molecule releases a hydrogen ion and forms a negative ion
laboratory reagent (LR) grade	chemical with moderate impurity levels suitable for general laboratory use
Le Chatelier's principle	Principle used to predict the direction of shift of the equilibrium of a reversible reaction in a closed system; 'equilibrium shifts to oppose change'
legend	explanation of information given in a graph or table
Lewis acid	electron pair acceptor
Lewis base	electron pair donor
litmus	dye from lichen that turns red in acid solution and blue in basic solution
meniscus	upper surface of a liquid
methyl orange	indicator used for end point of titrations of a strong acid with a weak base
millivoltmeter	meter measuring voltage in millivolts
moderated	adjusted; HSC school internal assessments are moderated according to the performance of the school's candidates in the exam paper
monoprotic	acid able to release only one proton per molecule
MSDS	material safety data sheet; risk and safety information about the properties and uses of a chemical substance
neutral	between acidic and basic; having neither acidic nor basic properties

neutralisation	reaction between an acid and base to form a salt and water
pH	'power of H <sup>+</sup> '; a measure of hydrogen ion concentration; pH = $-\log_{10}[\text{H}^+]$
phenolphthalein	indicator used for end point of titrations of a strong base with a weak acid
pH meter	millivoltmeter with scale marked in pH units that can be connected to a pH probe
pH probe	pH sensitive electrode system
pipette	tube for transferring a fixed volume of liquid from one container to another
pipette filler	rubber bulb with three valves that can be attached to a pipette and used to fill and empty the pipette
Polyatomic ion	ion consisting of more than one atom covalently bonded together
polyprotic	acid able to release more than one proton per molecule
reflux/refluxing	process of boiling liquid in a flask with a condenser attached so that vapours are condensed and returned to the liquid
shelf life	time for which a product may remain on the shelf and still be marketed
slurry	suspension of an insoluble solid and liquid like a thin paste
soda water	solution of carbon dioxide in water; carbonated water
standardisation	to bring to or make an established standard
standard solution	solution of accurately known concentration and composition
strong acid	an acid that is completely ionised in solution
strong base	a base that is completely ionised in solution
titration	technique in which a solution of unknown concentration is reacted with a solution of known concentration (a standard solution) until equivalence point is reached
titre	volume added from a burette in a titration
transgenic	organism containing a gene from a different type of organism

translocated gene	the gene that is moved from one organism to a different type of organism
triprotic	acid able to release up to three protons per molecule
volumetric analysis	quantitative analysis using volumes to find the unknown concentration of a solution
volumetric chemistry	branch of chemistry involving measurements of volumes of reactants and products established by Gay-Lussac
volumetric flask	flask used to prepare a solution of accurately known or measured concentration
weak acid	acid that is incompletely ionised in solution
weak base	base that is incompletely ionised in solution

# Chemistry

HSC Course

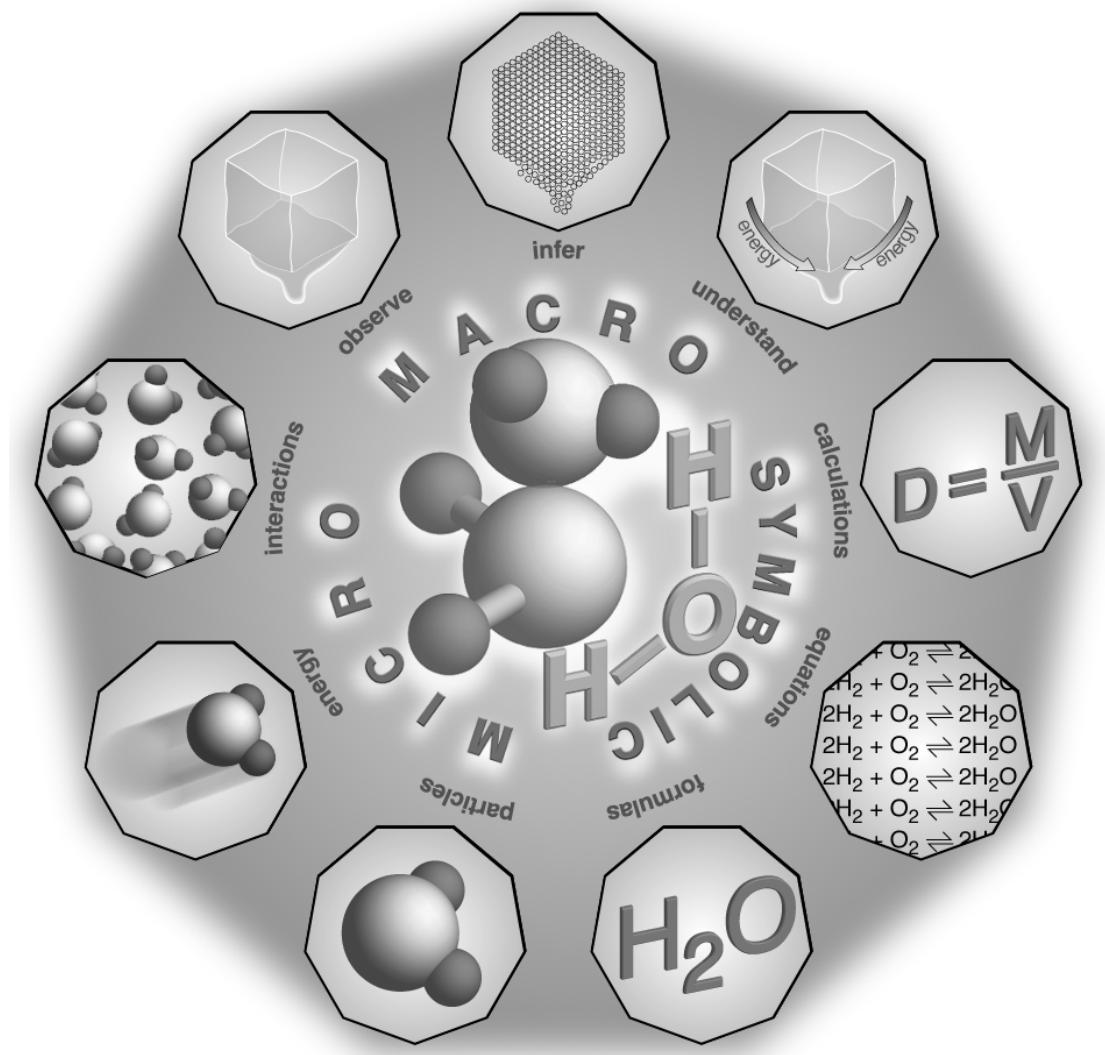
Stage 6

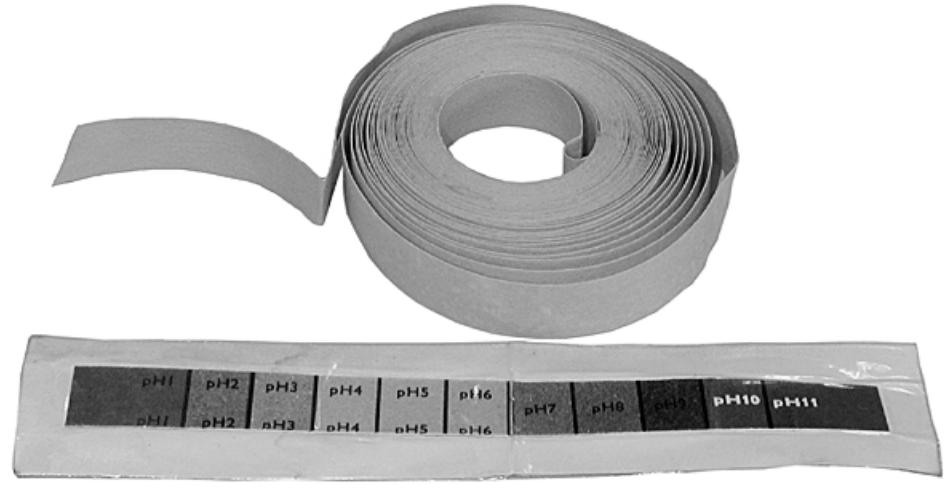
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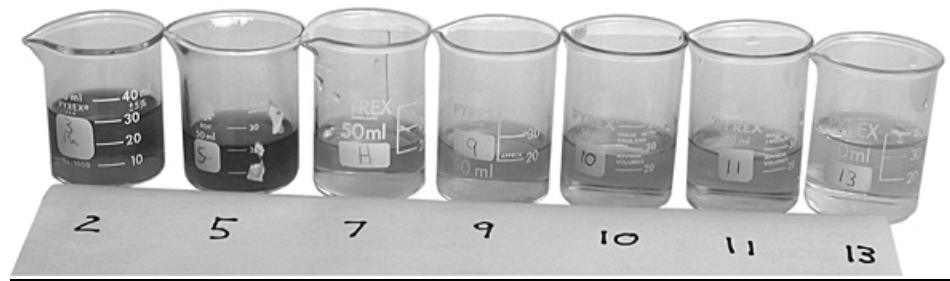
## The acidic environment

### Part 1: Indicators





Universal indicator paper and colour chart



Range of colours produced by red cabbage solution at different pHs



Colour images of these photographs can be accessed at the website  
[www.lmpc.edu.au/science](http://www.lmpc.edu.au/science)

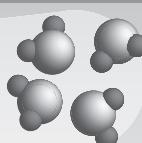
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understand



**SYMBOLIC**

$H_2O$  formulas  
equations  
calculations



**MICRO**

particles  
energy  
interactions

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# Introduction

Acid-base **indicators** are chemicals that can change colour with change in the acidic environment. Many of these indicators occur naturally in plants. Litmus is an indicator extracted from lichens that turns red in acid conditions and blue in base conditions. Hydrangea plants produce blue flowers in acid soils but pink flowers in basic soils. These natural indicators as well as synthetically made indicators are useful in testing for **acidity** – acid properties – and **basicity** – basic properties.

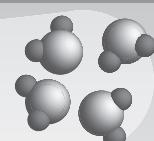
In Part 1 you will be given opportunities to learn to:

- classify common substances as acidic, basic or neutral
- 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
- identify and describe some everyday uses of indicators including the testing of soil acidity/basicity
- define Le Chatelier's principle
- identify factors which can affect the equilibrium in a reversible reaction.

In Part 1 you will be given opportunities to:

- perform a first-hand investigation to prepare and test a natural indicator
- identify data and choose resources to gather information about the colour changes of a range of indicators
- solve problems by applying information about the colour changes of indicators to classify some household substances as acidic, neutral or basic.

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# What is an indicator?

An indicator is a chemical which changes, usually in colour, when the chemical environment changes. Most indicators in chemistry are acid-base indicators. They change colour as the acid-base environment changes. In this module, and most other contexts in chemistry, indicator stands for acid-base indicator.

Here are a number of definitions for an acid-base indicator.

- A dye sensitive to acid or base.
- Organic molecules or ions with different coloured acidic and basic forms that change colour typically over about two pH units.
- A dye sensitive to concentration of hydrogen ions.
- Organic molecules whose colour depends on the acidity or basicity of the solution in which they are dissolved.
- Organic molecules changing colour over a narrow pH range.

You may already understand parts of some of these definitions. By the end of this module hopefully you will understand all of them. If you already understand all of them don't jump to the next module! There is a lot more to understanding acidic behaviour than just knowing about indicators!



Rearrange these definitions from the most complex to the simplest:

most complex

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simplest

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Compare your answer with the suggested answers.

# Preparing and testing a natural indicator



## What you will need:

- a few leaves of a red cabbage (sometimes called purple cabbage because of its colour)
- a small saucepan to be heated or a dish that can be microwaved
- water
- a means of heating the saucepan or a microwave oven
- two teaspoons and a tablespoon
- white vinegar (5% w/w acetic acid CH<sub>3</sub>COOH)
- baking soda or sodium hydrogen carbonate NaHCO<sub>3</sub>
- three clear glass containers such as glasses/beakers/test tubes.

## What you will do:

- 1 Tear or slice the leaves into small pieces and place them into the saucepan/microwavable dish. Just cover the red cabbage with water.
- 2 Boil for 20 minutes or microwave for 10 minutes. If water boils away replace the water.
- 3 Let the mixture cool until warm then remove the cabbage with the tablespoon. You have prepared a solution of a natural indicator.

Describe the colour of the solution \_\_\_\_\_

- 4 While the mixture is cooling half fill the three containers with water. Add the same amount of red cabbage solution to each of the containers.
- 5 Add a teaspoon of white vinegar to the first container and stir to dissolve. Colour of the solution? \_\_\_\_\_
- 6 Add a teaspoon of baking soda to the third container and stir to dissolve. Colour of the solution? \_\_\_\_\_

Save unused red cabbage solution for later activities.

## Conclusion:

Red cabbage solution is \_\_\_\_\_ (colour) in neutral water, \_\_\_\_\_ in acidic vinegar solution and \_\_\_\_\_ in basic baking soda solution.

If you cannot obtain some red cabbage here are some alternatives you could try:

- beetroot (beetroot in tins is usually in a vinegar solution)
- blueberries or blackcurrants or raspberries or mulberries
- blue or red grapes
- carrots or pumpkin or other orange or yellow coloured vegetable
- cherries or strawberries
- red radish
- rhubarb.

If it is difficult to extract an indicator from solid material you could use

- a cup of tea
- dark grape juice or red wine.



While you have the equipment available, consider carrying out Exercise 1.1. This requires you to extract natural dyes from flower petals and investigate changes in colour between acidic vinegar and basic baking soda solutions.

A science class tested a range of products from a supermarket. The students added water solutions of the product to red cabbage solution and observed the following results:

Product	Colour changes
caustic soda drain cleaner	purple → blue-green → yellow
diluted white vinegar 10% v/v	purple → pink
rust remover	purple → red
dry acid for pools	purple → light red
borax	purple → blue-green → green
sodium hydrogen carbonate	purple → blue-green
washing soda	purple → blue-green → green

Colour change observed as a solution of the product is slowly added to the red cabbage solution.

The purple red cabbage solution was **neutral** – between acidic and basic.



Use the colour changes of the red cabbage indicator solution to arrange the product solutions from most basic (yellow) to most acidic (red):

Acidity/basicity	Colour of solution	Product solution
most basic		
moderately basic	green	
weakly basic	blue-green	
neutral	purple	
weakly acidic	pink	
moderately acidic	light red	
most acidic		

Check your answers.

## Classifying household substances



### What you will need:

- red cabbage solution
- at least six different household substances (try the kitchen, laundry, cleaning solutions, bathroom, hardware chemicals that are water soluble). Read the label and heed the label of every substance that you are to use. Do not use the substance if it is marked CAUSTIC or POISON or S6, S7 or S8. (See the *Appendix* for an explanation of the S1 to S8 poisons schedule.)
- protective eyewear – either glasses or safety goggles
- old stainless steel knives or spoons for adding chemicals and stirring
- disposable paper towel for drying washed knives or spoons
- clear glass or plastic containers to hold red cabbage solution and added substances.

**What you will do:**

- 1 Always add a small amount of the household substance to a larger amount of the red cabbage solution.
- 2 Stir with the knife or spoon to help dissolve the substance.
- 3 Wash and completely dry the knife or spoon after use.
- 4 Record the colour of the red cabbage solution then decide whether the household substance is acidic, neutral or basic.

**Results:**

Household substance	Colour of red cabbage solution	Acidic/neutral/basic

Do particular groups of household substances such as foods or cleaning liquids show mostly acidic, neutral or basic properties?

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Foods are usually acidic – this provides some protection against micro-organisms.

Cleaning solutions are usually basic – this enables them to dissolve fats and greases.

Do Exercises 1.2 and 1.3 now.



# A brief history of indicators

The ancient Greeks used indicators over 2 000 years ago. Litmus had been extracted from certain lichens and the hardness of basic water was tested by its ability to change the color of wine.

In the 1700s Robert Boyle, famous author of *The Sceptical Chymist*, used blue 'syrup of violets' to detect acids (red) and bases (green). Gay-Lussac who laid the foundations of **volumetric chemistry** used litmus.

In the nineteenth century (1800s) syrup of violets was used to detect when equal amounts of acid and base had reacted. Litmus and cabbage liquor were commonly used to distinguish acidic and basic substances. By the end of the 1800s chemists knew that indicators were organic molecules that could exist in a molecular form and an ionic form. These two forms had different colours.

Different indicators change colour at different levels of acidity and basicity.

The **pH** scale (pH stands for power or concentration of hydrogen ion  $H^+$ ), typically from 0 (extremely acid) to 14 (extremely basic), enables a quantitative comparison of indicators:

Indicator	very acid	weakly acid	neutral	weakly basic	very basic											
<b>pH</b>																
	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14	
bromothymol blue	-----	yellow	-----	→	6.0	7.6	←	blue	-----	-----	-----	-----	-----	-----	-----	-----
litmus	-----	red	-----	→	5.0	8.0	←	blue	-----	-----	-----	-----	-----	-----	-----	-----
methyl orange	--	red	--	→	3.1	4.4	←	yellow	-----	-----	-----	-----	-----	-----	-----	-----
phenolphthalein	-----	colourless	-----	→	8.3	10.0	←	red	-----	-----	-----	-----	-----	-----	-----	-----

The pH range between →6.0 and 7.6← is where a colour change occurs.

Some indicator solutions such as universal indicator are a mixture of different indicators. Universal indicator changes colour from red to orange to yellow to green to blue as the pH rises. This sequence of colours should be familiar as the order of colours in the light spectrum.

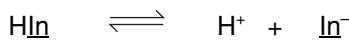
Red cabbage solution contains three different indicators. These produce the range of colour changes seen.

The activities you have undertaken in this Part 1 so far have involved observing and working at the macro level. The symbolic level uses equations and formulas.

## Symbolic representation of indicators

An indicator can be represented by  $\text{HIn}$ . H represents a hydrogen atom that can be released as a hydrogen ion  $\text{H}^+$ .  $\text{In}$  represents the rest of the organic molecule.

A chemical equilibrium exists in solution between the  $\text{HIn}$  molecule and the  $\text{In}^-$  ion. The  $\text{HIn}$  molecule and the  $\text{In}^-$  ion are different colours. The presence or absence of the hydrogen ion produces different colours.



	<b>molecule</b>	<b>ion</b>
bromothymol blue	yellow	blue
litmus	red	blue
methyl orange	red	yellow
phenolphthalein	colourless	red



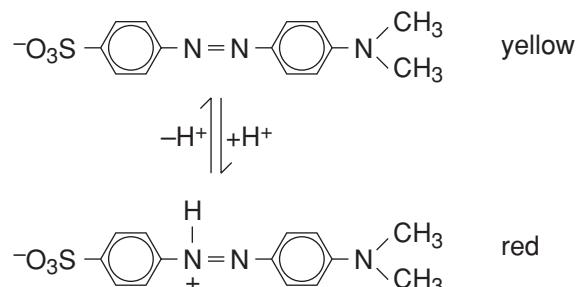
Remember mixing two coloured paints to make another colour when you were younger? If you can remember, then you should find it easier to predict the colour of indicator solutions containing about equal numbers of  $\text{HIn}$  molecules and  $\text{In}^-$  ions.

- bromothymol blue: yellow + blue  $\rightarrow$  gr\_\_\_\_\_
- litmus: red + blue  $\rightarrow$  p\_\_\_\_\_
- methyl orange: red + yellow  $\rightarrow$  \_\_\_\_\_
- phenolphthalein: colourless + red  $\rightarrow$  faint pink

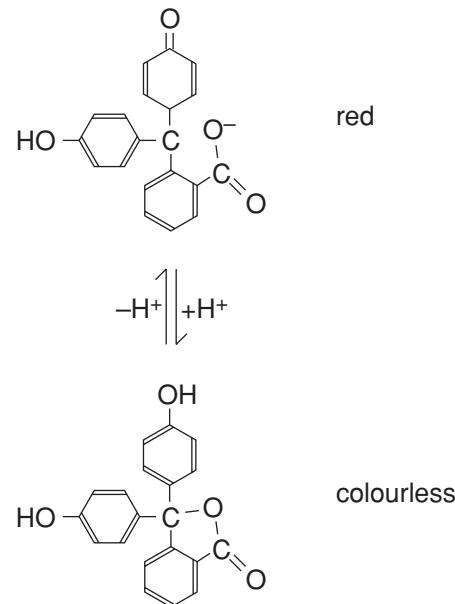
Check your answers.

- A benzene ring of six carbon atoms in a hexagon can be represented as 
- This shows alternating double and single covalent bonds.
- Another representation shows that all bonds in the hexagon are equal and contain **delocalised electrons** represented by a large ring 
- In these structures any carbon which is not shown with an attached bond has an attached hydrogen atom. If the benzene ring is not attached to any other atoms it is C<sub>6</sub>H<sub>6</sub>.
- Different texts use different representations. The examples of indicator structures below illustrate this.

In<sup>-</sup> and HIn forms of methyl orange:



In<sup>-</sup> and HIn forms of phenolphthalein:



# Uses of indicators

Indicators can be used in solution or adsorbed onto paper.

- **Soil testing**

A small sample of soil is mixed with universal indicator solution. White insoluble barium sulfate  $\text{BaSO}_4$  is sprinkled on the surface.  $\text{BaSO}_4$  provides a white background against which the colour of the indicator solution seeping through from the soil can be more easily seen. The colour seen is matched to a pH value on an indicator colour-pH chart.

pH of soil is important in determining what crops can be grown. Productivity per hectare can be significantly affected by the soil pH. Agricultural lime, calcium carbonate  $\text{CaCO}_3$ , increases the basicity. Manure increases the acidity.

Fertiliser addition can change soil pH. Liquid ammonia  $\text{NH}_3$  increases basicity while ammonium salts containing the ammonium ion  $\text{NH}_4^+$  increase acidity.

- **Maintaining swimming pool water pH close to 7.4 for human skin and eye comfort**

pH can be measured by taking a sample of water and testing with indicator solution or indicator paper.

If the pH needs to be lowered, that is made more acid, acids such as hydrochloric acid solution or solid sodium hydrogen sulfate  $\text{NaHSO}_4$  can be added.

The pH can be raised, that is made more basic, by adding sodium carbonate  $\text{Na}_2\text{CO}_3$  or sodium hydrogen carbonate  $\text{NaHCO}_3$ .

- **Maintaining aquarium water at a suitable pH**

Some aquarium animals are very sensitive to the pH of water. Indicators can be used to check that the pH of the aquarium water lies in a suitable range. Sea water is about pH 8.5.

- **Detecting the end point of quantitative acid-base reactions**

In Part 4 you will determine the concentration of an unknown acid solution by reacting a known volume with a known volume and concentration of a base solution. The end point of the reaction is detected by change in colour of a suitable indicator.

## MACRO

observe  
infer  
understand



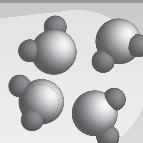
## SYMBOLIC

$\text{H}_2\text{O}$  formulas  
equations  
calculations



## MICRO

particles  
energy  
interactions



# Le Chatelier's principle

In 1885 a French chemist proposed a very important principle for predicting the effect of change on a system at equilibrium. This principle for reversible reactions became known as **Le Chatelier's principle**:

'If a system is at equilibrium and a change is made so the system is no longer at equilibrium, more reaction will occur to counteract the change.'

This occurs because a disturbed system will adjust the concentrations of reactants and products to re-establish equilibrium.

Le Chatelier derived this principle from studying chemical systems. However, there are many examples of this principle in fields as diverse as economics, human sociology and animal behaviour!

The principle can be stated more simply as:

'equilibrium shifts to oppose external change'

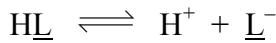
or more bluntly by people who don't like change and oppose it as:

'nature is a real &?@!#'

To emphasise the sorts of changes that occur in chemical systems, Le Chatelier's principle can be stated as:

'The concentrations of reactants and products in a mixture at equilibrium will alter so as to counteract any change in concentration, temperature or gas pressure'.

Consider a neutral solution of the indicator litmus. The litmus solution will be purple in colour – intermediate between the red  $\text{H}\underline{\text{L}}$  of an acidic solution and the blue  $\underline{\text{L}}^-$  of a basic solution. The constant purple colour of the solution shows that there is equilibrium between  $\text{H}\underline{\text{L}}$  and  $\underline{\text{L}}^-$ .



If acid, a solution containing  $\text{H}^+$ , is added to the purple neutral solution the change is an increased concentration of  $\text{H}^+$ . The equilibrium shifts to oppose this change by  $\text{L}^-$  reacting with some of the excess  $\text{H}^+$  and forming a higher concentration of  $\text{HL}$ . This causes the purple solution to turn red.

Conditions	Red $\text{HL}$ concentration	Blue $\text{L}^-$ concentration	Colour of solution
initial	moderate	moderate	purple
after adding $\text{H}^+$	high	low	red

The equilibrium shown is said to shift position to the left.

Consider another neutral solution of the indicator litmus.

If base, a solution that reacts with  $\text{H}^+$ , is added to the purple neutral solution the change is a decreased concentration of  $\text{H}^+$ . The equilibrium shifts to oppose this change by  $\text{HL}$  reacting to release  $\text{H}^+$  and at the same time forming a higher concentration of  $\text{L}^-$ . This causes the purple solution to turn blue.

Conditions	Red $\text{HL}$ concentration	Blue $\text{L}^-$ concentration	Colour of solution
initial	moderate	moderate	purple
after removing $\text{H}^+$	low	high	blue

The equilibrium is said to shift position to the right.



Complete the table and statement below for the addition of acid to a green solution of bromothymol blue       $\text{HB} \rightleftharpoons \text{H}^+ + \text{B}^-$

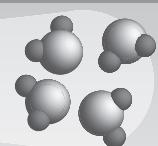
Conditions	Yellow $\text{HB}$ concentration	Blue $\text{B}^-$ concentration	Colour of solution
initial	moderate	moderate	
after adding $\text{H}^+$			

The equilibrium is said to shift position to the \_\_\_\_\_.

Check your answers.



Do Exercises 1.4 now.

**MACRO**observe  
infer  
understand**SYMBOLIC** $H_2O$  formulas  
equations  
calculations**MICRO**particles  
energy  
interactions

## Appendix

### Poisons schedule

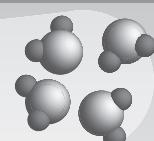
National Health and Medical Research Council (NHMRC) recommendation regarding classification of drugs and poisons.

Commercially available products have appropriate signal words on labels. S1 is no longer used.

Availability is in the order: no schedule, 5, 6, 2, 3, 7, 4, 8, 9

Schedule	Signal words to identify poisons
S2	PHARMACY MEDICINE
S3	PHARMACIST ONLY MEDICINE
S4	PRESCRIPTION ONLY MEDICINE
S5	CAUTION
S6	POISON
S7	DANGEROUS POISON
S8	CONTROLLED DRUG; POSSESSION WITHOUT AUTHORITY ILLEGAL
S9	PROHIBITED; ONLY AVAILABLE FOR APPROVED MEDICAL AND SCIENTIFIC RESEARCH



**MACRO**observe  
infer  
understand**SYMBOLIC** $H_2O$  formulas  
equations  
calculations**MICRO**particles  
energy  
interactions

# Suggested answers

## What is an indicator?

most complex	Organic molecules or ions with different coloured acidic and basic forms that change colour typically over about two pH units
	Organic molecules changing colour over a narrow pH range
	Organic molecules whose colour depends on the acidity or basicity of the solution in which they are dissolved
	A dye sensitive to concentration of hydrogen ions.
simplest	A dye sensitive to acid or base.

## Preparing and testing a natural indicator

Acidity/basicity	Colour of solution	Product solution
most basic	yellow	<i>caustic soda drain cleaner</i>
moderately basic	green	<i>washing soda, borax</i>
weakly basic	blue-green	<i>sodium hydrogen carbonate</i>
neutral	purple	<i>red cabbage solution</i>
weakly acidic	pink	<i>diluted white vinegar 10% v/v</i>
moderately acidic	light red	<i>dry acid for pools</i>
most acidic	red	<i>rust remover</i>

## Symbolic level representation of indicators

bromothymol blue: yellow + blue  $\rightarrow$  green

litmus: red + blue  $\rightarrow$  purple

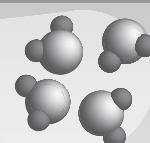
methyl orange: red + yellow  $\rightarrow$  orange

phenolphthalein: colourless + red  $\rightarrow$  faint pink

## Le Chatelier's principle

Conditions	Yellow HB concentration	Blue B <sup>-</sup> concentration	Colour of solution
initial	moderate	moderate	green
after adding H <sup>+</sup>	high	low	yellow

The equilibrium is said to shift position to the left. (*This refers to the equation: HB  $\rightleftharpoons$  H<sup>+</sup> + B<sup>-</sup>*).

**MACRO**observe  
infer  
understand**SYMBOLIC** $H_2O$  formulas  
equations  
calculations**MICRO**particles  
energy  
interactions

# Exercises – Part 1

Exercises 1.1 to 1.4

Name: \_\_\_\_\_

## Exercise 1.1: Extracting natural indicators from flower petals

Now you really need to get organised!. You will need to handle a lot of containers and will need to report your results. But you are guaranteed a colourful time!

**What you will need:**

- five different coloured flower petals. Include white flowers in your collection. Try to find out the names of the flowers.
- 15 test tubes or 5 hot water resistant containers and 10 other transparent containers
- container of hot water to hold 5 test tubes or 5 hot water resistant containers
- white vinegar
- solution of baking soda (sodium hydrogen carbonate  $NaHCO_3$ ) or washing soda (sodium carbonate  $Na_2CO_3$ )
- graphite pencil
- glove or small towel.

**What you will do:**

- 1 Get organised! You will be handling a number of containers, different solutions and labelling will be important. If you are using test tubes and they have a small paint spot you can write on this with a graphite pencil. Otherwise stick labels on the containers or organise them in a labelled area. If you do not have labels small squares of masking tape work well.
- 2 Keep the same colour petals from different flowers separate.

- 3 Tear the petals of a particular flower into small portions. Push the petals to the bottom of the test tube with the blunt end of the pencil. Add water until the petals are just covered. Label the test tube. Repeat for the other flowers.
- 4 Place the test tubes in hot, even boiling, water and leave for 10 minutes
- 5 Using a glove or small towel to prevent burning of your hand remove a test tube. Carefully pour the liquid extract into another test tube then discard the solid matter inside the original test tube. Repeat for all the heated test tubes.
- 6 Divide each extract solution amongst two other test tubes/containers so that you have three lots of each solution.
- 7 Add equal amounts of vinegar to each of the first lots.
- 8 Add water to each of the second lots
- 9 Add baking soda or washing soda solution to each of the third lots.
- 10 Record the colours you can see in the results table below.

**Results table:**

Name or description of flower	Colour in acid	Colour of extract	Colour in base	Suitable as an acid-base indicator?

## **Exercise 1.2: Identifying data and choosing resources to gather information about an indicator**

Here is a list of indicators not covered in the Part 1 notes:

congo red  
cresol red  
erythrosin  
orange IV  
thymol blue  
methyl red  
bromocresol purple  
phenol red  
alizarin yellow  
malachite green  
methyl violet

Using whatever resources you can access – texts, encyclopedias, libraries, internet search engines – select three of these indicators and complete the following table.

Indicator	Formula	pH	colour	pH	colour

### **Exercise 1.3: Investigating the pH of your mouth**

Does the pH of your mouth change through the day? Is there a pattern of pH change that is maintained from day to day? Does the pH vary much according to what you have just drunk or eaten?

Here is an opportunity to devise your own investigation to find out more about your own body. All you need is an indicator that can measure a range of pH from about 4 to 10 such as red cabbage solution or universal indicator.

Add some of your saliva to a small amount of red cabbage solution or a small piece of universal indicator paper. Do not put red cabbage solution or universal indicator paper in your mouth.

Dispose of the test solutions or test papers responsibly. You should be the only person to fill, handle, empty and wash out any glassware used with your saliva before placing it in an undiluted bleach (sodium hypochlorite) solution. Read and follow warnings on the labels of the bleach solution.

Describe what you did and what you found out about the pH of your mouth using an acid-base indicator:

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## **Exercise 1.4: Understanding the HSC – draft performance band descriptors**

The following extracts are from *Draft Performance Bands, Chemistry*. 1999. Board of Studies, NSW. This publication can be downloaded from:  
[http://www.boardofstudies.nsw.edu.au/syllabus99/syllabus2000\\_list.html](http://www.boardofstudies.nsw.edu.au/syllabus99/syllabus2000_list.html)

The Higher School Certificate awarded at the end of your course will use performance band descriptors on your HSC to describe your attainments. The average of your exam mark and **moderated** assessment mark will be used to award a composite HSC mark. This mark will place you in a descriptor band describing the standard the mark represents.

The descriptor bands given below are draft and could be modified.

The descriptor bands are Band 6 (the highest band) down to Band 2.

The descriptor bands shown below are not in order. Read through these band descriptors for chemistry then try to correctly label them Band 6, Band 5, Band 4, Band 3 and Band 2.

Remember Band 6 describes the highest attainment.

**Band** \_\_\_\_\_

- demonstrates a sound knowledge and understanding of the concepts of the chemistry course content including context, prescribed focus areas and domain
- describes concepts and information clearly in written, graphical and diagrammatic forms such as structural and electron-dot formulae, and applies these concepts in familiar situations
- demonstrates a broad ability to carry out calculations and or substitute into equations, to use relevant symbols and units when manipulating chemical data including stoichiometric data, and to construct balanced chemical equations
- displays proficiency in selecting relevant data from information given in written, tabular, graphical and diagrammatic form
- describes correct apparatus for a particular chemical measurement and has an adequate understanding of experimental methodology.

**Band** \_\_\_\_\_

- demonstrates a limited knowledge and understanding of the chemistry course content including context, prescribed focus areas and domain
- recalls elementary terminology and formulae related to some areas of chemistry
- makes simple substitutions of data in chemical calculations
- describes simple safety precautions in experimental procedure

**Band** \_\_\_\_\_

- demonstrates an extensive knowledge and understanding of the concepts of the chemistry course content including context, prescribed focus area and domain
- displays an outstanding ability to describe and explain chemistry concepts, including abstract ideas, clearly and accurately, and to apply the concepts to unfamiliar situations
- applies a high level of critical thinking skills in developing appropriate solutions to problems involving a long sequence of related tasks
- analyses, evaluates and extrapolates chemical data effectively, identifies complex relationships, quantifies explanations and descriptions, and synthesizes information to draw conclusions
- communicates succinctly, logically and sequentially using a variety of scientific formats
- demonstrates a high level ability to design an experimental procedure.

**Band** \_\_\_\_\_

- demonstrates a basic knowledge and understanding of the concepts of the chemistry course content including context, prescribed focus areas and domain
- uses simple chemistry definitions, terms, diagrams and graphs to communicate understanding of chemistry concepts
- substitutes data from information given in written, tabular, graphical and diagrammatic form, and manipulates basic chemical data including stoichiometric data

**Band** \_\_\_\_\_

- demonstrates a thorough knowledge and understanding of the concepts of the chemistry course including context, prescribed focus areas and domain
- effectively communicates a detailed understanding of chemistry concepts using appropriate chemistry terminology and scientific formats, and applies the concepts to unfamiliar situations
- analyses information given in written, tabular, graphical and diagrammatic forms and relates this to other relevant information
- displays competence in manipulating equations to solve problems involving a number of steps
- demonstrates a thorough knowledge of the use of appropriate experimental procedures.

# Chemistry

HSC Course

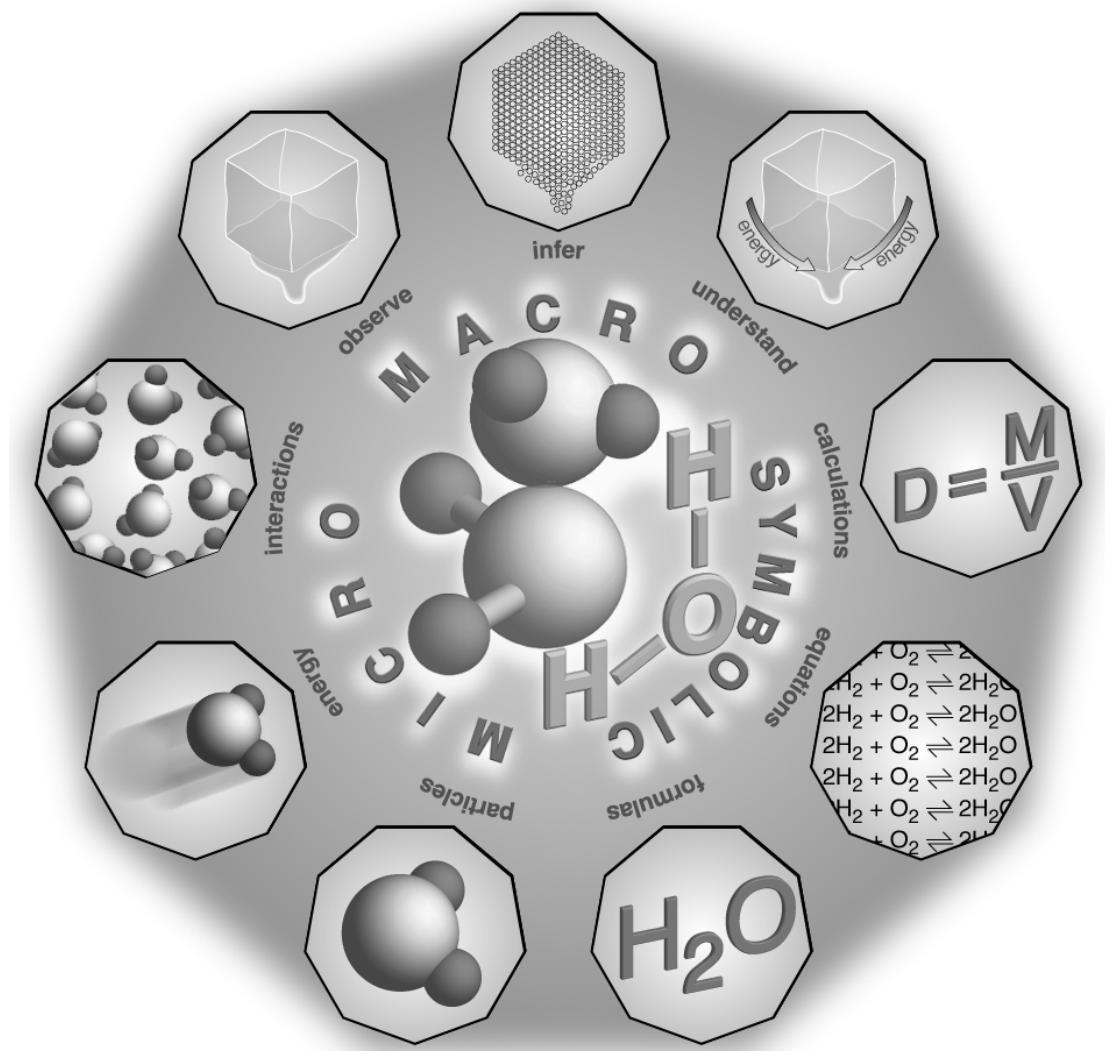
Stage 6

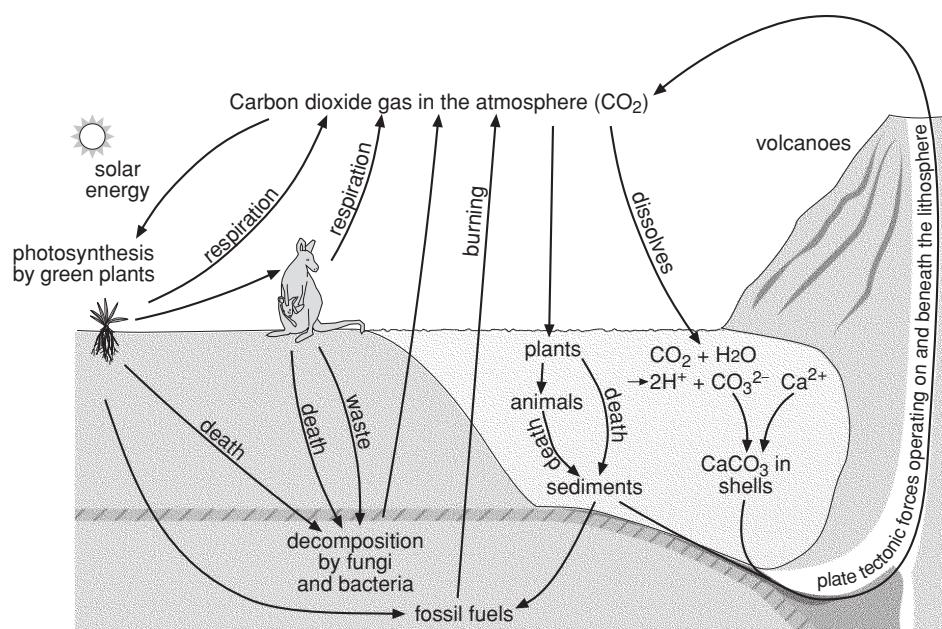
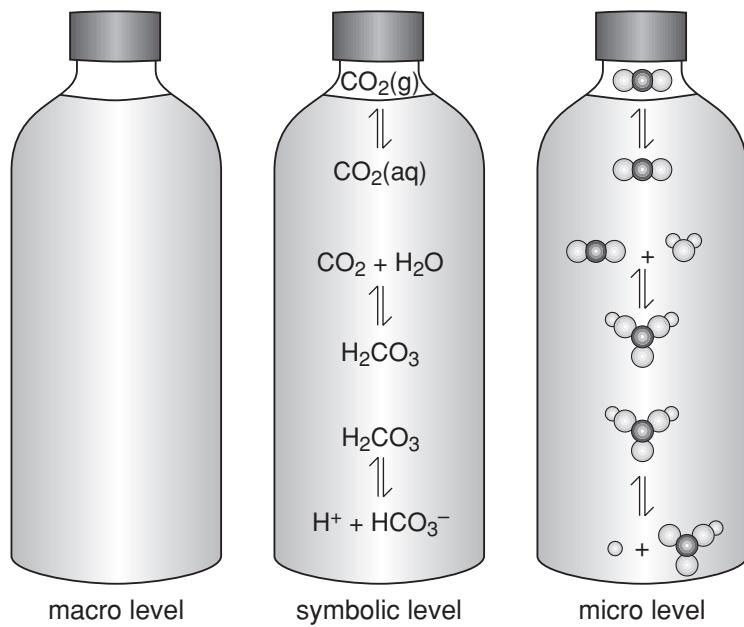
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DEPARTMENT  
OF EDUCATION  
AND TRAINING

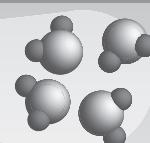


## The acidic environment

### Part 2: Acid in the atmosphere

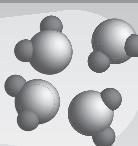




**MACRO**observe  
infer  
understand**SYMBOLIC** $H_2O$  formulas  
equations  
calculations**MICRO**particles  
energy  
interactions

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**MACRO**observe  
infer  
understand**SYMBOLIC** $H_2O$  formulas  
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energy  
interactions

# Introduction

The air around us contains acidic gases from natural sources such as carbon dioxide from respiration, nitrogen dioxide from lightning strikes and sulfur dioxide from volcanoes. The concentrations of acidic oxides in the atmosphere have been increasing over the last two hundred years as underground fossil fuels have been brought to the surface and burnt as a source of energy. Carbon dioxide ( $CO_2$ ) is the main product of fossil fuel combustion while high temperature combustion can provide sufficient energy for nitrogen and oxygen in the air to join forming NO which is easily oxidised to  $NO_2$ . Coal and petroleum typically contain a few per cent of sulfur which burns forming  $SO_2$  and can be further oxidised to  $SO_3$ .

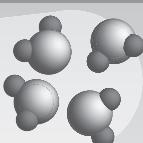
In Part 2 you will be given opportunities to learn to:

- identify oxides of non-metals which act as acids and describe the conditions under which they act as acids
- analyse the position of these non-metals in the Periodic Table and generalise about the relationship between position of elements in the Periodic Table and acidity/basicity of oxides
- describe the solubility of carbon dioxide in water under various conditions as an equilibrium process and relate this to Le Chatelier's principle
- identify natural and industrial sources of sulfur dioxide and oxides of nitrogen
- describe, using equations, examples of chemical reactions which release sulfur dioxide and chemical reactions which release oxides of nitrogen
- assess the evidence which indicates increases in atmospheric concentration of oxides of sulfur and nitrogen
- calculate volumes of gases given masses of some substances in reactions, and calculate masses of substances given gaseous volumes, in reactions involving gases at  $0^\circ C$  and  $101.3\text{ kPa}$  or  $25^\circ C$  and  $101.3\text{ kPa}$
- explain the formation and effects of acid rain.

In Part 2 you will be given opportunities to:

- 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
- process and present information from secondary sources to describe the properties of sulfur dioxide and the oxides of nitrogen
- choose resources, gather and 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.

Extracts from *Chemistry Stage 6 Syllabus* © Board of Studies NSW, originally issued 1999. The most up-to-date version can be found on the Board's web site at [http://www.boardofstudies.nsw.edu.au/syllabus99/syllabus2000\\_list.html](http://www.boardofstudies.nsw.edu.au/syllabus99/syllabus2000_list.html)

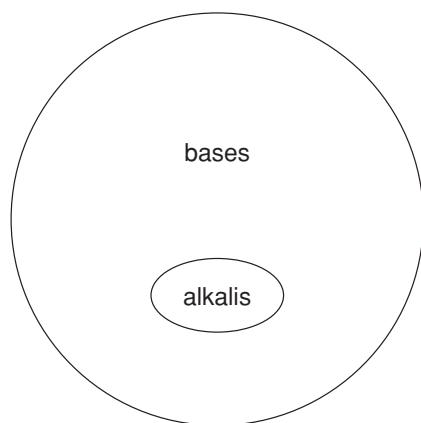
**MACRO**observe  
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energy  
interactions

# Acidity and basicity of oxides

The acidity/basicity of an oxide can be tested by

- dissolving soluble oxide in water and testing the pH of the solution with indicator paper such as litmus paper or universal indicator paper
- placing oxide in a small amount of alkaline or acidic solution and see if the acid or base properties are neutralised.

An **alkali** is a water soluble base. Most alkalis are metal hydroxides such as sodium hydroxide ( $NaOH$ ) or potassium hydroxide ( $KOH$ ). Alkalies are water soluble, cheap and the most commonly used bases in industry and school laboratories. Not all bases are water soluble.



Venn diagram showing that alkalis are a type of base.

An **alkaline** solution contains a high concentration of hydroxide ion  $OH^-$  which can react with  $H^+$  from acid forming water:  $H^+ + OH^- \rightarrow H_2O$



What conclusion would you come to if:

- 1 a white solid oxide dissolves in water producing a solution that turns litmus blue?

- 2 a colourless oxide gas dissolves in water producing a solution with a pH of 3?
- 
- 3 an oxide dissolved in an alkaline solution forming a solution with pH 7?
- 
- 4 a black oxide dissolved in an acidic solution of pH 3.6 producing a blue solution with a pH of 5?
- 
- 5 a gaseous oxide dissolved in an acidic solution of pH 3.6 producing a solution of pH 3.6?
- 

Check your answers.

The water insoluble oxides of beryllium, aluminium, zinc, tin and lead react with both alkaline solutions and acidic solutions.

Reaction with and dissolution in alkaline solution demonstrates acidic properties. Reaction with and dissolution in acid solution demonstrates basic properties. Chemicals which show both acidic and basic properties are called **amphoteric**. (Note that the terms alkaline and basic are used interchangeably to describe solutions of pH > 7.)



Elements can be divided into five groups for the acid-base behaviour of their oxides.

- Acidic oxides such as CO<sub>2</sub>, NO<sub>2</sub>, N<sub>2</sub>O<sub>3</sub>, N<sub>2</sub>O<sub>5</sub>, P<sub>2</sub>O<sub>3</sub>, P<sub>2</sub>O<sub>5</sub>, SO<sub>2</sub>, SO<sub>3</sub> react with water producing an acidic solution.

What do you notice about these elements combined with oxygen?

---

- Basic oxides such as Li<sub>2</sub>O, MgO, CaO, TiO<sub>2</sub>, Cr<sub>2</sub>O<sub>3</sub>, MnO<sub>2</sub>, Fe<sub>2</sub>O<sub>3</sub>, NiO, CuO, Ag<sub>2</sub>O react with acid solutions neutralising them.

What do you notice about these elements combined with oxygen?

---

Check your answers.

- Amphoteric oxides BeO, Al<sub>2</sub>O<sub>3</sub>, ZnO, SnO<sub>2</sub>, PbO react with both acid solutions and alkaline solutions.
- Neutral oxides CO, NO, N<sub>2</sub>O dissolve slightly in water without changing the pH.
- Elements which do not form oxides – the noble gases.



Use coloured pencils to divide the periodic table below into these five types of oxides. Draw a key to label the five groups. The boxes marked with a ? are elements made in such small amounts that the nature of their oxides are unknown.

A large grid puzzle consisting of several interconnected sections. On the left, there's a vertical stack of 5x2 rectangles. To the right of this is a long horizontal row of 10x2 rectangles. Further right is a section of 10x5 rectangles. Below these are two rows of 10x2 rectangles each. At the bottom, there's a dashed line with question marks above it, followed by a solid line with question marks. The entire grid is composed of thin black lines on a white background.

Use the trends you can see in your coloured and labelled periodic table to complete these statements.

- 1 Metal oxides are usually \_\_\_\_\_
  - 2 Non-metal oxides are usually \_\_\_\_\_
  - 3 Amphoteric oxides are formed by \_\_\_\_\_ elements close to the non-metals in the Periodic Table.
  - 4 The only elements that do not form oxides are the \_\_\_\_\_ in group \_\_\_\_\_.

Check your answers.

Non-metal oxides are compounds of a non-metal with a non-metal. Such compounds are made up of covalently bonded molecules. The intramolecular covalent bonds are strong but the intermolecular forces between the molecules are much weaker. The weak intermolecular forces are easily overcome so most non-metal oxides are gases at room temperature. The few exceptions are water which can hydrogen bond and is liquid at room temperature and large, heavy oxide molecules such as phosphorus (V) oxide which is a solid.

Non-metal oxides can act as acids when they come in contact with water and a base. Always moisten the indicator paper or use an aqueous indicator solution when testing an oxide to see if it is acidic or basic.

Compared with other gases non-metal oxide gases are much more soluble because they react with the water. The activity following requires you to measure the amount of gaseous carbon dioxide in soft drink.

**MACRO**observe  
infer  
understand**SYMBOLIC** $H_2O$  formulas  
equations  
calculations**MICRO**particles  
energy  
interactions

# Carbon dioxide

## Decarbonating soda water

**Soda water** or **carbonated water** is a solution of carbon dioxide in water. Soft drinks are carbonated water with added flavours and sweeteners. Some cola drinks even contain phosphoric acid  $H_3PO_4$ ! Read the labels of what you drink.

**What you will need:**

- a means of weighing equipment to at least the nearest gram
- an unopened small bottle (or can) of soda water with liquid level as low as possible; 250 mL or 300 mL size is adequate
- either a source of dry heat such as an electric hotplate or a saucepan in which the soda water can be gently warmed and a dry towel and a thermometer (the warming method)

OR 1g of table salt per 50 mL of soda water (the salting method)

**What you will do:**

- 1 Weigh the unopened bottle of soda water \_\_\_\_\_ g
- 2 Slowly remove the cap controlling the release of bubbles so that the soda water does not foam out the top and you lose soda water. If this happens you will need to buy another bottle and start all over again! Observe what happens to the release of bubbles when you retighten the cap.
- 3 Reweigh the bottle including the cap \_\_\_\_\_ g
- 4 Calculate the mass of carbon dioxide lost when the cap was removed
- 5 Now choose either the warming method or the salting method.

**Warming method – uses the reduced water solubility of gas with temperature rise**

- 6 Carefully stand the open bottle on an electric hotplate or in a saucepan of heated water. Insert the thermometer. Stirring encourages release of gas bubbles.
- 7 Heat the soda water bottle to about 38°C (human body temperature).
- 8 Remove the soda water bottle, dry it well with a towel, then reweigh the bottle and its cap.

Weight of bottle and cap after warming. \_\_\_\_\_ g

Calculate the change in weight due to loss of CO<sub>2</sub> gas to the air:

**Salting method – uses the addition of ions which attract water molecules and reduce the availability of water able to dissolve gas**

- 6 Weigh about 1 g of table salt for each 50 mL of soda water accurately. \_\_\_\_\_ g
- 7 Add the salt very slowly and carefully to the soda water so that the water does not **degas** too quickly. This is especially important in the early stages. If soda water foams over the top of the container you will lose weight and need to start all over again! Control spraying so that spray does not leave the container and affect weighing.
- 8 Reweigh the bottle and cap after salting. \_\_\_\_\_ g

Allowing for the weight of salt added to the soda water, calculate the change in weight due to loss of CO<sub>2</sub> gas to the air:

**Further calculations:**

In Part 5 of the *Metals* module in the Preliminary course you learnt that the molar volume of a gas at 25°C and 101.3 kPa pressure was 24.5 L.

In this activity you have measured the mass of CO<sub>2</sub> gas released. To calculate the volume that this mass of CO<sub>2</sub> gas would occupy you need to convert the mass in g to moles using:

$$n = \frac{m}{M} \quad \text{where } n = \text{no. of moles}$$

m = mass in g

M = molar mass in g/mole

The molar mass of CO<sub>2</sub> = 12 + (2 x 16) = 44 g/mole

$$\text{Thus no. of moles of CO}_2 = \frac{m}{44}$$



Once you have calculated the number of moles of CO<sub>2</sub> multiply this by 24.5 L to calculate the volume of gas at 25°C and 101.3 kPa.

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**Conclusions:**

Compare the volume of gas released with the volume of liquid soda water which contained that amount of gas.

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What do these volumes illustrate about the distance between CO<sub>2</sub> particles in the gas phase compared with the distance between CO<sub>2</sub> particles in solution?

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Explain why bubbles of CO<sub>2</sub> gas escape from solution in a bottle when the cap is undone.

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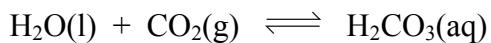
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Check your answers.

# Carbon dioxide, water and carbonic acid equilibrium

When carbon dioxide dissolves in water it forms carbonic acid solution.



The more carbon dioxide dissolves in water then the more acidic the solution becomes as more and more carbonic acid is formed. Another way of estimating the amount carbon dioxide dissolved in water is to measure the pH of the water.

Freshly distilled high purity water has a pH of 7. If this water is exposed to the air  $\text{CO}_2$  gas from the air dissolves to form a very dilute carbonic acid solution with a pH between 5.5 and 6.0. Thus pure distilled water you use in a laboratory does not normally have a pH of 7!



On a summer's day distilled water in a lab has a pH of 6.0 while on a winter's day in the same lab it is found to have a pH of 5.5.

Account for this difference.

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Check your answer.

Carbon dioxide is about 0.03% of the volume of air. This means that 0.03% of the molecules in air are  $\text{CO}_2$  and the pressure of  $\text{CO}_2$  is 0.03% of atmospheric pressure.

Pressure of  $\text{CO}_2$  gas in air is about  $(0.03/100) \times 100 \text{ kPa} = 0.03 \text{ kPa}$ .

If water is covered with pure  $\text{CO}_2$  at an atmospheric pressure of 100 kPa the carbonic acid concentration will be  $100/0.03 = 3\,000$  times higher. As a result the pH will be lower – typically pH 3 to 4 for soda water.

Carbonic acid lowers the pH because of the following reaction that produces hydrogen ions:





Describe what you hear and see when you remove the lid or cap of a bottle of carbonated drink (soda water, soft drink, beer, and so on):

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Do you think the carbon dioxide used to prepare the drink was at less than, about equal to or greater than 100 kPa? Justify your answer.

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Check your answers.

If bottles of carbonated drink are placed in a freezer they could burst for two reasons:

- Liquid water expands on freezing to ice. If the bottle is nearly full of liquid the volume of expanding ice could be greater than the internal volume of the bottle and the bottle bursts.
- As the liquid freezes there is less liquid to dissolve the CO<sub>2</sub>. A lot of CO<sub>2</sub> gas is released from the freezing liquid raising the gas pressure in the bottle. This can burst the bottle.

Try to avoid this happening. Finding all the scattered glass or plastic bottle fragments produced is not much fun!



What happens to a carbonated drink if its container is opened and it is left to stand for hours before it is drunk?

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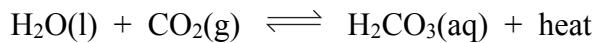
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Check your answer.

## Applying Le Chatelier's principle

The carbon dioxide, water and carbonic acid equilibrium can be written as:



If the forward reaction from left to right is exothermic what word would you use to describe the reverse reaction? \_\_\_\_\_

Check your answer.

Le Chatelier's principle can be stated as:

'equilibrium shifts to oppose external change'.

Consider a sealed container of soda water. This is a closed system containing carbonic acid solution in contact with carbon dioxide gas.

If the container is heated bubbles of CO<sub>2</sub> gas are released. Le Chatelier's principle predicts that as heat is added to the equilibrium  
 $\text{H}_2\text{O(l)} + \text{CO}_2(\text{g}) \rightleftharpoons \text{H}_2\text{CO}_3(\text{aq}) + \text{HEAT}$  it will shift to the left.



Use Le Chatelier's principle to predict what happens when the cap of a soda water container is released so that the CO<sub>2</sub> gas pressure above the solution is reduced.

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Explain how your prediction matches an observation you have made when opening a container of carbonated drink.

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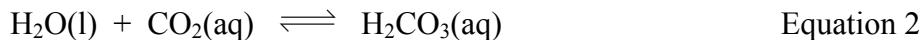
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Check your answers.

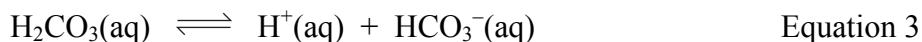
$\text{H}_2\text{O(l)} + \text{CO}_2(\text{g}) \rightleftharpoons \text{H}_2\text{CO}_3(\text{aq})$  is a simplification of the carbon dioxide, water and carbonic acid system. The system can be regarded as made up of three equilibria:



the solution of carbon dioxide gas in water

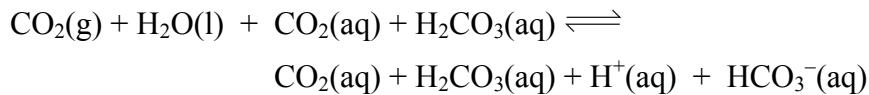


the reaction of dissolved carbon dioxide with water

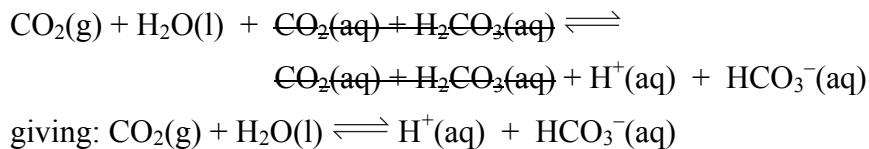


carbonic acid behaves as an acid by releasing a hydrogen ion.

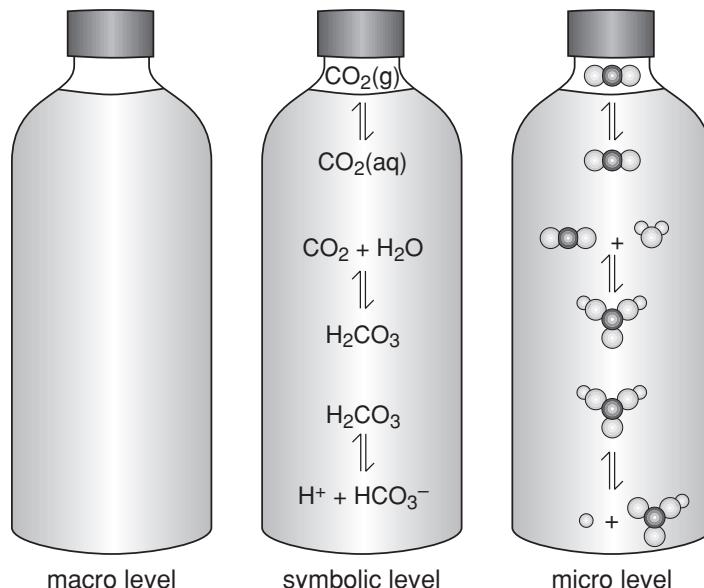
If these three equations are added together:



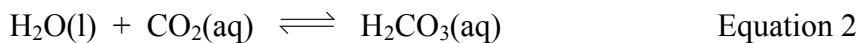
then simplified:



it is easier to see why increasing the pressure of  $\text{CO}_2$  gas increases the concentration of  $\text{H}^+$  and hence lowers the pH.



Three ways of seeing equilibrium in a closed system.



Consider the following changes:

### 1 An increase in CO<sub>2</sub> gas pressure

Le Chatelier's principle predicts that in Equation 1 an increase in CO<sub>2</sub> gas pressure would increase the concentration of dissolved CO<sub>2</sub>.

In Equation 2 this shifts the equilibrium to the right. The increased concentration of carbonic acid in Equation 3 results in an equilibrium shift to the right that increases the concentration of H<sup>+</sup> and lowers the pH.

This matches the observation that the more CO<sub>2</sub> is dissolved in water the lower the pH becomes.

### 2 Raising the temperature of the system

Raising the temperature increases the amount of heat energy available. The decomposition of a substance requires heat energy. With more heat energy available carbonic acid decomposes and the Equation 2 equilibrium shifts to the left.

The higher concentration of CO<sub>2</sub>(aq) will cause Equation 1 equilibrium to shift to the left also and more CO<sub>2</sub> gas is released.

This matches the observation that CO<sub>2</sub> gas is released as the temperature of a carbon dioxide solution is raised.



Fill in the underlined parts in the example below

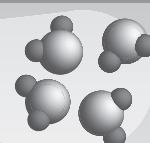
### 3 Adding alkali to the solution increases the solubility of CO<sub>2</sub>.

Hydroxide ion from the alkali will react with the H<sup>+</sup> shown in Equation \_\_\_\_.

The lowered concentration of H<sup>+</sup> causes the equilibrium to shift to the right. This lowers the concentration of carbonic acid and causes the equilibrium in Equation 2 to shift to the \_\_\_\_\_.

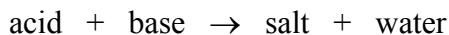
This lowers the \_\_\_\_\_ of dissolved CO<sub>2</sub> resulting in the equilibrium in Equation 1 shifting to the \_\_\_\_\_.

Check your answers.

**MACRO**observe  
infer  
understand**SYMBOLIC** $H_2O$  formulas  
equations  
calculations**MICRO**particles  
energy  
interactions

# Neutralisation

**Neutralisation** is the reaction between an acid and base to form a salt and water. The properties of the acid and properties of the base are lost and replaced by the properties of a salt in water.



Observable acid properties are:

- turn moist blue litmus red
- conduct electricity in solution
- sour taste.

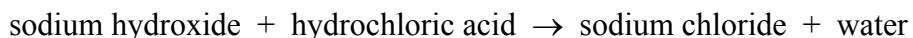
Observable base properties are:

- turn moist red litmus blue
- conduct electricity in solution
- bitter taste.

Over one hundred years ago chemists visualised acid solutions as containing hydrogen ions  $H^+$  and base solutions as containing hydroxide ions  $OH^-$ .

Thus hydrochloric acid solution  $HCl(aq)$  contained  $H^+$  and  $Cl^-$  while sodium hydroxide solution  $NaOH(aq)$  contained  $Na^+$  and  $OH^-$ .

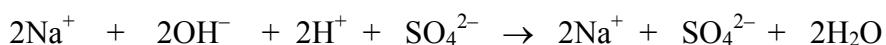
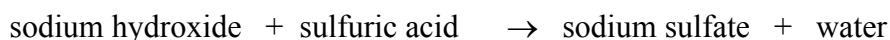
If an acid solution and a base solution were mixed the  $H^+$  and  $OH^-$  ions reacted producing water. The anion from the acid and the cation from the base formed the salt.



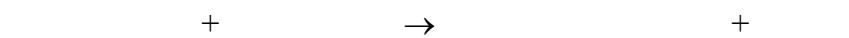
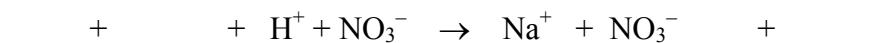
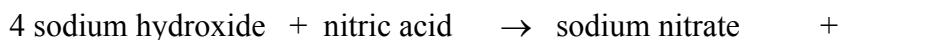
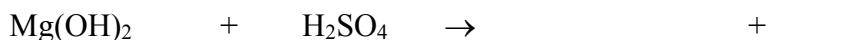
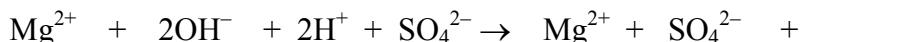
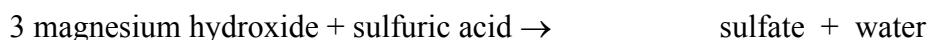
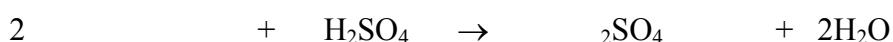
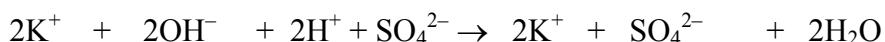
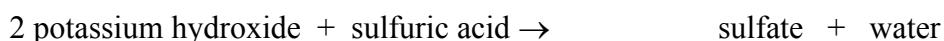
If the acid solution and base solution contained equal amounts of acid and base the salt solution formed did not have acid or base properties. The salt solution was called neutral.

Neutralisation is not an equilibrium situation. The reaction is said to go to completion. All the acid and base present react to produce products. Le Chatelier's principle is applied to systems where there are significant levels of reactant and product – this does not apply to neutralisation.

If NaOH is reacted with sulfuric acid H<sub>2</sub>SO<sub>4</sub> which can form two H<sup>+</sup> per molecule then two moles of NaOH are needed for each mole of H<sub>2</sub>SO<sub>4</sub>.



Complete the word, ionic and neutral formula balanced equations for neutralisations by filling in the parts marked \_\_\_\_\_.

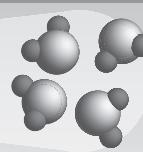


Check your answers.

In all these neutralisation reactions the only ions that react are the hydrogen ions and hydroxide ions that react to form water.



All other ions in the reactions are 'spectator' ions. They float around but don't actually react. The negative ions from the acid and the positive ions from the base are still floating around after reaction. If the salt solution formed is evaporated the salt ions will come together to form solid salt.

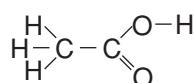
**MACRO**observe  
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understand**SYMBOLIC** $H_2O$  formulas  
equations  
calculations**MICRO**particles  
energy  
interactions

# Safe use of acids and bases

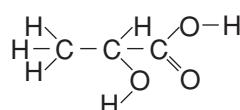
## Acids

Common laboratory acids are the inorganic acids hydrochloric acid ( $HCl$ ), sulfuric acid ( $H_2SO_4$ ) and nitric acid ( $HNO_3$ ). Common household acids are organic acids such as:

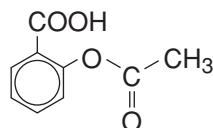
- acetic (ethanoic) acid in vinegar  $CH_3COOH$



- lactic (2-hydroxypropanoic) acid in yoghurt and recently exercised muscles  $CH_3CHOHCOOH$



- acetylsalicylic acid (aspirin) in aspirin tablets  $CH_3COOC_6H_4COOH$ .



Note that organic acids contain the  $-COOH$  group.

Whenever an acid is diluted with water the acid must be poured into water and not vice versa. This can be remembered using the mnemonic ARROW



When you read ArroW from left to right you go from Acid to Water.

**ALWAYS ADD ACID TO WATER!**

Stirring should be used while the acid is being added to water. This way the mixture is much less likely to boil and splash out acid.

Acids are corrosive, skin irritants and often release lung-irritant vapours.

Chemical splashes on the skin should be washed for 5 minutes or until all traces of the chemical have disappeared. If acid splashes in a person's eye they must immediately wash their eye with water for 15 minutes. The water flow should be slow and steady and eyelids should be held back. If the person is wearing a contact lens on the eye the contact lens should be removed shortly after washing starts to improve contact between water and eye.

## Bases

Common laboratory bases are sodium hydroxide NaOH, potassium hydroxide KOH and ammonia solution NH<sub>3</sub> (aq). These water soluble bases are called alkalis and form alkaline solutions.

Alkaline solutions are amongst the most dangerous substances used in school laboratories. When alkalis dissolve in water they generate a lot of heat and can cause the solution to boil or splash.

These strong alkalis are extremely corrosive. Eye and skin protection should always be worn when using them. They are more dangerous to the eyes than acid (acid **denatures** water soluble proteins in the liquid on an eyeball bringing them out of solution and forming a slightly protective film; alkali does not).

Alkali spilt on skin or eye must immediately be washed off with large amounts of water for many minutes. Be very careful that parts of your body that have been near alkali stay away from the eyes.

## Gas solutions

Hydrochloric acid is a solution of hydrogen chloride gas in water while ammonia solution is a solution of ammonia gas in water.

Both these solutions given off harmful fumes and containers should be well sealed. If hydrogen chloride fumes from hydrochloric acid enter your lungs they will dissolve forming an acid solution inside your lungs. Similarly if ammonia fumes from ammonia solution enter your lungs they will dissolve forming a basic solution inside your lungs. Your lungs operate best at pH 7.4, not in acid or base!

Approximately one third of the weight of both a concentrated hydrochloric acid solution and a concentrated ammonia solution is due to the dissolved gas.

If the lid is left off a container of hydrochloric acid or ammonia solution the escaping gas will lower the concentration of the solutions.

## Material Safety Data Sheets MSDS

A **hazardous** substance is one which has the potential to harm the health of people. By law manufacturers and suppliers of a hazardous substance are required to provide a material safety data sheet (**MSDS**). The MSDS of a chemical contains risk and safety information about the properties and uses of the substance.

A MSDS for the alkali sodium hydroxide is shown in the *Appendix*.



Exercise 2.1 requires you to read the MSDS in the *Appendix* and answer questions using the information supplied in the MSDS. Now could be a suitable time for you to do this exercise.

When you do not have access to a MSDS:

READ THE LABEL

and

HEED THE LABEL

for any chemical that you are going to use.

It is very important that you are mentally prepared with knowledge and understanding of a hazardous substance before you physically use it.



MSDSs can be accessed at a number of web sites.

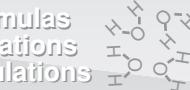
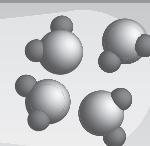
Try to find a web site supplying MSDSs for:

- sulfur dioxide SO<sub>2</sub>
- nitrogen monoxide NO
- nitrogen dioxide NO<sub>2</sub>

by using a search engine such as [www.google.com](http://www.google.com) or [www.northernlight.com](http://www.northernlight.com).

Record the address of any suitable web site.

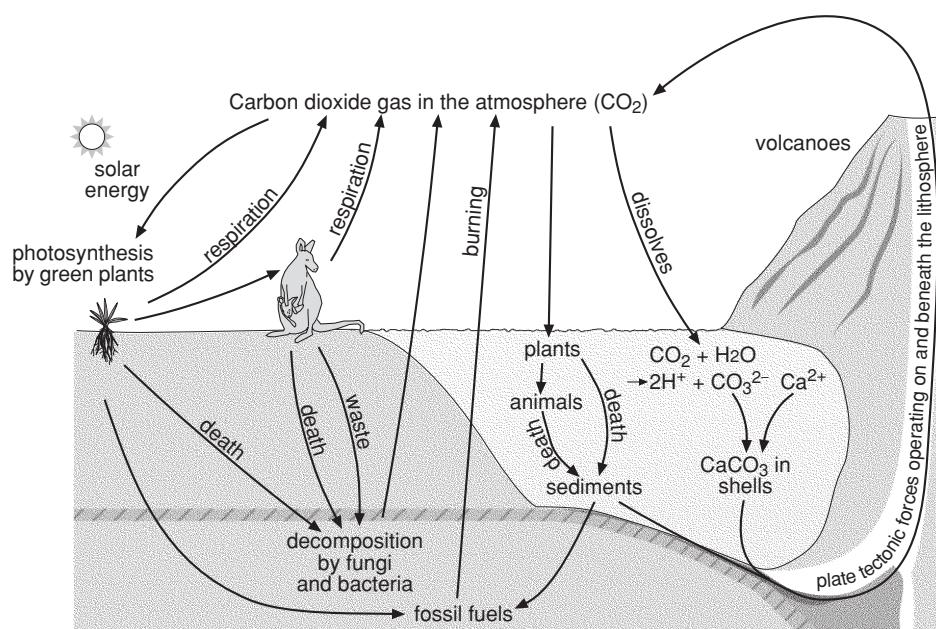
If you cannot find a suitable source of MSDSs go to [www.lpmc.edu.au/science](http://www.lpmc.edu.au/science) and look at links provided for this module.

**MACRO**observe  
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interactions

# Acidic oxides in the atmosphere

## Carbon dioxide $CO_2$

Of the over 100 billion tonnes of carbon released into the earth's atmosphere as  $CO_2$  each year only about 6 billion tonnes of carbon comes from the burning of fossil fuel. Some of this  $CO_2$  released from fossil fuel leads to more photosynthesis. However the rest is added to the  $CO_2$  level of the atmosphere which is increasing about 0.4% per year.



Using the diagram above explain how the carbon dioxide level in the atmosphere is partly controlled by the oceans and living things with shells.

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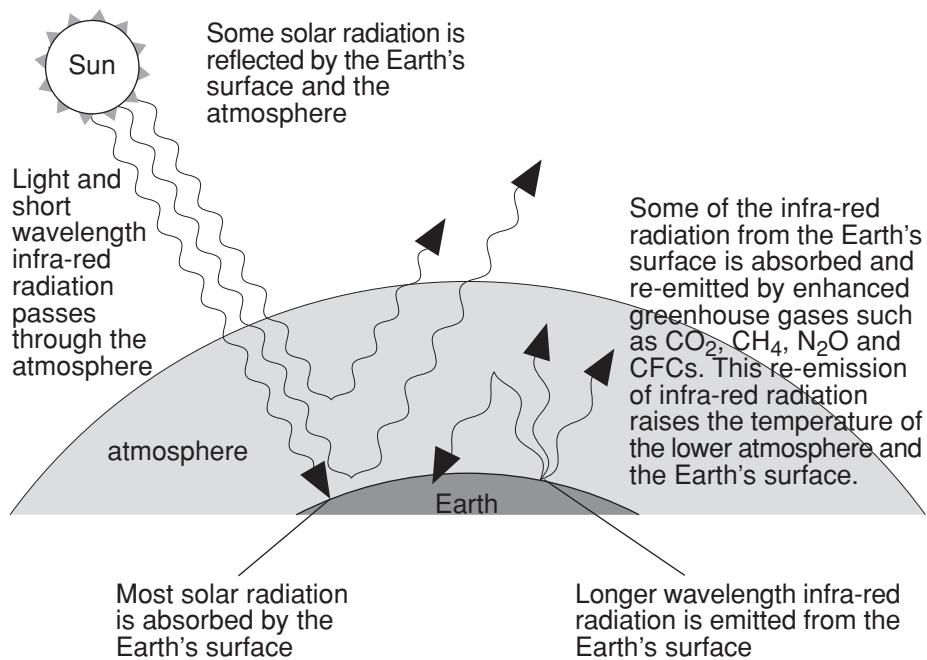


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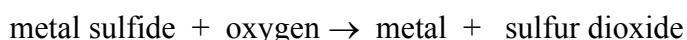
The increased CO<sub>2</sub> levels are believed by many scientists to contribute to an enhanced greenhouse effect that is raising global temperatures. International agreements to control CO<sub>2</sub> release by the nations of the world are actively being pursued through the United Nations.



## Sulfur dioxide SO<sub>2</sub>

Of the 100 million tonnes of sulfur released to the atmosphere each year only about one quarter comes from natural sources. Over three-quarters of the sulfur comes from human sources! Most of the sulfur emitted is in the form of SO<sub>2</sub> gas and 80% of that is emitted in the Northern Hemisphere.

Approximately two-thirds of the sulfur comes from power plants burning coal or oil and most of the rest from smelters and refineries processing sulfide ores.

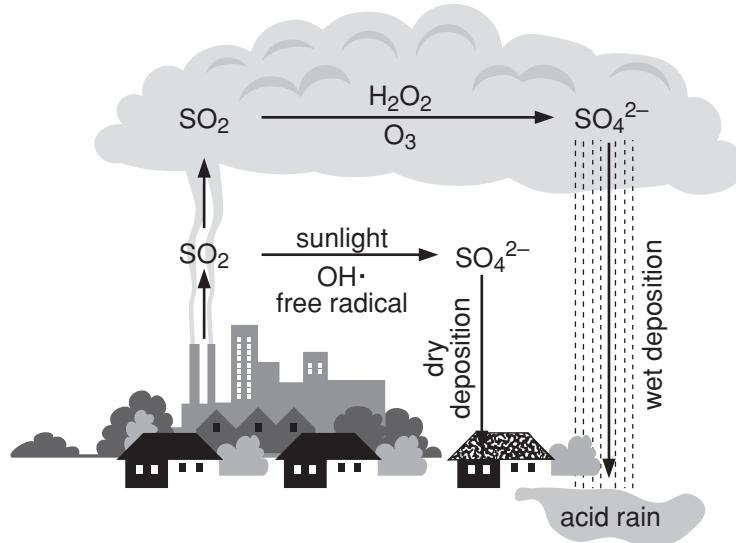
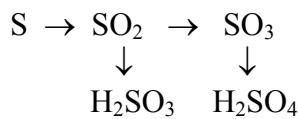


One of the reasons why Australia is the largest exporter of coal in the world is the low sulfur content of Australian coals.

The sulfur dioxide can react with water to form sulfuric acid:



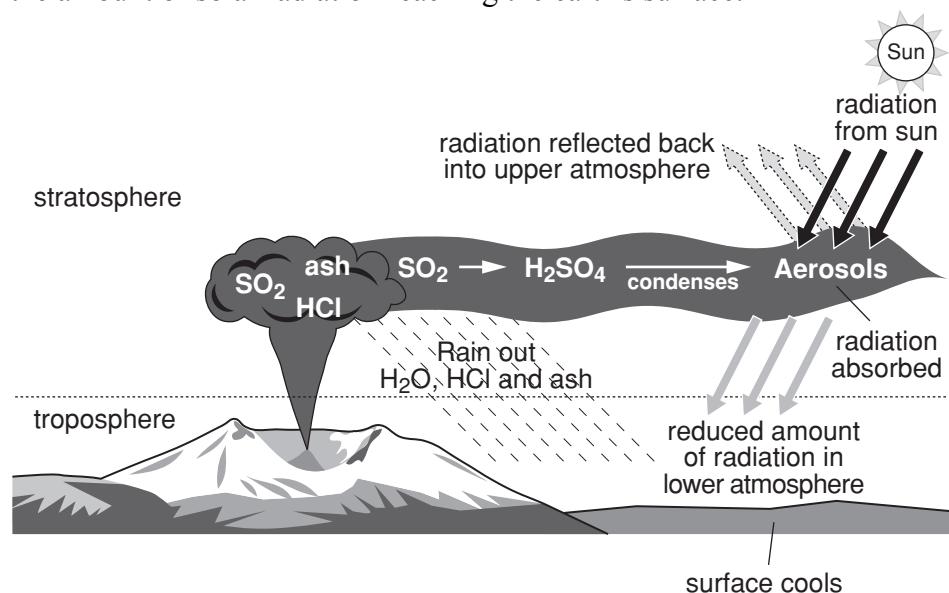
The sulfur dioxide and sulfurous acid can be oxidised to form sulfur trioxide  $\text{SO}_3$  and sulfuric acid  $\text{H}_2\text{SO}_4$ .



The sulfur is mostly returned to the earth's surface as

- wet deposition – sulfuric acid in acid rain
- dry deposition – sulfate particles.

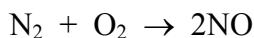
An unpredictable source of sulfur oxides are major volcanic eruptions. The  $\text{SO}_2$  released can produce large amounts of aerosols – very small solid or liquid particles suspended in air. Aerosol production can reduce the amount of solar radiation reaching the earth's surface.



The 1991 eruption of Mt. Pinatubo in the Philippines saw a reduction in global temperatures of 0.5°C. In the Northern Hemisphere the reduction in temperature was about 0.7°C.

## Nitrogen oxides NO<sub>x</sub>

Lightning and high temperatures over 1300°C reached in some combustions can cause reaction between the two most common gases in the atmosphere:



The colourless NO is called nitrogen monoxide, nitric oxide or nitrogen (II) oxide. NO reacts with oxygen in the air to form NO<sub>2</sub> – nitrogen dioxide or nitrogen (IV) oxide.



The nitrogen dioxide reacts with water to form a mixture of nitrous acid HNO<sub>2</sub> and nitric acid HNO<sub>3</sub>.



The NO and NO<sub>2</sub> in the atmosphere are called nitrogen oxides, NO<sub>x</sub>, NOx or NOX.

## Formation of acid rain

Pure rainwater is slightly acidic largely because of dissolved CO<sub>2</sub>. In Australia the pH of unpolluted rainwater ranges from about 6 to about 5. A pH as low as 2 in rainwater has been recorded in some parts of the Northern Hemisphere.

The total mass of NO<sub>x</sub> released by human activity into the atmosphere is about the same as for SO<sub>2</sub>. Acid produced from NO<sub>x</sub> as well as SO<sub>2</sub> contributes to acid rain. Most of this acid is precipitated, that is falls as rain, snow or hail, within hundreds of kilometres of its source.

Sources of NO<sub>x</sub> are more widespread around the Earth whereas sources of sulfur dioxide are fewer in number but larger. In Australia, lightning strikes and high temperature combustion reactions occur all over the continent. In contrast most sulfur dioxide is released at smelters in Mt Isa and Kalgoorlie mining regions or the coal burning power stations in the Hunter Valley (NSW) and Latrobe Valley (Victoria).

In the city of Melbourne 10% of the SO<sub>2</sub> and 60% of the NO<sub>x</sub> come from motor vehicle exhausts. A century ago most heating of Sydney homes was by burning coal and coal burning power stations operated in the city. The SO<sub>2</sub> released and formation of sulfates has caused considerable damage to old Sydney buildings made of sandstone.

## Effects of acid rain

If the quantity of acid rain deposited is greater than the capacity of the environment to neutralise it:

- soil chemistry can be changed making it difficult for plant roots to take up essential elements such as calcium and potassium
- micro-organisms in soil essential for recycling of nutrients are killed
- soil chemistry can be changed releasing high levels of normally insoluble elements such as aluminium and mercury in runoff
- crop production can be reduced as plants are sensitive to soil pH
- leaf damage can occur, most noticeably in trees, as protective waxes are removed from leaf surfaces
- buildings (especially limestone, marble or concrete) can be damaged
- smog and acid rain can combine to be a human health hazard
- most fish and shellfish die at pH 4.5 to 5.0
- aquatic life in lakes can be killed so the lake is crystal clear, beautiful looking but dead.

- 1 Summarise the industrial origins of sulfur dioxide and nitrogen oxides.



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- 2 Evaluate (make a judgement based on criteria) reasons for concern about release of sulfur dioxide and nitrogen oxides into the environment.

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Check your answers.

## Measuring concentrations of carbon dioxide, sulfur dioxide and nitrogen oxides

Have you noticed how air bubbles are trapped in water when it freezes? When a core of ice is extracted from the ice covering Antarctica it can contain air bubbles that were trapped many thousands of years ago. These bubbles are thought to represent samples of past atmospheres. Measurements of the proportion of carbon dioxide in these bubbles show that the concentration in the Earth's atmosphere remained constant at 280 parts per million (ppm) for thousands of years but has steadily increased to 360 ppm over the last 200 years.

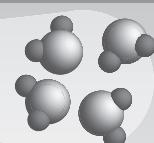
Similarly measurements of carbon isotope content of old trees, grass seeds collected over hundreds of years and the calcium carbonate in coral point to a steady increase in atmospheric carbon dioxide since the industrial revolution.

It is much more difficult to find evidence for changes in atmospheric sulfur dioxide and nitrogen oxides for the following reasons:

- SO<sub>2</sub> and NO<sub>x</sub> are found at much lower levels in the atmosphere – typically about 0.001 ppm in populated parts of the Earth but as low as 0.00005 ppm well away from human activity, electrical storms and volcanic action
- CO<sub>2</sub> forms carbonate salts which are water insoluble whereas most sulfate salts and all nitrate salts are water soluble. Carbonate salts remove carbon dioxide from the atmosphere and place it in a relatively inert form such as an animal shell that can act as a storehouse of chemical information particularly if measurements are made of isotope ratios using mass spectrometers. Soluble sulfate and nitrate salts circulate in the hydrosphere and biosphere and are chemically changed
- Chemical instrumentation able to measure very small amounts of sulfur and nitrogen have only become commercially available and readily accessible to scientists in the last twenty or thirty years.



Do Exercise 2.2 now.

**MACRO**observe  
infer  
understand**SYMBOLIC** $H_2O$  formulas  
equations  
calculations**MICRO**particles  
energy  
interactions

# Using the mole concept

The mole concept can be used in calculations involving some chemicals measured in mass and other chemicals measured as gas volumes.

Suppose that you wanted to calculate the volume of carbon dioxide gas produced by heating 1 000 g of calcium carbonate.

- 1 Make sure any formulas you write are correct.



- 2 Check that the equation is balanced



- 3 Underline the formula of any chemical with a known amount and write the amount underneath



1 000 g

- 4 Underline the formula of any chemical of unknown amount you are asked to find. Place '?' under each formula of unknown amount.



1 000 g ?

- 5 Use the mole ratio in the equation to construct a proportionality involving the known amount and the unknown amount

$$\frac{\text{molCaCO}_3}{\text{molCO}_2} = \frac{1}{1}$$

- 6 Rearrange the proportionality so that the unknown amount is on one side and all the known amounts are on the other side

$$\text{mol CO}_2 = \text{mol CaCO}_3$$

- 7 Change known quantity to moles.

$$\text{Formula mass of CaCO}_3 = 40 + 12 + (3 \times 16) = 100$$

$$1 000 \text{ g of CaCO}_3 = 1 000 / 100 = 10.0 \text{ mol}$$

8 Calculate mol of unknown quantities.

$$\text{mol CO}_2 = \text{mol CaCO}_3 = 10.0 \text{ mol}$$

9 Convert mol to grams or litres

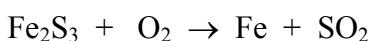
$$\text{At } 25^\circ\text{C and 101.3 kPa } 10.0 \text{ mol gas} = 10.0 \times 24.5 \text{ L} = 245 \text{ L.}$$



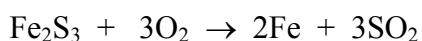
One thousand grams of an iron(III) sulfide ore sample was heated and released 98 L of SO<sub>2</sub> gas at 25°C and 101.3 kPa. Calculate the mass of iron(III) sulfide in the sample.

Complete the steps for calculating the answer.

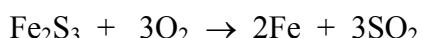
1 Make sure any formulas you write are correct.



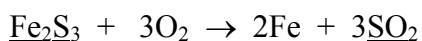
2 Check that the equation is balanced



3 Underline the formula of any chemical with a known amount and write the amount underneath.



4 Underline the formula of any chemical of unknown amount you are asked to find. Place '?' under each formula of unknown amount.



5 Use the mole ratio in the equation to construct a proportionality involving the known amount and the unknown amount

$$\frac{\text{mol Fe}_2\text{S}_3}{\text{mol SO}_2} = -$$

6 Rearrange the proportionality so that the unknown amount is on one side and all the known amounts are on the other side

$$\text{mol Fe}_2\text{S}_3 = - \times \text{mol SO}_2$$

7 Change known quantity to moles.

$$\text{At } 25^\circ\text{C and 101.3 kPa } 98 \text{ L of gas} = 98/24.5 \text{ mol} = 4.0 \text{ mol}$$

8 Calculate mol of unknown quantities.

$$\text{mol Fe}_2\text{S}_3 = \frac{1}{3} \times \text{mol SO}_2 = \frac{1}{3} \times 4.0 = 4/3 \text{ mol}$$

9 Convert mol to grams or litres

$$4/3 \text{ mol Fe}_2\text{S}_3 = (4/3) \times ( ) \text{ g} = \text{_____ g}$$

Check your answers.

**MACRO**

observe  
infer  
understand



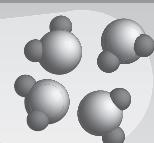
**SYMBOLIC**

$H_2O$  formulas  
equations  
calculations



**MICRO**

particles  
energy  
interactions



## Appendix: MSDS for sodium hydroxide

# sodium hydroxide

ChemWatch Material Safety Data Sheet (REVIEW)

CHEMWATCH 1823

Date of Issue: Thu 27-May-1999

## Identification

### Statement of hazardous nature

Hazardous according to WorkSafe Australia criteria.

## Supplier

Company: Tennant Limited

Address:

210 George St	492 St Kilda Rd	3/186 Hampden Rd	215 Adelaide St
Sydney	Melbourne	Nedlands	Brisbane
NSW 2000	VIC 3004	WA 6000	QLD 4000
Australia	Australia	Australia	Australia

Telephone: (02) 9321 0300/(03) 9820 1400

Telephone: (089) 389 1777/(07) 3229 2711

Emergency Tel: 1800 039 008

Fax: (02)9321 0333/(03)9820 1663

## **Chemwatch hazard ratings**

Flammability: 0

Toxicity: 2

Body Contact: 4

Reactivity: 2

Chronic: 0

Scale: Min/Nil=0 Low=1 Moderate=2 High=3 Extreme=4

## **Personal protective equipment for industrial/commercial environments**

Long Gloves

Overalls

Boots

Face Shield or Full Face Respirator

Product Name: sodium hydroxide

Other Names: sodium hydroxide

sodium hydroxide, flake, technical

CAS RN No(s): 1310-73-2

UN Number: 1823

Packaging Group: II

Dangerous Goods Class: 8

Subsidiary Risk: None

Hazchem Code: 2X

Poisons Schedule Number: S6 NZS3

## **Use**

Component of alkali cleaners. Manufacture of soap, pulp and paper; rayon.

Chemical manufacture. Neutralising agent in petroleum refining; manufacture of aluminium, detergents, textile processing, refining of vegetable oils. Laboratory reagent, for organic fusion, etching of metal.

Used for regenerating ion exchange resins, lye peeling of fruits and vegetables in the food industry.

## **Physical description/properties**

### **Appearance**

White hygroscopic, odourless, pellets, flakes, sticks or solid cast mass.

Vigorously exotherms when mixed with water. Explosive boiling and spitting will occur if added to hot water. Reacts violently with acids.

CAUSTIC alkali. Soluble in water, alcohol, ether, glycerol.

In the presence of moisture, highly corrosive to aluminium, zinc and tin.

HIGHLY reactive: with ammonium salts evolves ammonia gas. Rapidly picks up moisture from the air and with carbon dioxide in air forms sodium carbonate.

Boiling Point (deg C): 1390

Melting Point (deg C): 318.4

Vapour Pressure (kPa): Negligible.

Specific Gravity: 2.12 @ 20 C

Flash Point (deg C): Not applicable

Lower Explosive Limit (%): Not applicable

Upper Explosive Limit (%): Not applicable

Solubility in Water (g/L): Soluble.

## **Health hazard**

### **Acute health effects**

#### **Swallowed**

Considered an unlikely route of entry in commercial/industrial environments.

The solid is extremely corrosive to the gastro-intestinal tract and may be fatal if swallowed.

Ingestion may result in severe burns to the mouth, throat and stomach, pain, nausea and vomiting, swelling of the larynx and subsequent suffocation, perforation of the gastro-intestinal tract.

A 1% aqueous solution (pH 13.4) failed to cause gastric, oesophageal or other damage in rabbits.

## **Eye**

The solid/dust is extremely corrosive to the eyes and is capable of causing severe damage with loss of sight.

The material may produce severe irritation to the eye causing pronounced inflammation. Repeated or prolonged exposure to irritants may produce conjunctivitis.

## **Skin**

The solid/dust is highly discomforting and extremely corrosive to the skin and is capable of causing severe burns and ulceration.

Bare unprotected skin should not be exposed to this material.

The material may produce severe skin irritation after prolonged or repeated exposure, and may produce a contact dermatitis (nonallergic).

This form of dermatitis is often characterised by skin redness (erythema) and swelling (oedema) which may progress to vesication, scaling and thickening of the epidermis.

Histologically there may be intercellular oedema of the spongy layer (spongiosis) and intracellular oedema of the epidermis.

Prolonged contact is unlikely, given the severity of response, but repeated exposures may produce severe ulceration.

Burns are not immediately painful; onset of pain may be delayed minutes or hours; thus care should be taken to avoid contamination of gloves and boots.

A 5% aqueous solution applied to the skin of rabbits for 4 hours produced severe necrosis. Instillation of a 1% solution into the conjunctival sac failed to produce ocular or conjunctival injury in rabbits provided the eye was promptly irrigated with copious amounts of water.

## **Inhaled**

Generated dust may be highly discomforting and corrosive to the upper respiratory tract if inhaled and is capable of causing severe burns to the upper respiratory tract.

The material may produce respiratory tract irritation which produces an inflammatory response involving the recruitment and activation of many cell types, mainly derived from the vascular system. Unlike most organs the lung can respond to a chemical insult or agent by first trying to remove or neutralise the irritant and then repairing the damage. The repair process, which initially developed to protect mammalian lungs from foreign matter and antigens, may however, cause further damage the lungs when activated by hazardous chemicals. The result is often the impairment of gas exchange, the primary function of the lungs.

Severe acute dust inhalation exposure may be fatal due to spasm, inflammation and oedema of the larynx and bronchi, chemical pneumonitis and severe pulmonary oedema.

Symptoms of overexposure include burning sensation, coughing, wheezing, laryngitis, shortness of breath, headache, nausea and vomiting.

## **Chronic health effects**

The material is considered to be harmful by all exposure routes.

Principal routes of exposure are usually by skin contact with the material, eye contact with the material and inhalation of generated dust.

As with any chemical product, contact with unprotected bare skin; inhalation of vapour, mist or dust in work place atmosphere; or ingestion in any form, should be avoided by observing good occupational work practice.

## **First aid**

### **Swallowed**

DO NOT delay. .

If poisoning occurs, contact a doctor or Poisons Information Centre.

If swallowed, do NOT induce vomiting. Give a glass of water.

### **Eye**

DO NOT delay. If this product comes in contact with the eyes:

- 1: Immediately hold the eyes open and wash continuously for at least 15 minutes with fresh running water.
- 2: Ensure complete irrigation of the eye by keeping eyelids apart and away from eye and moving the eyelids by occasionally lifting the upper and lower lids.
- 3: Transport to hospital or doctor without delay.
- 4: Removal of contact lenses after an eye injury should only be undertaken by skilled personnel.

### **Skin**

DO NOT delay. If this product comes in contact with the skin:

- 1: Immediately flush body and clothes with large amounts of water, using safety shower if available.
- 2: Quickly remove all contaminated clothing, including footwear.

- 3: Wash affected areas with water (and soap if available) for at least 15 minutes.
  - 4: Transport to hospital, or doctor.
- or In case of burns:
- 1: Quickly immerse affected area in cold running water for 10 to 15 minutes.
  - 2: Bandage lightly with a sterile dressing. Treat for shock if required.
  - 3: Lay patient down. Keep warm and rested.
  - 4: Transport to hospital, or doctor.

### **Inhaled**

- 1: If dust is inhaled, remove to fresh air.
- 2: Encourage patient to blow nose to ensure clear breathing passages.
- 3: Ask patient to rinse mouth with water but to not drink water.
- 4: Seek immediate medical attention.

### **Advice to doctor**

For acute or short-term repeated exposures to highly alkaline materials:

1. Respiratory stress is uncommon but presents occasionally because of soft tissue edema.
2. Unless endotracheal intubation can be accomplished under direct vision, cricothyroidotomy or tracheotomy may be necessary.
3. Oxygen is given as indicated.
4. The presence of shock suggests perforation and mandates an intravenous line and fluid administration.
5. Alkali corrosives damage occurs by liquefaction necrosis whereby the saponification of fats and solubilisation of proteins allow deep penetration into the tissue. Alkalies continue to cause damage after exposure.

### **Ingestion:**

1. Milk and water are the preferred diluents. No more than 2 glasses of water should be given to an adult.
2. Neutralising agents should never be given since exothermic heat reaction may compound injury.
  - \* Catharsis and emesis are absolutely contra-indicated.
  - \* Activated charcoal does not absorb alkali.
  - \* Gastric lavage should not be used.

Supportive care involves the following.

1. Withhold oral feedings initially.
2. If endoscopy confirms transmucosal injury start steroids only within the first 48 hours.
3. Carefully evaluate the amount of tissue necrosis before assessing the need for surgical intervention.
4. Patients should be instructed to seek medical attention whenever they develop difficulty in swallowing (dysphagia).

#### **Skin and eye**

1. Injury should be irrigated for 20-30 minutes.

Eye injuries require saline. [Ellenhorn & Barceloux: Medical Toxicology].

## **Precautions for use**

### **Exposure standards**

TLV C: 2 mg/m<sup>3</sup>

ES Peak: 2 mg/m<sup>3</sup>

OES STEL: 2 mg/m<sup>3</sup>

IDLH Level: 250 mg/m<sup>3</sup>

The TLV-C is recommended based on concentrations that produce noticeable but not excessive, ocular and upper respiratory tract irritation.

### **Engineering controls**

Use in a well-ventilated area

DO NOT handle directly. Wear gloves and use scoop / tongs / tools.

If risk of overexposure exists, wear SAA approved respirator.

If conditions where worker exposure potential is high, wear full-face air-supplied breathing apparatus and full protective suit.

### **Personal protection**

#### **Eye**

Safety glasses with side shields Chemical goggles. Full face shield.

Contact lenses pose a special hazard; soft lenses may absorb irritants and all lenses concentrate them.

### **Hands/feet**

DO NOT handle directly. Wear gloves and use scoop / tongs / tools  
Elbow length PVC gloves or Butyl rubber gloves or Neoprene rubber gloves.  
Safety footwear.

### **Other**

- 1: Overalls. 2: PVC Apron.
- 3: PVC protective suit may be required if exposure severe.
- 4: Eyewash unit.
- 5: Ensure there is ready access to a safety shower.

## **Safe handling**

### **Storage and transport**

#### **Suitable container**

Plastic bag or Packaging as recommended by manufacturer. Glass container.  
Polyethylene or polypropylene container or Polylined drum.  
DO NOT use aluminium, galvanised or tin-plated containers.  
Check that containers are clearly labelled.

#### **Storage incompatibility**

Keep dry. Reacts violently with water.  
Segregate from water, strong oxidisers, strong acids, organic materials, ammonium compounds, nitro compounds and trichlorethylene.

#### **Storage requirement**

Keep dry. Reacts violently with water.

- 1: Store in original containers.
- 2: Keep containers securely sealed.
- 3: Store in a cool, dry, well-ventilated area.
- 4: Store away from incompatible materials and foodstuff containers.
- 5: Protect containers against physical damage and check regularly for leaks.
- 6: Observe manufacturer's storing and handling recommendations.

DO NOT use aluminium, galvanised or tin-plated containers.

## **Transportation**

Class 8 Corrosives shall not be loaded in the same vehicle or packed in the same freight container with:

- Class 1 – Explosives;
- Class 4.3 – Dangerous when wet substances;
- Class 5.1 – Oxidising agents;
- Class 5.2 – Organic peroxides;
- Class 7 – Radioactive substances;
- Class 8 – Acids only;
- Food and food packaging in any quantity.

## **Spills and disposal**

### **Minor spills**

DO NOT touch the spill material. Slippery when spilt. Clean up all spills immediately.  
Control personal contact by using protective equipment.  
Use dry clean up procedures and avoid generating dust.  
Place in suitable containers for disposal.

### **Major spills**

DO NOT touch the spill material . Slippery when spilt.  
Keep dry. Reacts violently with water.  
Clear area of personnel and move upwind.  
Alert Fire Brigade and tell them location and nature of hazard.  
1: Wear full body protective clothing with breathing apparatus.  
2: Prevent, by any means available, spillage from entering drains or water courses.  
Shut off all possible sources of ignition and increase ventilation.  
Stop leak if safe to do so.  
Use dry clean up procedures and avoid generating dust.  
Collect recoverable product into labelled containers for recycling. Collect residues and seal in labelled drums for disposal.  
Wash area down with large quantity of water and prevent runoff into drains.  
If contamination of drains or waterways occurs, advise emergency services.  
After clean up operations, decontaminate and launder all protective clothing and equipment before storing and re-using.

## **Disposal**

- 1: Recycle wherever possible or consult manufacturer for recycling options.
- 2: Consult State Land Waste Management Authority for disposal.
- 3: Treat and neutralise with dilute acid at an effluent treatment plant.
- 4: Recycle containers, otherwise dispose of in an authorised landfill.

## **Waste disposal procedures**

- "Wear eye protection, protective clothing and nitrile rubber gloves to control personal contact from sodium hydroxide."
- Add the compound to a large volume of ice water. Neutralise by adding 5% hydrochloric acid and empty into the drain [Armour 1996].

## **Spillage disposal**

- "Wear eye protection, protective clothing and nitrile rubber gloves to control personal contact from sodium hydroxide. Scoop the contents into a container and add small portions into a large volume of ice water."
- Neutralise with 5% hydrochloric acid and empty into the drain. Wash the area of the spill with water [Armour 1996].

## **Fire/explosion hazard**

- 1: Non combustible.
- 2: Not considered to be a significant fire risk, however containers may burn.

Solid in contact with water or moisture reacts violently, and solutions are highly alkaline and may cause severe skin burns.

## **Contact point**

Australian Poisons Information Service

24 hour service: 13 11 26

Police or fire brigade: 000 (Exchange: 1100)

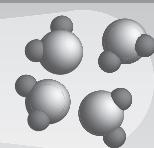
End of Report (review)

Date of Preparation: Thu 27-May-1999

Print Date: Wed 31-Jan-2001

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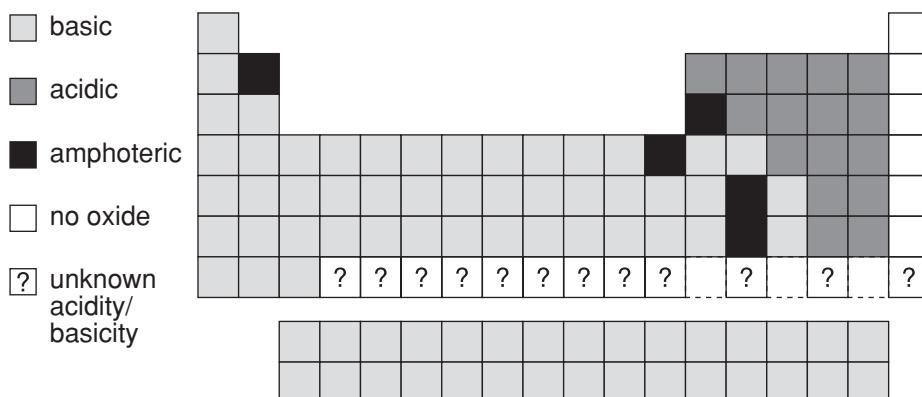
# Suggested answers

## Acidity and basicity of oxides

- 1 The white solid oxide is a basic oxide because it dissolved producing a basic solution from neutral water.
- 2 The colourless oxide gas is acidic because it dissolved making an acidic solution from neutral water.
- 3 The oxide must have been acidic because it neutralised the alkaline solution producing a neutral pH.
- 4 The black oxide must have been basic as some of the acidity of the pH 3.6 solution has been partly neutralised
- 5 The gaseous oxide must have been neutral as it has not produced a change in pH.

Elements combined with oxygen in acidic oxides are non-metals.

Elements combined with oxygen in basic oxides are metals.



- 1 Metal oxides are usually *basic*.
- 2 Non-metal oxides are usually *acidic*.
- 3 Amphoteric oxides are formed by *metallic* elements close to the non-metals in the Periodic Table.

- 4 The only elements that do not form oxides are the *noble gases* in group 18.

## Decarbonating soda water

Each g of CO<sub>2</sub> = 1/44 mole

$$= (1/44) \times 24.5 \text{ L} = 0.56 \text{ L at } 25^\circ\text{C and } 101.3 \text{ kPa}$$

Volume of gas released is a few times greater than the volume of liquid soda water which contained that gas.

These volumes indicate that CO<sub>2</sub> molecules are further apart when in the gas phase than when dissolved in soda water.

When the cap is undone the pressure of CO<sub>2</sub> acting on the water surface decreases. The number of CO<sub>2</sub> entering the water is now much less than the number of CO<sub>2</sub> leaving the water. Equilibrium shifts to oppose the change by CO<sub>2</sub> passing from solution to gas phase.

## Carbon dioxide, water and carbonic acid equilibrium

In winter the water is at a lower temperature and can dissolve more CO<sub>2</sub> from the air. Also during winter more people inside and the combustion of heating fuels can raise CO<sub>2</sub> levels. This produces more carbonic acid, more H<sup>+</sup> and lower pH.

When you remove the lid or cap of a bottle of carbonated water you first hear release of the higher gas pressure then see bubbles of gas released from within the solution.

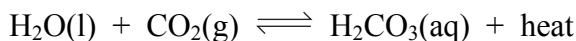
The carbon dioxide would have been at a greater pressure than 100 kPa. The gas pressure was heard to change and CO<sub>2</sub> gas came out of solution to replace the reduced CO<sub>2</sub> gas pressure above the solution.

If a carbonated drink is opened and left to stand for hours a lot of the dissolved gas escapes and the drink tastes flat. Few bubbles are released in the mouth when the flat drink is drunk. Exposed to the air it will eventually reach equilibrium with atmospheric CO<sub>2</sub>(g).

## Applying Le Chatelier's principle

Endothermic. The reverse reaction has to be endothermic if the forward reaction is exothermic.

Le Chatelier's principle predicts that in the equilibrium:



when the  $\text{CO}_2$  gas pressure is reduced the equilibrium shifts to the left.

This prediction matches the observation that when a container of carbonated drink is opened bubbles of  $\text{CO}_2$  gas are released from the solution.

Hydroxide ion from the alkali will react with the  $\text{H}^+$  shown in Equation 3.

The lowered concentration of  $\text{H}^+$  causes the equilibrium to shift to the right. This lowers the concentration of carbonic acid and causes the equilibrium in Equation 2 to shift to the *right*.

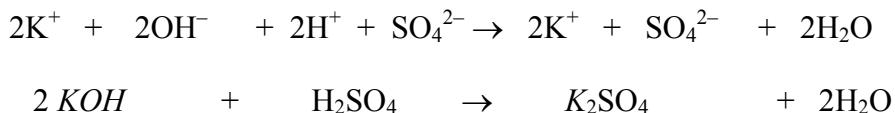
This lowers the *concentration* of dissolved  $\text{CO}_2$  resulting in the equilibrium in Equation 1 shifting to the *right*.

## Neutralisation

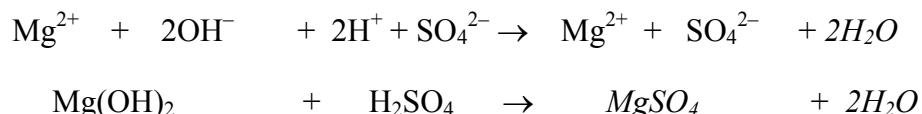
1 potassium hydroxide + *hydrochloric* acid  $\rightarrow$  potassium chloride + water



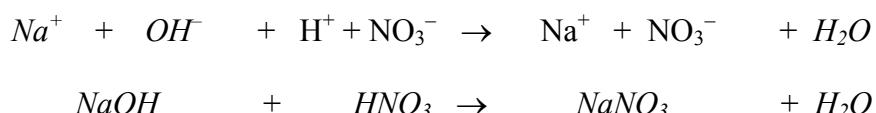
2 potassium hydroxide + sulfuric acid  $\rightarrow$  potassium sulfate + water



3 magnesium hydroxide + sulfuric acid  $\rightarrow$  *magnesium* sulfate + water



4 sodium hydroxide + nitric acid  $\rightarrow$  sodium nitrate + water



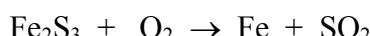
## Carbon dioxide

The larger the volume of water in the oceans and the lower the temperature of this water the more CO<sub>2</sub> can dissolve. Living things with shells require CaCO<sub>3</sub> to make their shells. The more shells are made the lower the concentration of carbonate ions and the more CO<sub>2</sub> is dissolved in water.

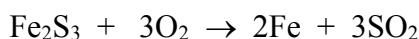
## Effects of acid rain

- 1 SO<sub>2</sub> burning of fossil fuels containing S, smelting of sulfide ores.  
NO electrical discharges (sparking) and high temperature combustion. NO<sub>2</sub> is formed by oxidation of NO by O<sub>2</sub>.
- 2 SO<sub>2</sub> and NO<sub>2</sub> react with water producing acid rain. Acid rain can damage soil, plants, animals, limestone, marble, concrete and some sandstones. This damage harms the natural and human-built environments and detracts from human enjoyment of life.

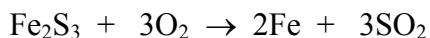
- 1 Make sure any formulas you write are correct.



- 2 Check that the equation is balanced

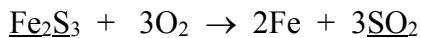


- 3 Underline the formula of any chemical with a known amount and write the amount underneath



98 L

- 4 Underline the formula of any chemical of unknown amount you are asked to find. Place '?' under each formula of unknown amount.



? 98 L

- 5 Use the mole ratio in the equation to construct a proportionality involving the known amount and the unknown amount

$$\frac{\text{mol Fe}_2\text{S}_3}{\text{mol SO}_2} = \frac{1}{3}$$

- 6 Rearrange the proportionality so that the unknown amount is on one side and all the known amounts are on the other side

$$\text{mol Fe}_2\text{S}_3 = \frac{1}{3} \times \text{mol SO}_2$$

- 7 Change known quantity to moles.

At 25°C and 101.3 kPa 98 L of gas = 98/24.5 mol = 4.0 mol

8 Calculate mol of unknown quantities.

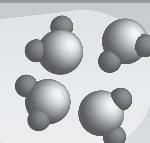
$$\text{mol Fe}_2\text{S}_3 = \frac{1}{3} \times \text{mol SO}_2 = \frac{1}{3} \times 4.0 = 4/3 \text{ mol}$$

9 Convert mol to grams or litres

$$4/3 \text{ mol Fe}_2\text{S}_3 = (4/3) \times (208) \text{ g} = 277 \text{ g}$$

1 000 g of the ore sampled contains 277 g of iron(III) sulfide.



**MACRO**observe  
infer  
understand**SYMBOLIC** $H_2O$  formulas  
equations  
calculations**MICRO**particles  
energy  
interactions

## Exercises – Part 2

Exercises 2.1 to 2.2

Name: \_\_\_\_\_

### Exercise 2.1: Understanding a MSDS

The MSDS for sodium hydroxide is in the *Appendix*.

Read through this MSDS so that you

- become familiar with what is probably the most hazardous substance you will use in this course
- become familiar with the layout of information in a MSDS.

Now answer the following questions, arranged in the order of information in the MSDS:

1 Which of the five Chemwatch hazard ratings is highest for NaOH?

---

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2 Explain why it is difficult to accurately weigh out a fixed amount of NaOH in the air.

---

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3 Outline the first aid procedure if NaOH comes in contact with an eye.

---

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---

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- 4 Outline the first aid procedure if NaOH comes in contact with skin.

---

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- 5 Explain why wearing contact lenses can be a special hazard.

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- 6 Why do you think NaOH cannot be stored in aluminium, galvanised or tin plated containers?

---

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---

---

---

---

- 7 What is the phone number of the Australian Poisons Information Centre? Note: This is only to be used in an emergency.

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---

---

---

---

## Exercise 2.2: Comparing SO<sub>2</sub>, NO and NO<sub>2</sub>

### Properties:

Complete this table of properties by inserting the formulas for sulfur dioxide, nitrogen monoxide and nitrogen dioxide in the formula row.

Formula			
colour	colourless	colourless	brown
odour	choking, can trigger asthma attacks	not detectable, reacts with O <sub>2</sub> in air	choking
acidity/basicity	acid	neutral	acid
MP (°C)	-75	-164	-11
BP (°C)	-10	-152	21
water solubility	reacts, very soluble	does not react, 0.05 g/L	reacts, very soluble



### Sources:

You can find out information about sources of sulfur dioxide pollutant and oxides of nitrogen pollutant in Australia from the National Pollutant Inventory Database:

[www.environment.gov.au/epg/npi/database](http://www.environment.gov.au/epg/npi/database)

Access this web site and report on the listed source of SO<sub>2</sub> and NO<sub>x</sub> closest to where you live.

a) sulfur dioxide

---

b) oxides of nitrogen (NO<sub>x</sub>)

---

# Chemistry

HSC Course

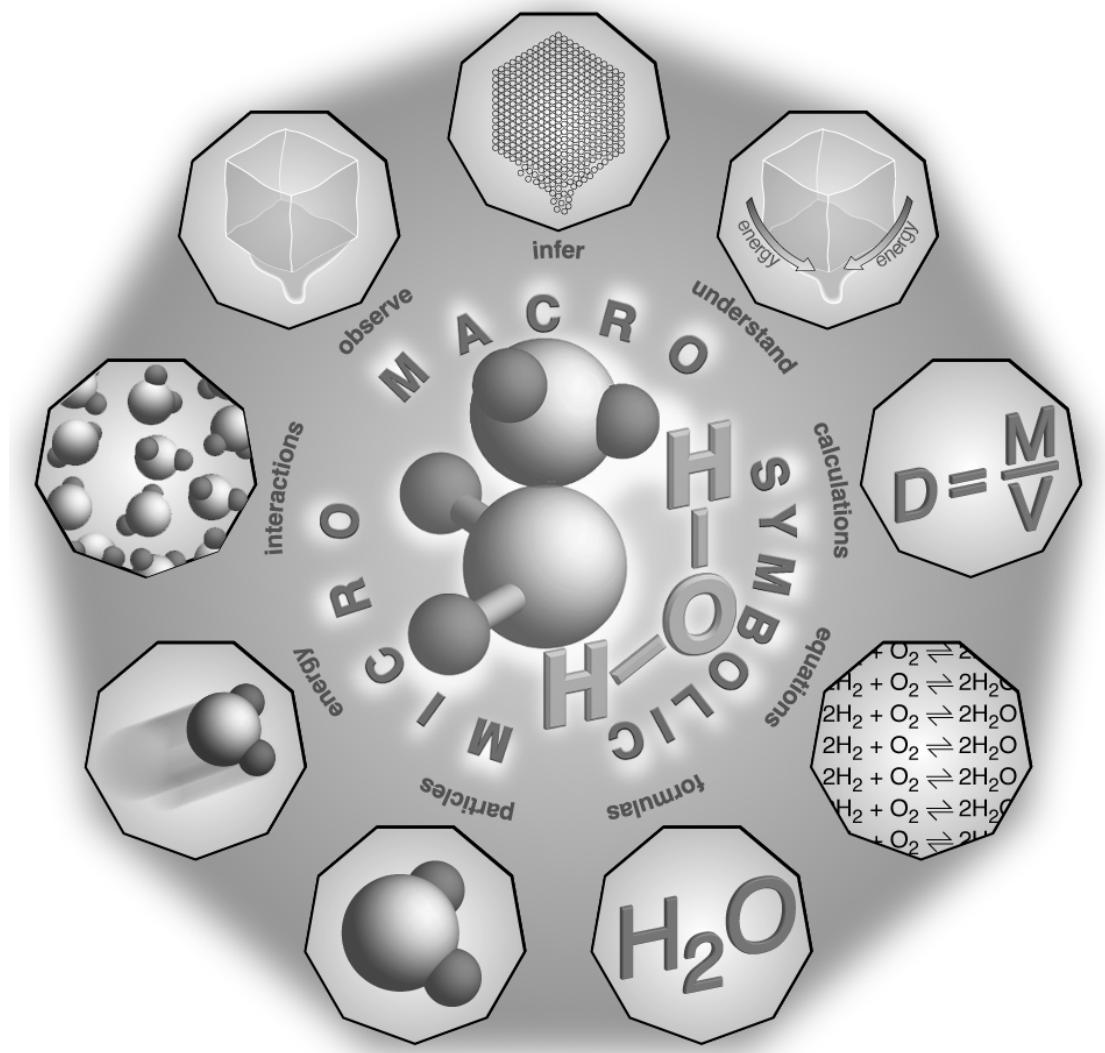
Stage 6

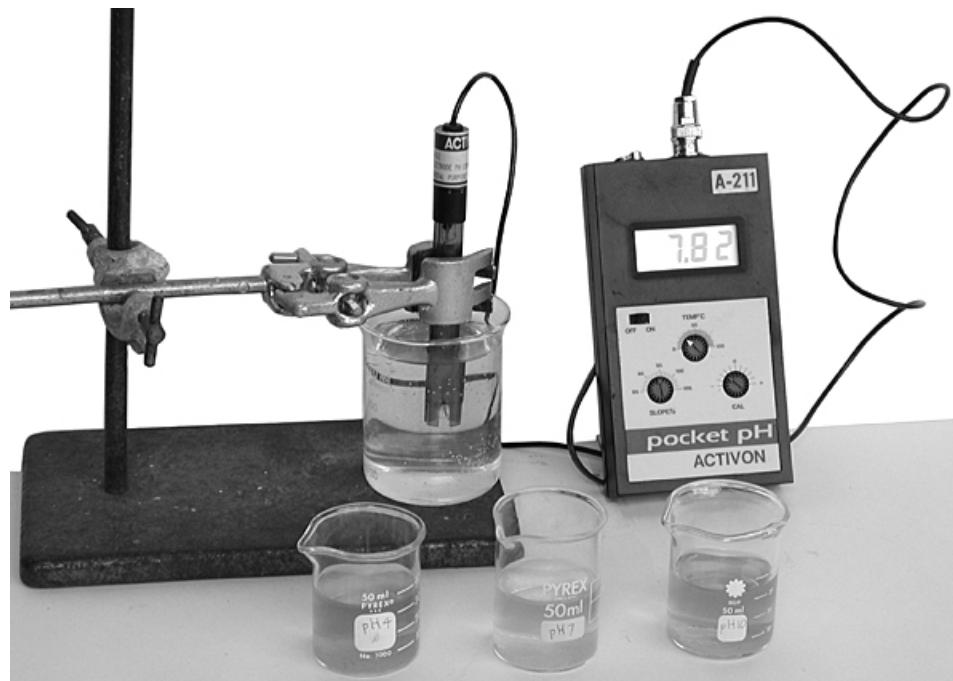
NEW SOUTH WALES  
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## The acidic environment

### Part 3: Acids in water



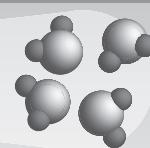


Pocket sized pH meter and pH 4.00, 7.00 and 10.00 calibration buffer solutions



'Stick' pH meter

Glass bulbs and protective plastic surrounds

**MACRO**observe  
infer  
understand**SYMBOLIC** $H_2O$  formulas  
equations  
calculations**MICRO**particles  
energy  
interactions

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## MACRO

observe  
infer  
understand



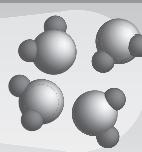
## SYMBOLIC

$\text{H}_2\text{O}$  formulas  
equations  
calculations



## MICRO

particles  
energy  
interactions



# Introduction

Acids have wide commercial applications and are part of most foods and drinks. Acids and bases are practically always used in water solution. To measure the acidity and basicity of water solutions the pH scale is used. It is important for students to understand the pH concept and relate it to the ions and molecules which produce pH changes.

In Part 3 you will be given opportunities to learn to:

- define acids as proton donors and describe the ionisation of acids in water
- identify acids such as acetic acid (ethanoic acid), citric acid (2-hydroxypropane-1,2,3-tricarboxylic acid), vitamin C and hydrochloric acid as naturally occurring acids, and acids such as sulfuric acid and hydrobromic acid as manufactured acids
- describe the use of the pH scale in comparing the concentrations of acids and alkalis
- describe acids and their solutions with the appropriate use of the terms strong, weak, concentrated and dilute
- identify pH as  $-\log_{10} [\text{H}^+]$  and explain that a change in pH of 1 means a ten-fold change in  $[\text{H}^+]$
- compare the relative strengths of equal concentrations of citric, acetic and hydrochloric acids and relate this to the degree of ionisation of their molecules
- describe the difference between a strong and a weak acid in terms of an equilibrium between the intact molecule and its ions.

In Part 3 you will be given opportunities to:

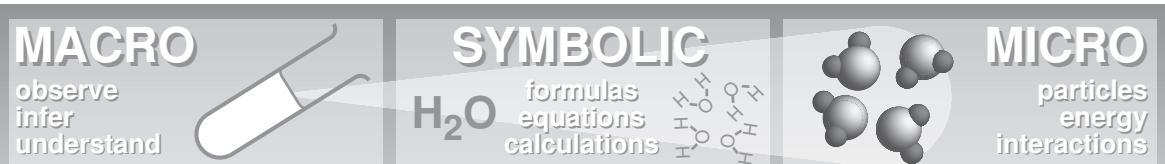
- solve problems and perform a first-hand investigation to use pH meters/probes and indicators to distinguish between acidic, basic and neutral chemicals
- plan and perform a first-hand investigation to measure the pH of identical concentrations of strong and weak acids

- gather and process information from secondary sources to write ionic equations to represent the ionisation of acids
- use available evidence to model the molecular nature of acids and simulate the ionisation of strong and weak acids
- gather and process information from secondary sources to explain the use of acidic oxides such as sulfur dioxide as food additives
- identify data, gather and process information from secondary sources to identify examples of naturally occurring acids and bases, their chemical composition and describe their pH in their naturally occurring form
- process information from secondary sources to calculate pH of strong acids given appropriate hydrogen ion concentrations
- analyse information from secondary sources to assess the use of neutralisation reactions as a safety measure or to minimise damage in accidents or chemical spills.

Also covered in this part is the section below in *italics* taken from the Preliminary Course 8.2.1

- apply systematic naming of inorganic compounds (binary compounds, hydroxides, acids, *salts of acids including the hydrogen salts of diprotic and tripotic acids*) as they are introduced in the laboratory.

Extracts from *Chemistry Stage 6 Syllabus* Board of Studies NSW, originally issued 1999. The most up-to-date version can be found on the Board's web site at [http://www.boardofstudies.nsw.edu.au/syllabus99/syllabus2000\\_list.html](http://www.boardofstudies.nsw.edu.au/syllabus99/syllabus2000_list.html)



# Hydrogen ions and protons in solution



The nucleus of most hydrogen atoms contains a single proton. (Most hydrogen atoms contain no neutrons. Neutrons are needed in the nucleus of all other atoms to keep the positively charged protons from repelling one another). Most hydrogen atoms and their nuclei can be represented by  $^1\text{H}$ .

- 1 a) Draw a  ${}^1_1\text{H}$  atom.

b) Draw a  ${}^1_1\text{H}$  nucleus.

2 If a hydrogen atom ionises losing an electron the change can be represented by:

hydrogen atom  $\rightarrow$  hydrogen ion + electron

Complete this chemical equation below using symbols for the underlined parts:

H  $\rightarrow$  \_\_\_\_\_ + \_\_\_\_\_

3 Compare your drawings in 1 with your equation in 2. Explain why the terms hydrogen ion and proton describe the same particle.

### Check your answers

In explaining acidity you will use the terms hydrogen ion and proton interchangeably. They both represent the same thing in water solutions.

# Acids in water

In Part 2 you learnt that acids in water:

- turn moist blue litmus red
- conduct electricity in solution
- have a sour taste.

These are properties observable in macro level (practical) situations.

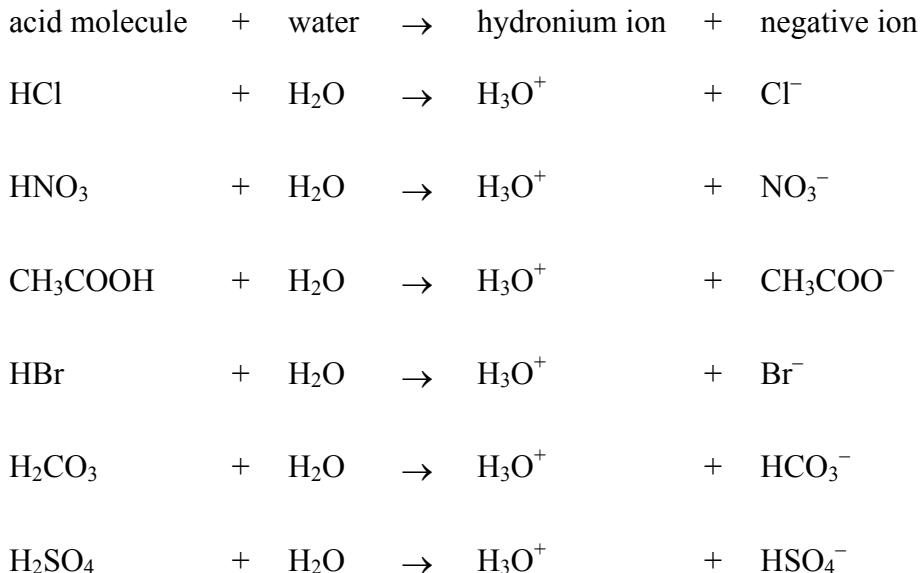
These properties can be explained at the micro (particle) level by the production of hydrogen ions/protons when an acid is in water solution:

- a hydrogen ion attaching itself to a blue litmus ion will produce a litmus molecule with a red colour
- an acid solution contains positive hydrogen ions and negative ions from the acid. These are mobile in solution enabling conduction.
- hydrogen ions stimulate sensors on the surface of the human tongue producing the sensation of sour taste.

An acid can be defined as a proton donor. When an acid molecule dissolves in water it can donate a proton/hydrogen ion to a water molecule.

The proton/hydrogen ion in water solution can be represented as  $\text{H}^+(\text{aq})$  or  $\text{H}_3\text{O}^+(\text{aq})$ .

$\text{H}_3\text{O}^+(\text{aq})$  is called the **hydronium** ion or hydrated hydrogen ion.



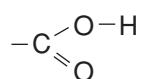


- 1 Print the acid names acetic, carbonic, hydrobromic, hydrochloric, nitric and sulfuric under the appropriate formula in the equation.
- 2 Print the negative ion names acetate, bromide, chloride, hydrogen carbonate, hydrogen sulfate, nitrate under the appropriate formula.

Check your answers.

These equations are ionic equations representing the ionisation of acid molecules in water.

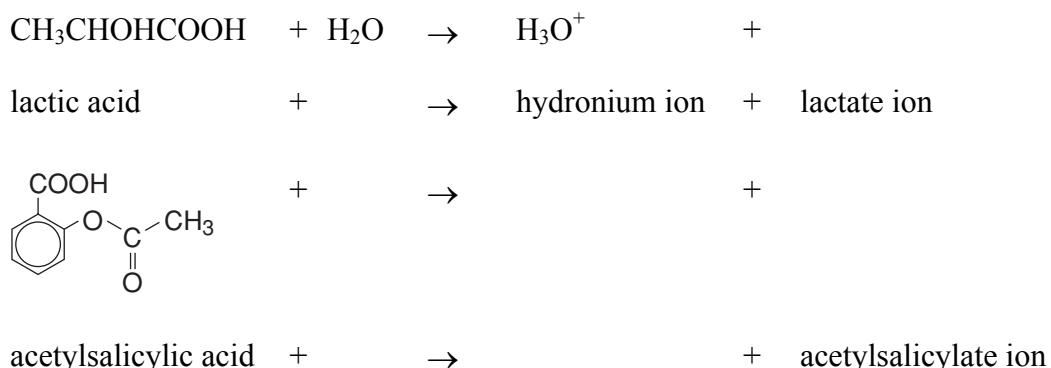
Most naturally occurring acids are organic compounds that contain the –COOH functional group.



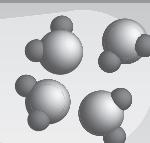
The shared electron pair (covalent bond) between the H and its attached O atom is weak. The electrons in this bond are attracted towards the two electronegative (electron-attracting) oxygen atoms. When these organic acids come in contact with water the hydrogen in the –COOH can be attracted away as a  $\text{H}^+$  to a polar water molecule. This leaves a  $\text{---COO}^-$  behind on the acid, changing the acid molecule to a negative ion.



Complete these ionic equations and word equations showing lactic acid (2-hydroxypropanoic acid) and acetylsalicylic acid (aspirin) undergoing ionisation in water solution. Note that a H atom covalently bonded to C as in  $\text{CH}_3$  never ionises. The acidic  $\text{H}^+$  comes from a –COOH group.



Check your answers.

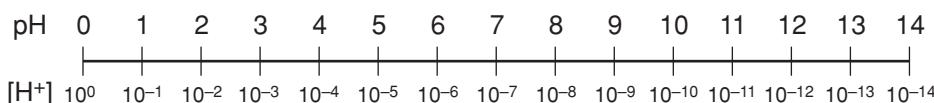
**MACRO**observe  
infer  
understand**SYMBOLIC** $\text{H}_2\text{O}$  formulas  
equations  
calculations**MICRO**particles  
energy  
interactions

# pH scale

The pH scale is a way of expressing the concentration of hydrogen ions  $\text{H}^+$  as a small number. pH is usually between 0 (extremely acid) and 14 (extremely basic). A pH of 7 represents a neutral solution.

pH can be calculated from the molar  $\text{H}^+$  concentration using:

$$\text{pH} = -\log_{10} [\text{H}^+] \quad \text{where } [ ] \text{ indicates molar concentration.}$$



pH values and corresponding  $[\text{H}^+]$ .

You will need a calculator that has a log key for calculating pH from molar  $\text{H}^+$  concentration. Instructions for a typical scientific calculator and a graphic (graphing) calculator are given below. You may need to modify these instructions for your calculator. It is very important that you know how to calculate pH on the type of calculator you will use in tests and exams. This skill is assumed knowledge for chemistry exams.

### Scientific calculator

- 1 Type in the  $[\text{H}^+]$  eg. 0.020.
- 2 Tap the log button (not the ln or  $\log_e$  buttons).
- 3 Tap the  $+-$  or  $(-)$  button.
- 4 Round off the figures you see to an appropriate number of significant figures; pH is 1.7 for a  $[\text{H}^+]$  of 0.020.

### Graphic calculator

- 1 Tap the  $+-$  or  $(-)$  button.
- 2 Tap the log button (not the ln or  $\log_e$  buttons).
- 3 Type in the  $[\text{H}^+]$  eg. 0.020.
- 4 Round off the figures you see to an appropriate number of significant figures; pH is 1.7 for a  $[\text{H}^+]$  of 0.020.



The table below shows the pH and  $[H^+]$  for a number of common solutions. Complete the table by using your calculator to calculate the missing values.

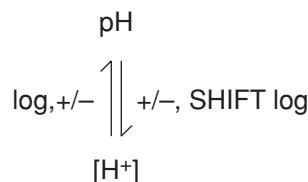
Solution	pH	$[H^+]$
human stomach gastric juice	2.0	$1.0 \times 10^{-2}$
lemon	2.3	$5.0 \times 10^{-3}$
vinegar		$1.6 \times 10^{-3}$
diet cola drink		$1.0 \times 10^{-3}$
lemonade	3.2	$6.3 \times 10^{-4}$
wine		$3.2 \times 10^{-4}$
tomato juice	4.2	$6.3 \times 10^{-5}$
black coffee		$1.0 \times 10^{-5}$
cow's milk		$4.0 \times 10^{-7}$
human saliva		$2.5 \times 10^{-7}$
human blood	7.4	$4.0 \times 10^{-8}$
fresh eggs		$1.5 \times 10^{-8}$
sea water	8.5	$3.2 \times 10^{-9}$
ammonia solution	11.0	$1.0 \times 10^{-11}$

Check your answers.

Now that you know how to work out pH from  $[H^+]$  can you work out a method to calculate  $[H^+]$  from pH? You will need to use a key marked SHIFT or INV or 2nd function before using the log key.

Use the figures in the suggested answers table to check that your method works.

If you are having difficulties and are using a scientific calculator try using the diagram below.



Remember you are expected to have developed the skill of calculations involving [H<sup>+</sup>] and pH.

The space below is available for you to record the methods that work for your calculator. Record your methods either as instructions or a diagram. You may be very glad you did so in a month's time or even the night before a major exam!

# Ionisation constant of water $K_w$

The product of  $[H^+]$  and  $[OH^-]$  in pure water, acidic solutions, basic solutions and salt solutions is constant for a given temperature.

This constant is called the ionisation constant of water  $K_w$ .

$$K_w = [H^+] \times [OH^-] = 1.0 \times 10^{-14} \text{ at } 25^\circ\text{C}.$$

In pure water  $[H^+] = [OH^-] = 10^{-7} \text{ mol L}^{-1}$  and  $pH = 7$ .

In an acidic solution  $[H^+] > 10^{-7} \text{ mol L}^{-1}$  and  $pH < 7$ .

In a basic solution  $[H^+] < 10^{-7} \text{ mol L}^{-1}$  and  $pH > 7$ .

Solution	pH	$[H^+]$	$[OH^-]$	$[H^+] \times [OH^-]$
1 M hydrochloric acid	0	$10^0 = 1$	$10^{-14}$	$10^{-14}$
0.1 M hydrochloric acid	1	$10^{-1}$	$10^{-13}$	$10^{-14}$
0.01 M hydrochloric acid	2	$10^{-2}$	$10^{-12}$	$10^{-14}$
carbonated drink	3	$10^{-3}$	$10^{-11}$	$10^{-14}$
	4	$10^{-4}$	$10^{-10}$	$10^{-14}$
	5	$10^{-5}$	$10^{-9}$	$10^{-14}$
	6	$10^{-6}$	$10^{-8}$	$10^{-14}$
pure water, no dissolved gas	7	$10^{-7}$	$10^{-7}$	$10^{-14}$
	8	$10^{-8}$	$10^{-6}$	$10^{-14}$
	9	$10^{-9}$	$10^{-5}$	$10^{-14}$
	10	$10^{-10}$	$10^{-4}$	$10^{-14}$
household ammonia	11	$10^{-11}$	$10^{-3}$	$10^{-14}$
0.01 M sodium hydroxide	12	$10^{-12}$	$10^{-2}$	$10^{-14}$
0.1 M sodium hydroxide	13	$10^{-13}$	$10^{-1}$	$10^{-14}$
1 M sodium hydroxide	14	$10^{-14}$	$10^0 = 1$	$10^{-14}$

Note that a change in pH of 1 unit means a ten-fold change in  $[H^+]$ .



Complete the following sentences:

- 1 1 M hydrochloric acid solution has \_\_\_\_ times as many H<sup>+</sup> than 0.1 M hydrochloric acid. The pH of the 1 M hydrochloric acid solution is \_\_\_\_ unit less than the 0.1 M hydrochloric acid.
- 2 1 M hydrochloric acid solution has \_\_\_\_ times as many H<sup>+</sup> than 0.01 M hydrochloric acid. The pH of the 1 M hydrochloric acid solution is \_\_\_\_ units less than the 0.01 M hydrochloric acid.
- 3 1 M sodium hydroxide solution has \_\_\_\_ times fewer H<sup>+</sup> than 0.1 M sodium hydroxide. The pH of the 1 M sodium hydroxide solution is \_\_\_\_ unit greater than the 0.1 M sodium hydroxide.
- 4 1 M sodium hydroxide solution has \_\_\_\_ times fewer H<sup>+</sup> than 0.01 M sodium hydroxide. The pH of the 1 M sodium hydroxide solution is \_\_\_\_ units \_\_\_\_ than the 0.01 M sodium hydroxide.
- 5 A change in pH of 1 unit means a 10<sup>1</sup> change in [H<sup>+</sup>].
- 6 A change in pH of \_\_\_\_ units means a 10<sup>2</sup> change in [H<sup>+</sup>].
- 7 A change in pH of \_\_\_\_ units means a 10<sup>3</sup> change in [H<sup>+</sup>].
- 8 A change in pH of 4 units means a \_\_\_\_ change in [H<sup>+</sup>].
- 9 A change in pH of \_\_\_\_ units means a 10<sup>7</sup> change in [H<sup>+</sup>].
- 10 A change in pH of \_\_\_\_ units means a 10<sup>2.5</sup> change in [H<sup>+</sup>].

Check your answers.

If the [H<sup>+</sup>] in a solution is known the [OH<sup>-</sup>] can be calculated using:

$$[\text{H}^+] \times [\text{OH}^-] = 10^{-14}.$$

Rearranging:  $[\text{OH}^-] = \frac{10^{-14}}{[\text{H}^+]}$

If an acid solution has [H<sup>+</sup>] = 10<sup>-4</sup> then [OH<sup>-</sup>] = 10<sup>-14</sup> / 10<sup>-4</sup> = 10<sup>-10</sup>

Similarly [H<sup>+</sup>] can be calculated from [OH<sup>-</sup>] using  $[\text{H}^+] = \frac{10^{-14}}{[\text{OH}^-]}$

If a base solution has [OH<sup>-</sup>] = 10<sup>-2.5</sup> then [H<sup>+</sup>] = 10<sup>-14</sup> / 10<sup>-2.5</sup> = 10<sup>-11.5</sup>

# pH measurement

## 1 Using an acid-base indicator

### a) On paper

A drop or two of a water solution to be tested is placed on the indicator paper. If the substance tested is a solid, liquid or gas and not in water solution the indicator paper must be first moistened with water.

The colour of the paper where the drop was placed is compared with a chart or table labelled with different colours and the corresponding pHs.

### b) In solution

A measured amount of indicator solution is added to the solution to be tested. The colour produced is either compared with a chart labelled with different colours and pHs or a table of written colours and pHs. If a chart or table is not available the colour is compared with the colours obtained for the indicator solution added to solutions of known pH.

## 2 Using a pH meter/probe system

Such a system consists of:

- an electrochemical cell **probe** (also called a combination electrode) made up of a pH sensitive glass electrode connected to a reference electrode; the potential (voltage) between the two electrodes depends on the  $[H^+]$  of the solution they are placed in
- a **millivoltmeter** designed to convert millivolts to pH units
- two buffer solutions of accurately known pH for **calibration** (**standardisation**) of the meter; for very accurate measurements the unknown pH value should be between the pHs of two **buffer** solutions used as **standards**.

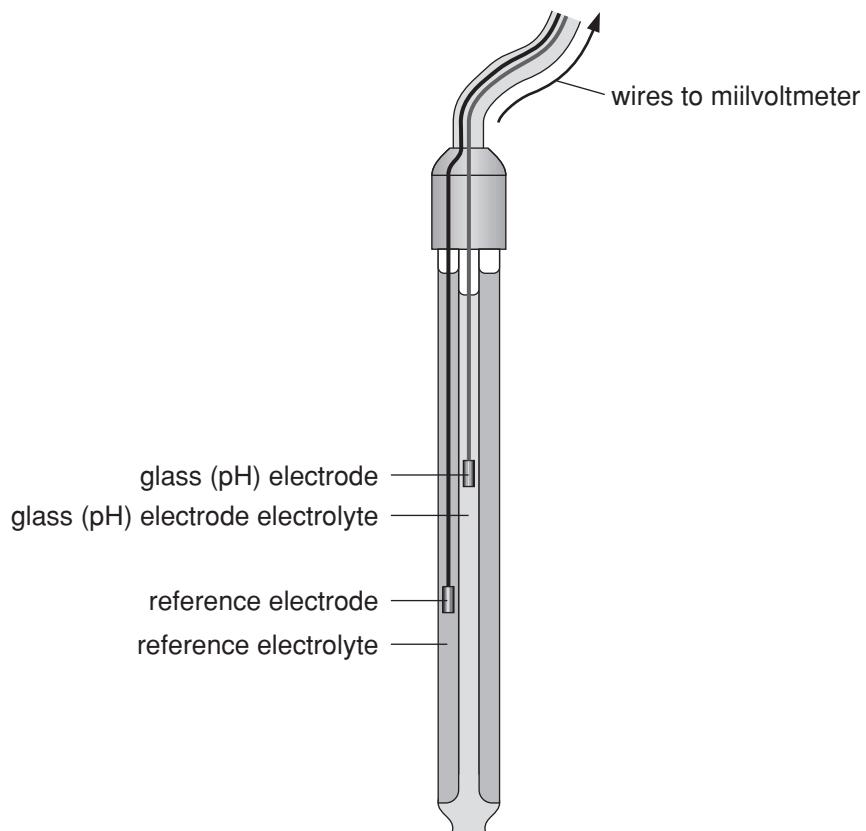
Simpler cheap and compact pH meters can be purchased from swimming pool and aquarium suppliers without buffer solutions for calibration.

These are suitable for comparing different solutions but without calibration do not give accurate readings.

The probe (combination electrode) should be stored in distilled water. If left dry it must be soaked in distilled water for at least 24 hours before becoming useable again.

Steps to be followed in using a pH meter are:

- 1 Adjust any temperature calibration on the meter for the temperature of the solutions being measured.
- 2 Calibrate the meter.
  - a) Remove the probe from its storage solution.
  - b) Wash the sensitive end of the probe with distilled water.
  - c) Place probe in a standard solution of known pH.
  - d) Adjust the meter to read pH of the standard solution.
  - e) Repeat steps b) to d) for another standard solution of different pH. Ideally the pHs of the two standard solutions should span the pH range to be measured.
- 3 Wash the sensitive end of the probe with distilled water before immersion in each solution to be tested.
- 4 When finished wash the probe with distilled water and store in storage solution.



pH probe combination electrode.

The very thin, scratch sensitive glass bulb is usually protected by a plastic surround.

# Using pH meters/probes and indicators



## What you will need:

- pH meter probe/system
- indicator such as universal indicator paper or red cabbage solution
- a variety of household chemicals already in water solution or able to dissolve in water.

Read and heed the label of any chemical before use. Do not use it if you do not have appropriate protective equipment or clothing.

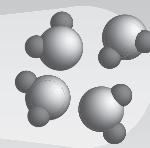


One of the pH test methods involves destructive testing. The solution tested should be disposed of responsibly after use. The other pH test method involves non-destructive testing – the solution tested is unaffected. Which method involves non-destructive testing?

## What you will do:

Test each chemical in turn and record its pH in the table below. Use the results to decide if the chemical is acidic, basic or neutral.

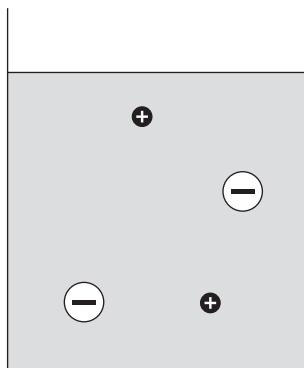
Chemical	pH	Acidic/neutral/basic

**MACRO**observe  
infer  
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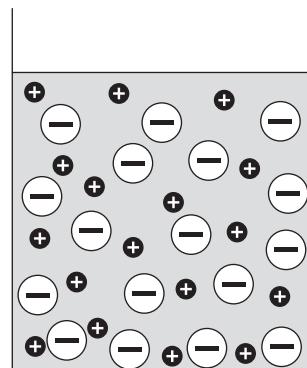
## Describing acid solutions

A **dilute** solution contains a small amount of solute in a given amount of solvent. Human stomach gastric juice could be described as a dilute hydrochloric acid solution averaging about  $0.01 \text{ molL}^{-1}$  HCl.

A **concentrated** solution contains a large amount of solute in a given amount of solvent. The hydrochloric acid manufactured industrially (about six million tonnes per year worldwide) could be described as concentrated hydrochloric acid solution averaging about  $10 \text{ molL}^{-1}$  HCl.



Dilute hydrochloric acid solution



Concentrated hydrochloric acid solution



The  $1000 \text{ kg}$  of gastric juice produced naturally each year in the stomachs of six billion humans totals about  $6 \times 10^9 \times 10^3 = 6 \times 10^{12} \text{ kg}$  of  $0.01 \text{ M}$  HCl.

If the  $6 \times 10^9 \text{ kg}$  of  $10 \text{ M}$  HCl manufactured industrially each year was diluted a thousand-fold it would produce about  $6 \times 10^{12} \text{ kg}$  of  $0.01 \text{ M}$  HCl.

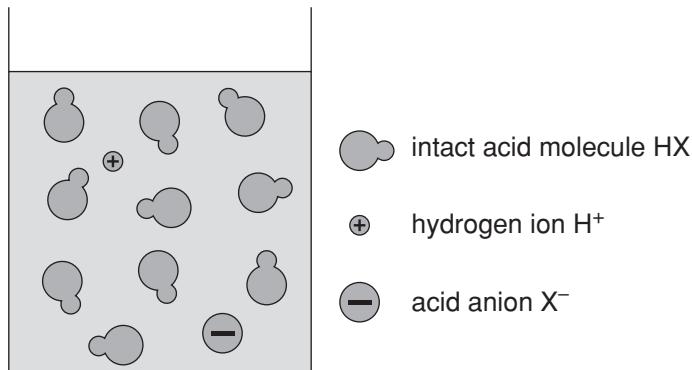
Should hydrochloric acid be called a naturally occurring acid or a manufactured acid? If hydrochloric acid was made using human genes **translocated** to another organism would it be natural or manufactured?



Citric acid is manufactured by fermentation of cane sugar or molasses using the fungus *Aspergillus niger*. Citric acid is found in most animal and plant tissues and fluids. Should citric acid be called manufactured or natural?

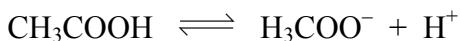
# Describing acids

Acids are described as weak or strong according to the degree of ionisation in solution. Acetic acid CH<sub>3</sub>COOH is called a **weak acid** because only a small proportion of the molecules in water are ionised.

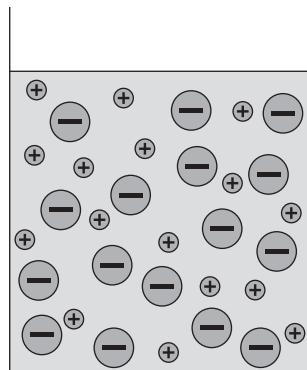


Weak acid solution.

When you write an equation for the ionisation of acetic acid in water it is best to use  $\rightleftharpoons$  to show that reaction is not complete. There is an equilibrium between intact unionised CH<sub>3</sub>COOH molecules and hydrogen ions H<sup>+</sup> and acetate ions CH<sub>3</sub>COO<sup>-</sup>.



Hydrochloric acid HCl is called a **strong acid** because the HCl molecules are practically all ionised in solution.



Strong acid solution

When you write an equation for the ionisation of hydrogen chloride gas in water it is best to use  $\rightarrow$  to show that the reaction goes to completion. Practically all the hydrochloric acid is in the form of ions rather than intact molecules.

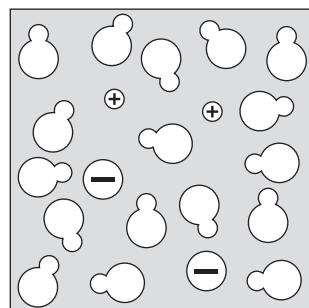
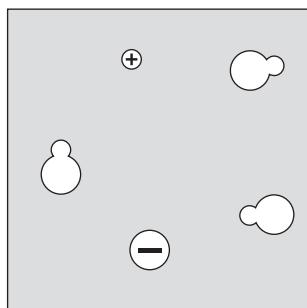
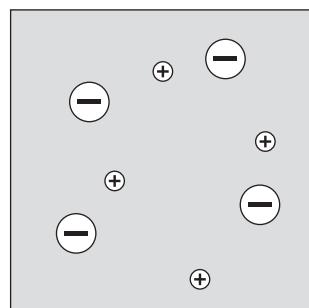
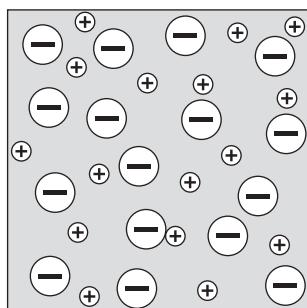




The four diagrams below show a:

- dilute weak acid solution
- concentrated weak acid solution
- dilute strong acid solution
- concentrated strong acid solution

Label each of the diagrams appropriately:



Check your answers.



Using equations explain why 1 M hydrochloric acid has a pH of about 0 while 1 M acetic acid has a pH of about 2.6.

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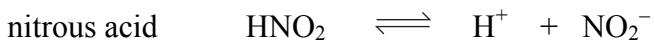
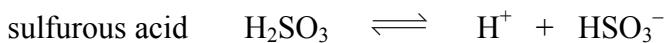
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Check your answer.

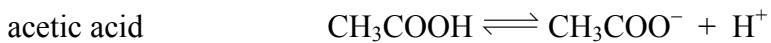
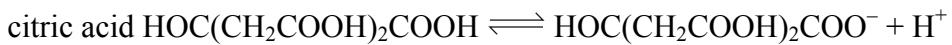
Strong acids, close to 100% ionised, include:



Acids, typically 10 to 20 % ionised, include:



Acids, typically 1% or less ionised, include:



as well as vitamin C, aspirin, lactic acid and most acids found in living things.

The degree (%) of ionisation depends on the concentration of the solution. In a dilute solution of an acid the degree of ionisation is greater than in a concentrated solution of the same acid.

The degree of ionisation of different acids can be compared by preparing solutions of the same molar concentration and measuring their pH. A pH meter is normally used as these can give accurate measurements to at least 0.1 pH units. Indicator paper or solution usually only give measurements to the nearest pH unit.

From the pH measurement the  $[\text{H}^+]$  can be calculated. Each molecule of acid that ionised would produce a  $\text{H}^+$ .  $[\text{H}^+]$  compared with the molar concentration of acid used to prepare the solution can be changed to a % degree of ionisation.

Consider the measurements made by a student on 0.10 M solutions of hydrochloric acid, acetic acid and citric acid:

<b>0.10 M acid solution</b>	<b>pH</b>	<b>[H<sup>+</sup>]</b>	<b>[H<sup>+</sup>] [acid]</b>	<b>[H<sup>+</sup>] [acid] × 100% (% ionisation)</b>
hydrochloric	1.0	$10^{-1} = 0.10$	0.10/0.10	100%
acetic	3.0	$10^{-3} = 0.001$	0.001/0.10	1%
citric	2.1	$10^{-2.1} = 0.008$	0.008/0.10	8%

- 1 Use the information in the table to list the three acids in order of strength. Place the strongest acid first and the weakest acid last.




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- 2 Another way to compare the strength of acids is to measure the electrical conductivity of acid solutions. Briefly explain how you would do this and why this method works.

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Check your answers.

## Polyprotic acids

Most acids are **monoprotic**, that is each molecule can release one proton in water.

**Polyprotic** acid molecules have the ability to release more than one proton per molecule. Sulfuric acid H<sub>2</sub>SO<sub>4</sub> and carbonic acid H<sub>2</sub>CO<sub>3</sub> are called **diprotic** while citric acid HOOC(CH<sub>2</sub>COOH)<sub>2</sub>COOH and phosphoric acid H<sub>3</sub>PO<sub>4</sub> are called **triprotic**.



- Why do you think sulfuric acid is called diprotic while phosphoric acid is called triprotic?

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Check your answer.

The release of the protons can be shown in stages or steps or in a single equation. For simplification H<sub>2</sub>O and states will be left out of equations.

1 Sulfuric acid H<sub>2</sub>SO<sub>4</sub>



Adding the two equations and cancelling the hydrogen sulfate ion on both sides gives:



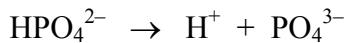
2 Carbonic acid H<sub>2</sub>CO<sub>3</sub>



Adding the two equations and cancelling the hydrogen carbonate ion on both sides gives:



3 Phosphoric acid H<sub>3</sub>PO<sub>4</sub>



Write the names dihydrogen phosphate ion, hydrogen phosphate ion and phosphate ion after the appropriate formula in the three equations.

Check your answers.

Overall the change is H<sub>3</sub>PO<sub>4</sub> → 3H<sup>+</sup> + PO<sub>4</sub><sup>3-</sup>

In a phosphoric acid solution there will be a mixture of intact H<sub>3</sub>PO<sub>4</sub> molecules, H<sub>2</sub>PO<sub>4</sub><sup>-</sup> ions, HPO<sub>4</sub><sup>2-</sup> ions, PO<sub>4</sub><sup>3-</sup> ions and H<sup>+</sup> ions.

### Hydrogen salts

Salts of polyprotic acids which still contain some hydrogen from the acid are called hydrogen salts or bisalts. An older name, acid salts, is misleading because some hydrogen salts produced basic solutions.

Sodium hydrogen carbonate NaHCO<sub>3</sub> is often labelled sodium bicarbonate.

Sodium hydrogen sulphate  $\text{NaHSO}_4$  similarly can be labelled sodium bisulfate.

You need to be very careful putting charges on hydrogen phosphate ions.

Dihydrogen phosphate ion is formed when a neutral phosphoric acid molecule  $\text{H}_3\text{PO}_4$  has lost one hydrogen ion. Thus it has a charge of one minus and is written as  $\text{H}_2\text{PO}_4^-$ .

Hydrogen phosphate ion is formed when  $\text{H}_3\text{PO}_4$  loses two hydrogen ions. Thus its charge is two minus and is  $\text{HPO}_4^{2-}$ .

## Acids can be ions as well as molecules

Not all acids are molecular. **Polyatomic ions** that contain hydrogen can often release protons in water.



Find four polyatomic ions in the previous section that can release a proton. Name each ion and write an equation showing each behaving as an acid.

Ion name	Equation showing ion as an acid

Check your answers.



A student prepared 1 M solutions of various salts and measured the pH of the solutions. Here are the results.

- |                                |                               |
|--------------------------------|-------------------------------|
| sodium hydrogen carbonate      | $\text{NaHCO}_3$ pH 9         |
| potassium hydrogen carbonate   | $\text{KHCO}_3$ pH 9          |
| sodium hydrogen sulfate        | $\text{NaHSO}_4$ pH 2         |
| potassium hydrogen sulfate     | $\text{KHSO}_4$ pH 2          |
| potassium hydrogen phosphate   | $\text{K}_2\text{HPO}_4$ pH 9 |
| potassium dihydrogen phosphate | $\text{KH}_2\text{PO}_4$ pH 5 |

Explain why:

- 1 NaHCO<sub>3</sub> and KHCO<sub>3</sub> solutions have the same pH.

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- 2 K<sub>2</sub>HPO<sub>4</sub> and KH<sub>2</sub>PO<sub>4</sub> solutions have a different pH.

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Check your answers.

Further testing enabled the student to arrange some molecular and ionic acids from strongest to weakest:

<b>strongest acid</b>	HCl
	H <sub>2</sub> SO <sub>4</sub>
	HNO <sub>3</sub>
	HSO <sub>4</sub> <sup>-</sup>
	H <sub>2</sub> SO <sub>3</sub>
	H <sub>3</sub> PO <sub>4</sub>
	CH <sub>3</sub> COOH
	H <sub>2</sub> CO <sub>3</sub>
	H <sub>2</sub> S
	HSO <sub>3</sub> <sup>-</sup>
	H <sub>2</sub> PO <sub>4</sub> <sup>-</sup>
	NH <sub>4</sub> <sup>+</sup>
	H <sub>2</sub> O
	HCO <sub>3</sub> <sup>-</sup>
	HPO <sub>4</sub> <sup>2-</sup>
<b>weakest acid in this list</b>	HS <sup>-</sup>



- 1 Write the name of each acid to the left of its formula. If you are not sure of the acid name you will find it in the following alphabetic list: acetic, ammonium ion, carbonic, dihydrogen phosphate ion, hydrochloric, hydrogen carbonate ion, hydrogen phosphate ion, hydrogen sulfate ion, hydrogen sulfide, hydrogen sulfide ion, hydrogen sulfite ion, nitric, phosphoric, sulfuric, sulfurous, water.
- 2 Write an equation after each formula showing it behaving as an acid.

Check your answers.

## Conductivity of pure acid and acid solution

Electrical conductivity measurements can give useful indications of the species present in pure liquid acid and in acid solutions.

If the conductivities of pure liquid sulfuric acid and pure liquid phosphoric acid are tested they are found to be extremely low. If the samples are left exposed to air and tested later, the conductivity is found to increase. The higher the humidity of the air the greater the increase in conductivity.

- 1 Are pure liquid acids such as  $\text{H}_2\text{SO}_4$  and  $\text{H}_3\text{PO}_4$  made up of molecules or ions? Explain your reasoning.



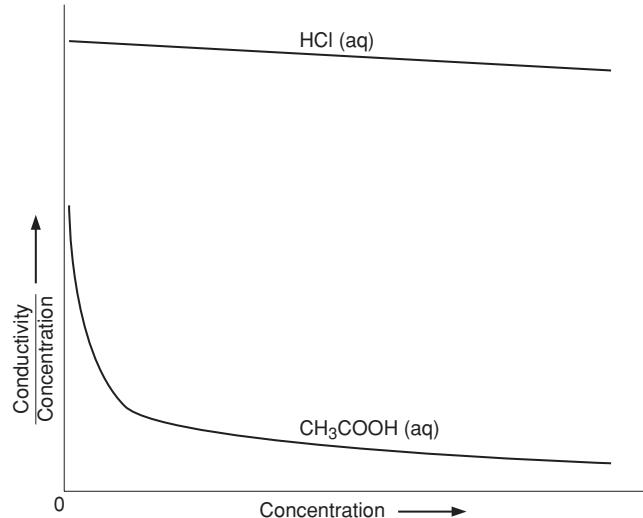
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- 2 Explain why the conductivities of pure  $\text{H}_2\text{SO}_4$  and  $\text{H}_3\text{PO}_4$  increase on exposure to air.

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- 3 You would expect that the conductivity of an acid solution is proportional to the concentration (doubling the concentration could double the number of ions and this would double the conductivity). If the conductivity is divided by the concentration and plotted against concentration different patterns are obtained for hydrochloric acid and acetic acid solutions. Explain the different patterns shown.

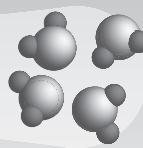


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Check your answers.

**MACRO**observe  
infer  
understand**SYMBOLIC** $H_2O$  formulas  
equations  
calculations**MICRO**particles  
energy  
interactions

## Acids as food additives

Acids such as vinegar have been added to foods as preservatives for thousands of years. The acid prevents spoiling by microbes. The acidic oxide  $SO_2$  is the strongest acid used as a food additive. In Australia dried fruit is allowed up to 3 g/kg while wine can contain up to 0.3 g/L of  $SO_2$ .

The sulfur dioxide can be added as the gas  $SO_2$ , its solution in water  $H_2SO_3$  or as a sulfite, bisulfite or metabisulfite salt. Any product with food codes 220 to 225 or 228 has one of these acidic substances added. The concentration of the additive decreases with time and this limits the **shelf life** of the food. Some people, often asthmatics, are allergic to these sulfur dioxide producing additives

Here are some acidic food preservatives listed from strongest down to weakest acid:

sulfur dioxide  $SO_2$

lactic acid  $CH_3CHOHCOOH$

nitrous acid  $HNO_2$

formic acid  $HCOOH$

benzoic acid  $C_6H_5COOH$

sorbic acid  $CH_3CH=CHCH=CHCOOH$  acetic acid  $CH_3COOH$

propionic (propanoic) acid  $CH_3CH_2COOH$

Just as sulfur dioxide is sometimes added as a sulfite salt these preservatives can be added to food as lactate, nitrite, formate, benzoate, sorbate, acetate and propionate (propanoate) salts.



Examine the labels on food containers in a kitchen or supermarket for food additive codes 220, 221, 222, 223, 224, 225 or 228 or the terms sulfur dioxide, sulfite, bisulfite or metabisulfite. List three products found to contain one of these sulfur dioxide preservatives.

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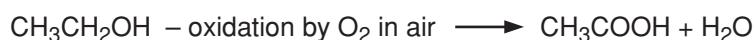
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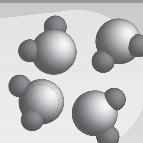
Some food acids such as phosphoric acid are mostly manufactured industrially while others such as citric acid in citrus fruits occur naturally. The phosphoric acid used in cola drinks and the citric acid added as a food preservative to many foods have usually been manufactured industrially. Acetic acid, citric acid, ascorbic acid (vitamin C) and hydrochloric acid all occur naturally but when required in food processing are increasingly being manufactured synthetically.

Pure acids which are manufactured in large quantities worldwide are shown in the table above.

<b>Acid</b>	<b>Formula</b>	<b>Uses</b>
sulfuric	H <sub>2</sub> SO <sub>4</sub>	fertilisers, petrochemicals, metal treatment
phosphoric	H <sub>3</sub> PO <sub>4</sub>	fertilisers, metal treatment, cola drinks
nitric	HNO <sub>3</sub>	fertilisers, explosives, dyes
hydrochloric	HCl	food, metallurgy and oil industries
acetic	CH <sub>3</sub> COOH	food preservation, plastics, solvents

Most acetic acid is made industrially by the petrochemical industry. Only small amounts are made by using micro-organisms to oxidise the alcohol CH<sub>3</sub>CH<sub>2</sub>OH in red wine or white wine:



**MACRO**observe  
infer  
understand**SYMBOLIC** $H_2O$  formulas  
equations  
calculations**MICRO**particles  
energy  
interactions

# Calculating pH of solutions

The  $[H^+]$  of strong acid solutions and the  $[OH^-]$  of strong base solutions can be calculated from the concentration of acid or base. The pH can then be calculated from the  $[H^+]$  or  $[OH^-]$ .

**Examples:**

1 Calculate the pH of the following strong acids

- a) 0.01 M HCl(aq)

Each HCl molecule produces one  $H^+$  on ionisation.

$$[H^+] = [HCl] = 0.01 \text{ M} \quad pH = -\log[H^+] = -\log 0.01 = 2.0$$

- b) 0.05 M HNO<sub>3</sub>(aq)

Each HNO<sub>3</sub> molecule produces one  $H^+$  on ionisation.

$$[H^+] = [HNO_3] = 0.05 \text{ M} \quad pH = -\log[H^+] = -\log 0.05 = 1.3$$

2 Calculate the pH of the following strong bases

- a) 0.01 M NaOH(aq)

Each unit of NaOH that dissolves releases one  $OH^-$

$$[OH^-] = 0.01 \text{ M} \quad [H^+] = 10^{-14} / 10^{-2} = 10^{-12}$$

$$pH = -\log[H^+] = -\log 10^{-12} = 12$$

- b) 0.005 M Ba(OH)<sub>2</sub>(aq)

Each unit of Ba(OH)<sub>2</sub> dissolves releasing two  $OH^-$

$$[OH^-] = 2 \times 0.005 \text{ M} = 0.01 \text{ M} \quad [H^+] = 10^{-14} / 10^{-2} = 10^{-12}$$

$$pH = -\log[H^+] = -\log 10^{-12} = 12$$

## MACRO

observe  
infer  
understand



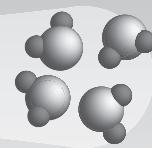
## SYMBOLIC

$H_2O$  formulas  
equations  
calculations



## MICRO

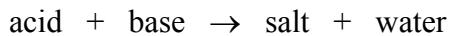
particles  
energy  
interactions



# Appendix: Cleaning up spills

Remember: acid or base spilt on any part of your body must be washed away with a steady flow of water for at least 5 minutes.

This appendix outlines what to do if acid or base is spilt but not on a person's body. Neutralisation reactions are used to clean up spills of acid or base.



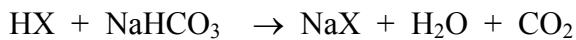
## Acid spills

Acids can be in solid form or liquid form.

If the acid is solid scoop it up on to paper being careful not to produce dust in the air. Wrap the solid in paper and dispose of in the garbage.

If a liquid acid is spilt add sodium hydrogen carbonate  $NaHCO_3$  (also known as sodium bicarbonate/bicarb of soda/baking soda). If the liquid was concentrated or pure acid put on protective clothing and eyewear. Mix and add water if necessary to form a **slurry**. (Caution: this reaction may be vigorous producing heat and corrosive fumes).

The acid  $HX$  reacts with the  $NaHCO_3$  producing salt  $NaX$ , carbon dioxide and water:



The slurry of salt and water should be mopped up and washed down the sink. Wash the spill site with  $NaHCO_3$  solution then rinse with water.

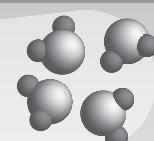
## **Alkali (water soluble base) spills**

Wear gloves, protective eyewear and clothing.

If solid has been spilt sweep up, transfer to a plastic bucket (alkali reacts with some metals) and neutralise with vinegar. Carry out the addition of vinegar slowly as heat may be released.

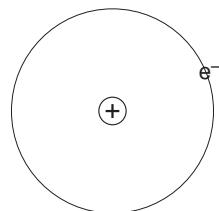
If liquid alkali solution is spilt add vinegar or  $\text{NaHCO}_3$ , mop up and wash down the sink.

An area where ammonia solution has been spilt must be well-ventilated before cleaning up.

**MACRO**observe  
infer  
understand**SYMBOLIC** $\text{H}_2\text{O}$  formulas  
equations  
calculations**MICRO**particles  
energy  
interactions

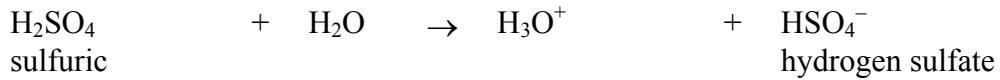
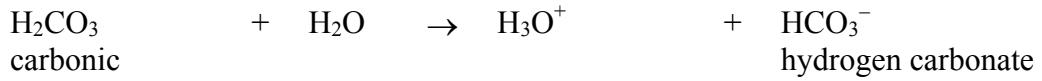
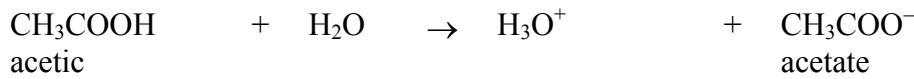
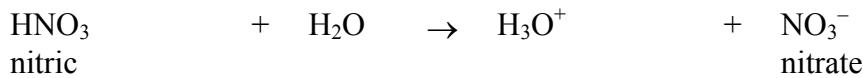
# Suggested answers

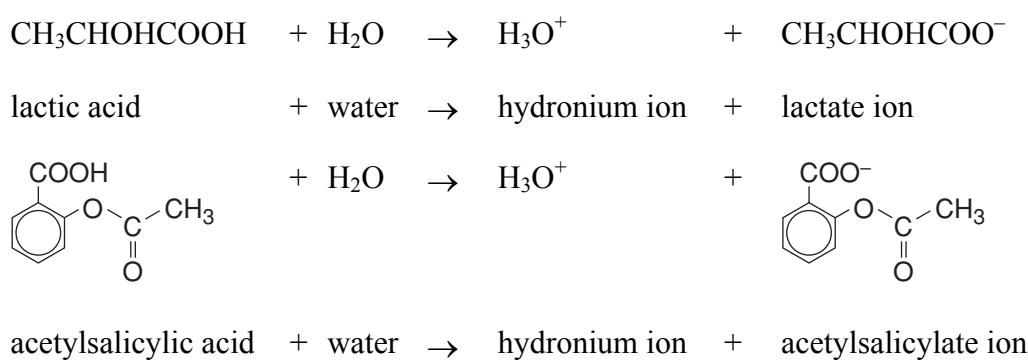
## Hydrogen ions and protons

1 a)  ${}^1_1\text{H}$  atomb)  ${}^1_1\text{H}$  nucleus

3 When a hydrogen atom loses an electron and forms a hydrogen ion the  $\text{H}^+$  consists of a single proton. Thus the terms hydrogen ion and proton can be used interchangeably when discussing acidity.

## Acids in water

acid molecule + water  $\rightarrow$  hydronium ion + negative ion



## pH scale

Solution	pH	$[\text{H}^+]$
human stomach gastric juice	2.0	$1.0 \times 10^{-2}$
lemon	2.3	$5.0 \times 10^{-3}$
vinegar	2.8	$1.6 \times 10^{-3}$
diet cola drink	3.0	$1.0 \times 10^{-3}$
lemonade	3.2	$6.3 \times 10^{-4}$
wine	3.5	$3.2 \times 10^{-4}$
tomato juice	4.2	$6.3 \times 10^{-5}$
black coffee	5.0	$1.0 \times 10^{-5}$
cow's milk	6.4	$4.0 \times 10^{-7}$
human saliva	6.6	$2.5 \times 10^{-7}$
human blood	7.4	$4.0 \times 10^{-8}$
fresh eggs	7.8	$1.5 \times 10^{-8}$
seawater	8.5	$3.2 \times 10^{-9}$
ammonia solution	11.0	$1.0 \times 10^{-11}$

## **Ionisation constant of water $K_w$**

- 1 1 M hydrochloric acid solution has *10* times as many  $H^+$  than 0.1 M hydrochloric acid. The pH of the 1 M hydrochloric acid solution is *1* unit less than the 0.1 M hydrochloric acid.
- 2 1 M hydrochloric acid solution has *100* times as many  $H^+$  than 0.01 M hydrochloric acid. The pH of the 1 M hydrochloric acid solution is *2* units less than the 0.01 M hydrochloric acid.
- 3 1 M sodium hydroxide solution has *10* times fewer  $H^+$  than 0.1 M sodium hydroxide. The pH of the 1 M sodium hydroxide solution is *1* unit greater than the 0.1 M sodium hydroxide.
- 4 1 M sodium hydroxide solution has *100* times fewer  $H^+$  than 0.01 M sodium hydroxide. The pH of the 1 M sodium hydroxide solution is *2* units *greater* than the 0.01 M sodium hydroxide.
- 5 A change in pH of 1 unit means a  $10^1$  change in  $[H^+]$ .
- 6 A change in pH of 2 units means a  $10^2$  change in  $[H^+]$ .
- 7 A change in pH of 3 units means a  $10^3$  change in  $[H^+]$ .
- 8 A change in pH of 4 units means a  $10^4$  change in  $[H^+]$ .
- 9 A change in pH of 7 units means a  $10^7$  change in  $[H^+]$ .
- 10 A change in pH of 2.5 units means a  $10^{2.5}$  change in  $[H^+]$ .

## **Using pH meter/probes and indicators**

Provided the test probe is washed with distilled water between uses the pH meter/probe method is non-destructive testing.

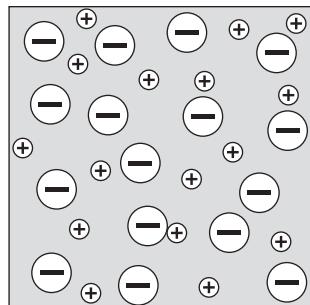
## **Describing acid solutions**

If quality counts most the hydrochloric acid manufactured industrially is purer (as well as more concentrated) than the hydrochloric acid in the mixtures produced in animal stomachs and gases emitted by volcanoes. If quantity counts most the hydrochloric acid produced naturally in the stomachs of other animals as well as humans would be greater than manufactured HCl. Also large amounts of HCl(g) which can dissolve in water forming HCl(aq) are produced naturally by volcanoes.

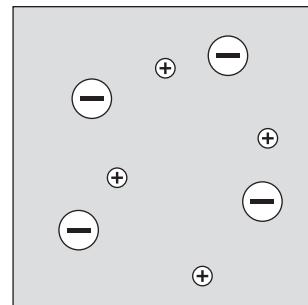
Increasingly chemicals are being made industrially using **transgenic** organisms. Chemicals such as insulin for diabetics and ethanol for renewable fuel are made by fermentation organisms containing translocated genes.

Whether citric acid is called manufactured or natural depends on context.  
At a factory or food processing plant you would probably call it  
manufactured; if investigating animal or plant biochemistry natural.

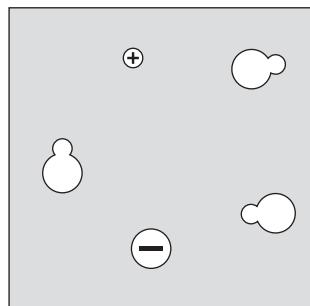
## Describing acids



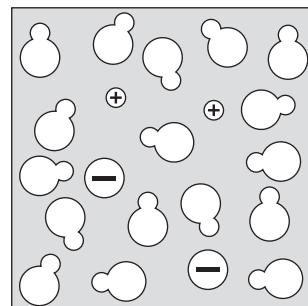
Concentrated strong acid solution.



Dilute strong acid solution.



Dilute weak acid solution.



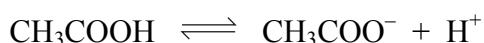
Concentrated weak acid solution.

1 M hydrochloric acid solution has a pH of about 1 because practically every HCl molecule ionises to produce a H<sup>+</sup>:



The concentration of H<sup>+</sup> equals the concentration of strong acid used to prepare the solution = 1 M = 10<sup>0</sup> M.

1M acetic acid solution has a pH of about 2.6 because only a small proportion of the weak acid molecules ionise:



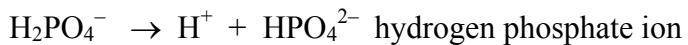
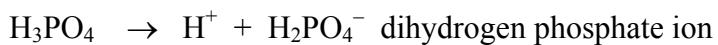
The acetic acid concentration is close to 1 (actually 1–10<sup>-2.6</sup>) M while the concentration of H<sup>+</sup> equals 10<sup>-2.6</sup>.

- 1 STRONGEST hydrochloric citric acetic WEAKEST
- 2 Apply a fixed DC voltage between two electrodes and measure the current that flows. Strong acid solutions contain more ions than weak acid solutions of the same acid concentration. Thus strong acid solutions have higher electrical conductivity.

## Polyprotic acids

Sulfuric acid is called a diprotic acid because each molecule can release up to two protons:  $\text{H}_2\text{SO}_4 \rightarrow 2\text{H}^+ + \text{SO}_4^{2-}$

Phosphoric acid is called a triprotic acid because each molecule can release up to three protons:  $\text{H}_3\text{PO}_4 \rightarrow 3\text{H}^+ + \text{PO}_4^{3-}$



## Acids can be ions as well as molecules

ion name	equation showing ion as an acid
hydrogen sulfate	$\text{HSO}_4^- \rightarrow \text{H}^+ + \text{SO}_4^{2-}$
hydrogen carbonate	$\text{HCO}_3^- \rightarrow \text{H}^+ + \text{CO}_3^{2-}$
dihydrogen phosphate	$\text{H}_2\text{PO}_4^- \rightarrow \text{H}^+ + \text{HPO}_4^{2-}$
hydrogen phosphate	$\text{HPO}_4^{2-} \rightarrow \text{H}^+ + \text{PO}_4^{3-}$

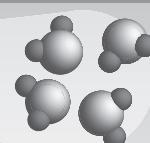
- 1  $\text{HCO}_3^-$  ions at the same concentration in both  $\text{NaHCO}_3$  and  $\text{KHCO}_3$  produce the same concentration of  $\text{H}^+$  ions
- 2 The  $\text{HPO}_4^{2-}$  and  $\text{H}_2\text{PO}_4^-$  ions have different acid strengths. The solution with the lowest pH contains the strongest acid.



phosphoric acid	$\text{H}_3\text{PO}_4 \rightarrow \text{H}^+ + \text{H}_2\text{PO}_4^-$
acetic acid	$\text{CH}_3\text{COOH} \rightarrow \text{H}^+ + \text{CH}_3\text{COO}^-$
carbonic acid	$\text{H}_2\text{CO}_3 \rightarrow \text{H}^+ + \text{HCO}_3^-$
hydrogen sulfide	$\text{H}_2\text{S} \rightarrow \text{H}^+ + \text{HS}^-$
hydrogen sulfite ion	$\text{HSO}_3^- \rightarrow \text{H}^+ + \text{SO}_3^{2-}$
dihydrogen phosphate ion	$\text{H}_2\text{PO}_4^- \rightarrow \text{H}^+ + \text{HPO}_4^{2-}$
ammonium ion	$\text{NH}_4^+ \rightarrow \text{H}^+ + \text{NH}_3$
water	$\text{H}_2\text{O} \rightarrow \text{H}^+ + \text{OH}^-$
hydrogen carbonate ion	$\text{HCO}_3^- \rightarrow \text{H}^+ + \text{CO}_3^{2-}$
hydrogen phosphate ion	$\text{HPO}_4^{2-} \rightarrow \text{H}^+ + \text{PO}_4^{3-}$
hydrogen sulfide ion	$\text{HS}^- \rightarrow \text{H}^+ + \text{S}^{2-}$

## Conductivity of pure acid and acid solution

- 1 The extremely low conductivity indicates that pure sulfuric acid and phosphoric acid are made up of molecules. Molecules are electrically neutral and so liquid containing only molecules is non-conducting.
- 2 Exposure to air enables the absorption of water molecules (typically 1 to 3% of air). These react with the acid molecules producing ions. Ions moving in a liquid can conduct electricity.
- 3 As the concentration of HCl increases practically all of the HCl remains in ionic form as  $\text{H}^+$  and  $\text{Cl}^-$ . In very dilute acetic acid many molecules have ionised but as the  $\text{CH}_3\text{COOH}$  concentration increases a lower proportion of the molecules are ionised. Degree of ionisation of:
  - a weak acid such as  $\text{CH}_3\text{COOH}$  depends on its concentration
  - a strong acid such as HCl is independent of its initial concentration.

**MACRO**observe  
infer  
understand**SYMBOLIC** $H_2O$  formulas  
equations  
calculations**MICRO**particles  
energy  
interactions

## Exercises – Part 3

Exercises 3.1 to 3.4

Name: \_\_\_\_\_

### Exercise 3.1: Planning an investigation to measure the pH of identical concentrations of strong and weak acids

Sodium hydrogen sulfate  $NaHSO_4$  forms a moderately strong acid solution. Sodium hydrogen carbonate  $NaHCO_3$  does not.

Both of these substances are white powders.

- a) Suppose you had to prepare identical concentrations of these two salt solutions so that you could compare their pHs. Would it be fairest to compare identical w/w, v/v, w/v or molar solutions? Explain your reasoning.

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- b) List the equipment you would need to carry out this activity.

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- c) Assess any risks. What safety precautions should you take?

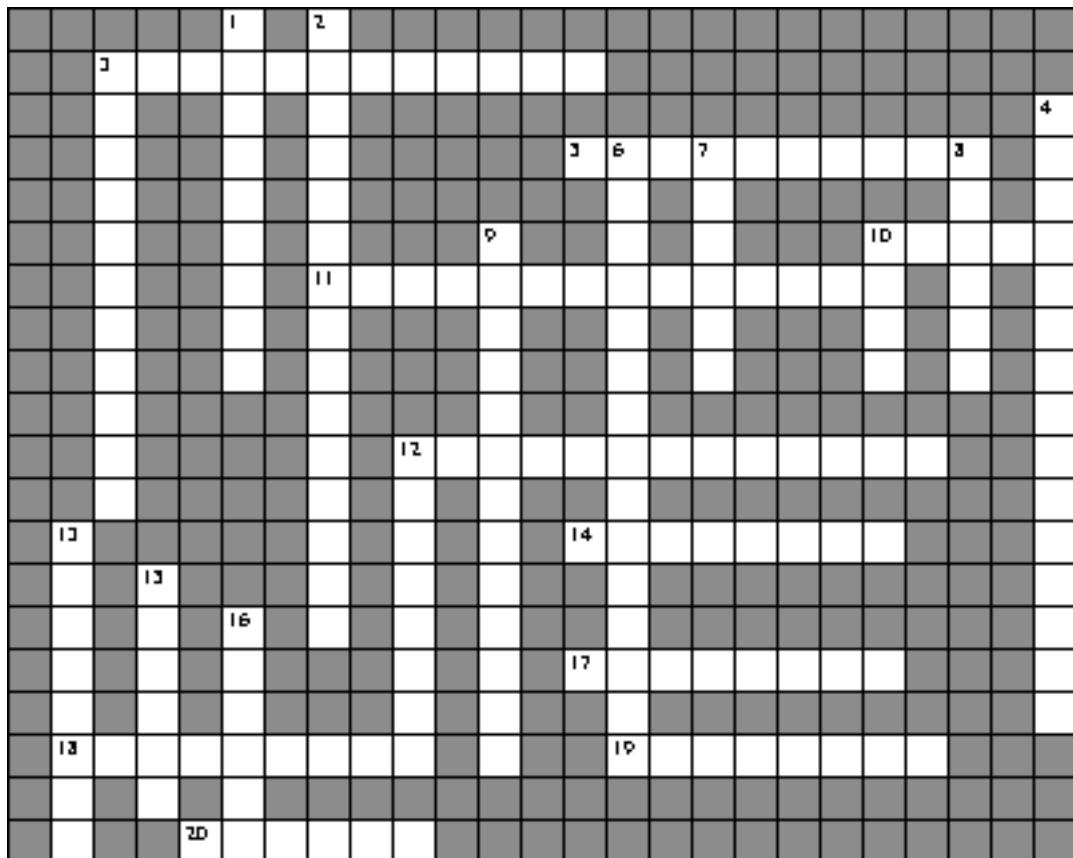
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## Exercise 3.2: Acid crossword

Complete this crossword before doing Exercise 3.3.



### Across clues

- 3 acid in stomach gastric juice
- 5 acid in some cola drinks
- 10 main solvent for acids
- 11 produced by oxidation of  $\text{SO}_2$
- 12 acidic oxide in carbonated drinks
- 14 acid most produced worldwide
- 17 ascorbic acid's vitamin name
- 18 acid releasing three protons
- 19 systematic name for acetic acid
- 20 acid in vinegar

### Down clues

- 1 systematic name for propionic
- 2 aspirin acid
- 3 similar properties to  $\text{HCl(aq)}$
- 4  $+ \text{H}_2\text{O} \rightarrow \text{HNO}_2 + \text{HNO}_3$
- 6 rotten egg gas
- 7 acid that completely ionises
- 8 acid in citrus fruits
- 9 strong food preservative acid
- 10 acid with low ionisation
- 12 acid in soda water
- 13 acid releasing two protons
- 14 acid formed in yoghurt
- 16 acid made from nitrate salts

### **Exercise 3.3: Naturally occurring acids and bases**

Identify two naturally occurring acids and two naturally occurring bases. These could be compounds or mixtures. Find out their chemical composition (formula of compounds, other components of mixtures) and their pH in their naturally occurring form and concentrations. Use this information to complete the table.

Name	Composition	pH

### **Exercise 3.4: Calculating pH of solutions**

- 1 Calculate the pH of a 0.008 M solution of the strong acid HBr(aq).

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- 2 a) Calculate the hydroxide ion concentration of a limewater solution. Limewater is saturated solution of the strong base calcium hydroxide  $\text{Ca}(\text{OH})_2$ . Limewater has a  $\text{Ca}(\text{OH})_2$  concentration of 0.015 M.

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- b) Use the calculated  $[\text{OH}^-]$  to calculate the pH of the limewater solution.

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# Chemistry

HSC Course

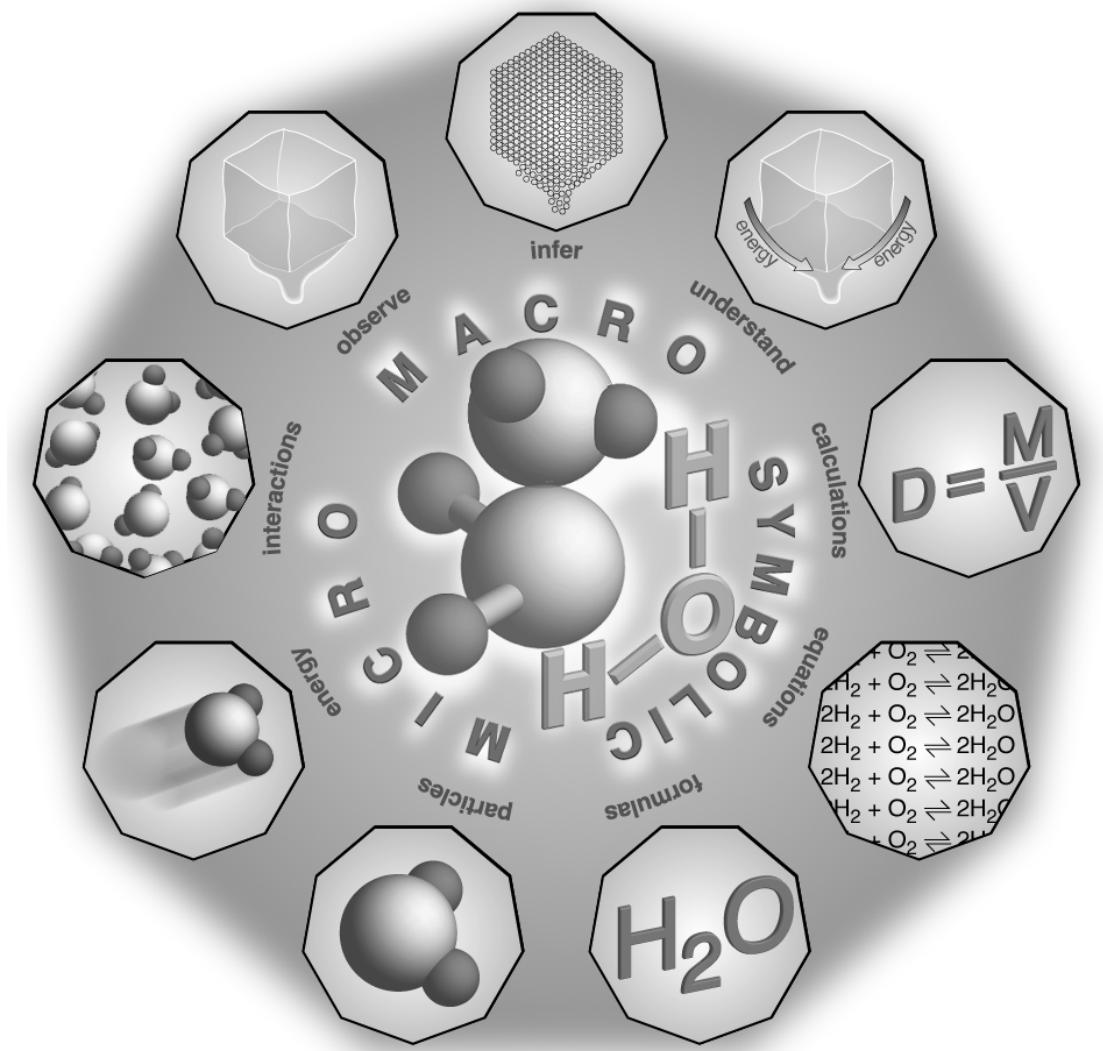
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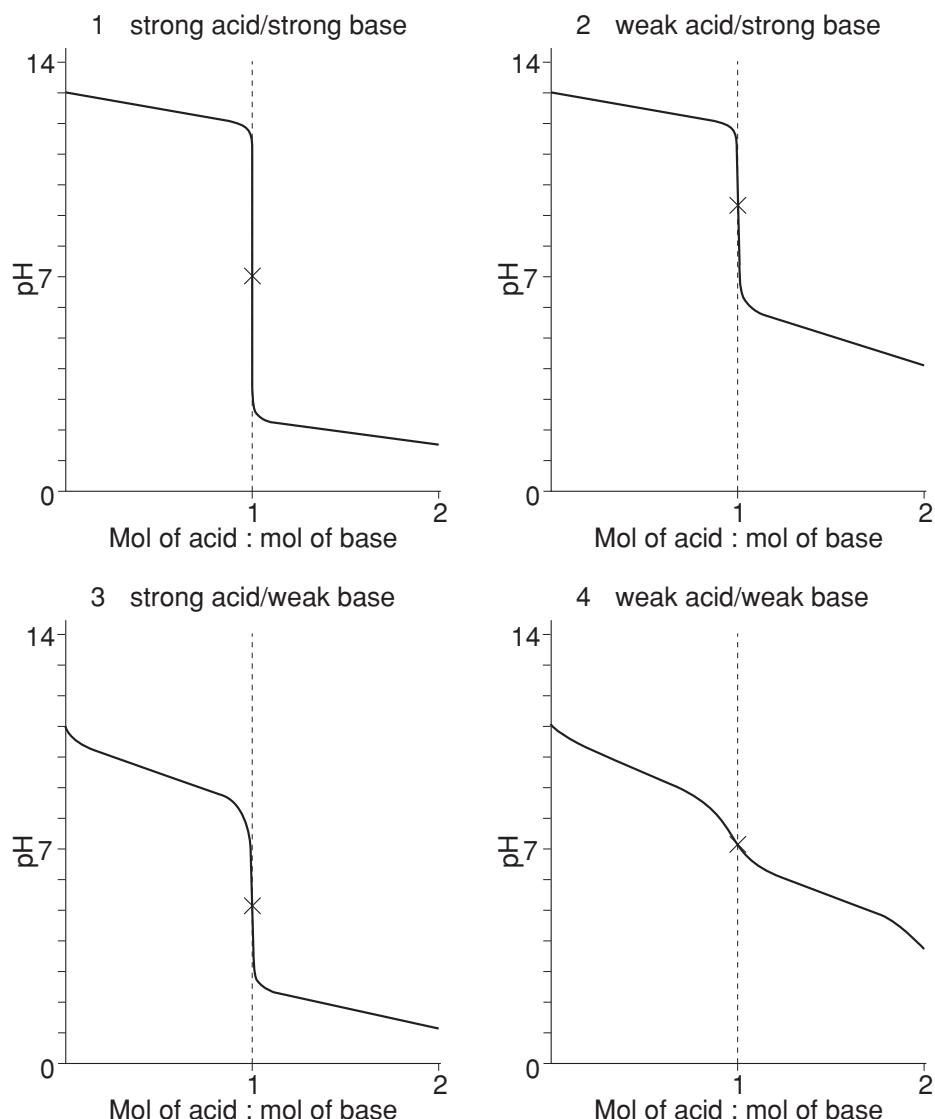
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AND TRAINING



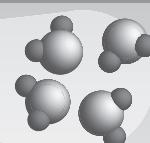
## The acidic environment

### Part 4: Acid concepts and titrations





Pipette filler with three valves.

**MACRO**observe  
infer  
understand**SYMBOLIC** $H_2O$  formulas  
equations  
calculations**MICRO**particles  
energy  
interactions

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## MACRO

observe  
infer  
understand



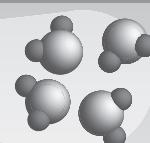
## SYMBOLIC

$H_2O$  formulas  
equations  
calculations



## MICRO

particles  
energy  
interactions



# Introduction

Acids and bases have been used by humans for thousands of years. Tannic acids were used in tanning leather, acetic acid (vinegar) and citric acid (citrus fruits) to alter taste of and preserve foods. Ashes from the burning of timber, consisting of metal oxides and carbonates, were used to change fats and oils to soap, to neutralise acids and to make glass.

As chemical theory developed, definitions of acids and bases changed from operational (based on macro level practical observations) to conceptual (based on micro level ideas about particles). Refining of the acid definition has seen the concept that an acid 'contains replaceable hydrogen' change to 'a producer of hydrogen ions in water' and then to 'a proton donor in both aqueous and non-aqueous solvents'. Similarly the base definition has changed from 'a producer of hydroxide ions in water' to 'a proton acceptor in any solvent'.

All the practical activities in this module are based upon water as the solvent. The Bronsted-Lowry theory (defining an acid as a proton donor and a base as a proton acceptor) will be used to interpret the reactions occurring in practical activities.

In Part 4 you will be given opportunities to learn to:

- outline the historical development of ideas about acids including those of:
  - Arrhenius
  - Lavoiser
  - Davy
- outline the Bronsted-Lowry theory of acids and bases
- describe the relationship between an acid and its conjugate base and a base and its conjugate acid
- identify a range of salts which form acidic, basic or neutral solutions and explain their acidic, neutral or basic nature
- identify conjugate acid/base pairs

- identify amphiprotic substances and write equations to describe their behaviour in acidic and basic solutions
- identify neutralisation as a proton transfer reaction which is exothermic
- describe the correct technique for conducting titrations and preparation of standard solutions
- qualitatively describe the effect of buffers with reference to a specific example in a natural system
- outline the Lewis definition of an acid
- assess the importance of each definition in terms of understanding.

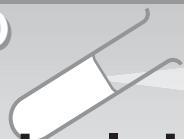
In Part 4 you will be given opportunities to:

- choose equipment and perform a first-hand investigation to identify the pH of a range of salt solutions
- analyse information from secondary sources to predict combining volumes and masses of acids and bases and products in neutralisation reactions
- 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
- process information from secondary sources to:
  - visualise the rearrangement of particles, and change in electrical conductivity and pH, which occurs in the reaction vessel during titration
  - explain the purpose of accuracy during titration
  - identify and describe modern analytical methods used to accurately measure concentrations
- perform a first-hand investigation to determine the concentration of a commercial acidic or alkaline substance such as vinegar, orange juice or window cleaner
- perform a first-hand investigation to measure temperature change during neutralisation and calculate the molar heat of reaction
- gather and process information from secondary sources to trace and describe developments in understanding and describing acid/base reactions and use available evidence to explain the expanded view of an acid as developed by Lewis.

Extracts from *Chemistry Stage 6 Syllabus* Board of Studies NSW, originally issued 1999. The most up-to-date version can be found on the Board's web site at [http://www.boardofstudies.nsw.edu.au/syllabus99/syllabus2000\\_list.html](http://www.boardofstudies.nsw.edu.au/syllabus99/syllabus2000_list.html)

**MACRO**

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infer  
understand



**SYMBOLIC**



formulas  
equations  
calculations



**MICRO**

particles  
energy  
interactions

## Historical development of ideas about acids

### Lavoisier

Antoine Lavoisier, the French chemist who established the quantitative science of chemistry investigated oxides of different elements. In the 1780s he found that non-metal oxides reacted with water producing acidic solutions. From this he concluded that 'an acid must contain oxygen'. However discoveries that HCl and HCN gases containing no oxygen react with water producing acid solutions contradicted his idea.

### Davy

In 1815 the English chemist Humphry Davy noted that all known 'acids contained hydrogen that could be replaced by reaction with a metal' and noted that compounds of metal with oxygen were bases.

Both Lavoiser's 'contains oxygen' and Davy's 'contains hydrogen' definitions were directly based on practical macro level observations. After their time the particle theory was largely accepted by chemists and successful definitions were conceptual. Conceptual definitions are based on concepts about the type of particles in acids and bases.

### Arrhenius

In 1884 the Swedish chemist Svante Arrhenius proposed that 'an acid was a substance that produced hydrogen ions  $\text{H}^+$  when dissolved in water'. He also proposed that 'a base was a substance that produced hydroxide ions  $\text{OH}^-$  in water'. These definitions, commonly used for acids and bases in junior high school, are restricted to the solvent water.

## Bronsted and Lowry

In 1923 the Danish chemist Johannes Bronsted and the British chemist Thomas Lowry independently realised that acid-base reactions involved proton transfer. The Bronsted-Lowry theory they developed defined an 'acid as a proton donor' and a 'base as a proton acceptor'.

These are the definitions used in senior high school. You will use these conceptual definitions to explain what is happening in your practical activities.

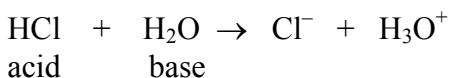
## Lewis

Towards the end of this part you will be introduced to acid-base definitions developed by the American chemist G. N. Lewis. The Lewis acid and Lewis base concepts were developed about the same time as the Bronsted-Lowry concepts but highlight the role of the electron pair. Today Lewis's definitions are used extensively for university level study especially in explaining the reactions of carbon compounds and biologically active molecules such as chlorophyll and haem (in the haemoglobin of red blood cells) that require metal ions to function.

## Bronsted-Lowry theory of acids and bases

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

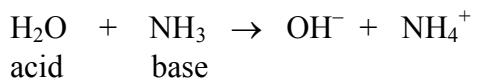
When a hydrogen chloride gas molecule dissolves in water a proton transfers from the HCl molecule to a water molecule:



An acid-base reaction is a proton transfer reaction.

(You have studied transfer reactions before. The metal reactions in the Preliminary module *Metals* and the electrochemistry reactions in the first HSC module were all electron transfer reactions commonly called oxidation-reduction reactions).

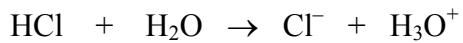
When an ammonia gas molecule dissolves in water a proton is transferred from a water molecule to the NH<sub>3</sub> molecule.



Note that water acts as a base  $\text{H}^+ + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+$  with HCl but as an acid  $\text{H}_2\text{O} \rightarrow \text{H}^+ + \text{OH}^-$  with  $\text{NH}_3$ .

This is another very important property of water to be added to the ones studied in the preliminary module *Water*. Water can act as an acid or a base depending on the chemical that it is in contact with.

HCl is a strong acid compared with water because the equilibrium position is largely on the right. Nearly 100% of the HCl molecules dissolved in water are ionised at any one time.



The  $\rightarrow$  signifies that a reaction goes to completion, that is, the equilibrium is largely to the right hand side.

$\text{CH}_3\text{COOH}$  is a weak acid in water because the equilibrium position is more to the left than to the right. Only about 1% of the  $\text{CH}_3\text{COOH}$  molecules dissolved in water are ionised at any one time.



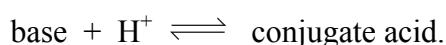
**MACRO**observe  
infer  
understand**SYMBOLIC**
 $\text{H}_2\text{O}$   
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 equations  
 calculations
**MICRO**particles  
energy  
interactions

# Conjugate acid-base pairs

In a conjugate acid-base pair:



OR

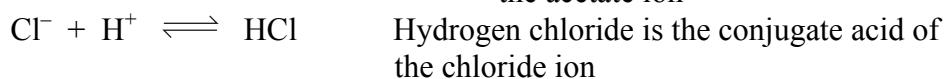
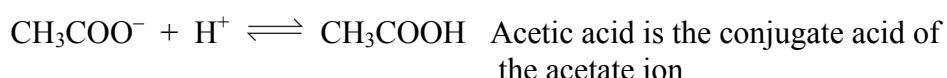
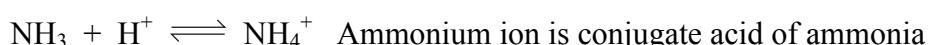
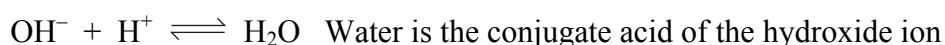


Every acid has a conjugate base:



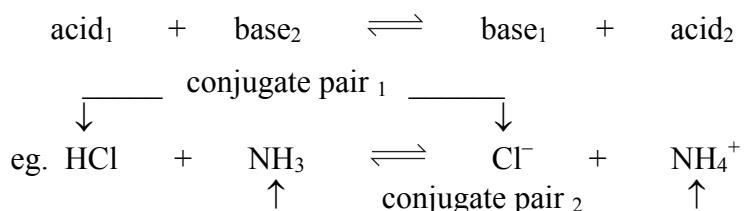
Note that the conjugate base has one fewer H and one more negative charge than its acid.

Similarly every base has a conjugate acid:



Note that the conjugate acid has one more H and one more positive charge than its base.

An acid-base reaction involves two different conjugate pairs





- 1 Give the formula and name of the conjugate base of:
  - a)  $\text{HNO}_3$
  - b)  $\text{H}_2\text{SO}_4$
  - c)  $\text{HCO}_3^-$
  - d)  $\text{HSO}_4^-$
- 2 Give the formula and name of the conjugate acid of
  - a)  $\text{S}^{2-}$
  - b)  $\text{HS}^-$
  - c)  $\text{CO}_3^{2-}$
  - d)  $\text{SO}_4^{2-}$
- 3 Give the formulas for the acid-base conjugate pairs in the following molecular equations. Part a) answer is given.
  - a)  $\text{NaOH} + \text{HCl} \rightarrow \text{NaCl} + \text{H}_2\text{O}$   
 *$\text{OH}^-$  and  $\text{H}_2\text{O}$ ,  $\text{HCl}$  and  $\text{Cl}^-$*
  - b)  $\text{KOH} + \text{HBr} \rightarrow \text{KBr} + \text{H}_2\text{O}$
  - c)  $\text{LiOH} + \text{HNO}_3 \rightarrow \text{LiNO}_3 + \text{H}_2\text{O}$
  - d)  $\text{NH}_3 + \text{HBr} \rightarrow \text{NH}_4\text{Br}$
  - e)  $\text{NaOH} + \text{NaHSO}_4 \rightarrow \text{Na}_2\text{SO}_4 + \text{H}_2\text{O}$
  - f)  $\text{H}_2\text{O} + \text{HS}^- \rightarrow \text{OH}^- + \text{H}_2\text{S}$

Check your answers.

# Strengths of acid-base pairs

Strongest acid		Weakest base in this list
HCl	strong	Cl <sup>-</sup> extremely weak
H <sub>2</sub> SO <sub>4</sub>	strong	HSO <sub>4</sub> <sup>-</sup> extremely weak
HNO <sub>3</sub>	strong	NO <sub>3</sub> <sup>-</sup> extremely weak
H <sub>3</sub> O <sup>+</sup>		H <sub>2</sub> O
HSO <sub>4</sub> <sup>-</sup>		SO <sub>4</sub> <sup>2-</sup>
H <sub>2</sub> SO <sub>3</sub>		HSO <sub>3</sub> <sup>-</sup>
H <sub>3</sub> PO <sub>4</sub>		H <sub>2</sub> PO <sub>4</sub> <sup>-</sup>
CH <sub>3</sub> COOH		CH <sub>3</sub> COO <sup>-</sup>
H <sub>2</sub> CO <sub>3</sub>		HCO <sub>3</sub> <sup>-</sup>
H <sub>2</sub> S		HS <sup>-</sup>
HSO <sub>3</sub> <sup>-</sup>		SO <sub>3</sub> <sup>2-</sup>
H <sub>2</sub> PO <sub>4</sub> <sup>-</sup>		HPO <sub>4</sub> <sup>2-</sup>
NH <sub>4</sub> <sup>+</sup>		NH <sub>3</sub>
HCO <sub>3</sub> <sup>-</sup>		CO <sub>3</sub> <sup>2-</sup>
HPO <sub>4</sub> <sup>2-</sup>		PO <sub>4</sub> <sup>3-</sup>
H <sub>2</sub> O	extremely weak	OH <sup>-</sup> strong
HS <sup>-</sup>	extremely weak	S <sup>2-</sup> strong
OH <sup>-</sup>	extremely weak	O <sup>2-</sup> strong

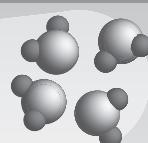
## Weakest acid in this list

Note that the stronger an acid the weaker is its conjugate base.

Similarly the stronger a base the weaker its conjugate acid.

Most acids and bases are called weak – both they and their conjugates are present in equilibrium in significant concentrations.

## Strongest base

**MACRO**observe  
infer  
understand**SYMBOLIC** $H_2O$  formulas  
equations  
calculations**MICRO**particles  
energy  
interactions

# Salt solutions

## Measuring the pH of salt solutions

**What you will need:**


- pH meter or a broad range indicator such as universal indicator paper or red cabbage solution (the red cabbage range of colours can be seen at [www.lmpc.edu.au/science](http://www.lmpc.edu.au/science))
- a range of salt solutions containing 5 % w/w salt. Use salts such as:
  - sodium chloride ( $NaCl$ ) table salt
  - sodium hydrogen carbonate ( $NaHCO_3$ ) baking soda
  - sodium carbonate ( $Na_2CO_3$ ) washing soda
  - sodium hydrogen sulfate ( $NaHSO_4$ ) toilet powder cleaner
  - magnesium sulfate ( $MgSO_4$ ) Epsom salt.

**What you will do:**

- 1 Measure the pH of each solution. Record in the table below.
- 2 Also measure the pH of the water used to prepare the solutions. \_\_\_\_\_
- 3 Classify the salts as acidic, basic or neutral. A salt is neutral if its pH is close to that of the water used to prepare the solution. Record in the table.

Salt name	Salt formula	Solution pH	Acidic/basic/neutral

Which salt solution contained the strongest acid? \_\_\_\_\_

Which salt solution contained the strongest base? \_\_\_\_\_

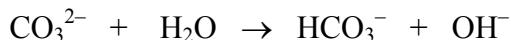
# Explaining acidic and basic salt solutions

Salts that produce acidic or basic solutions contain an ion which behaves as an acid (producing an acidic solution) or a base (producing a basic solution). Ions can react with water producing  $\text{H}_3\text{O}^+$  ions (acidic solution) or  $\text{OH}^-$  ions (basic solution).

If a salt solution is neutral either the ions do not react with water significantly, eg.  $\text{Na}^+\text{Cl}^-$  sodium chloride or the positive ions and negative ions have reacted with water to the same extent so that their acidic and basic effects cancel, eg.  $\text{CH}_3\text{COO}^-\text{NH}_4^+$  ammonium acetate.

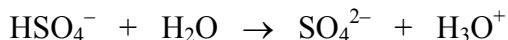
Adding a suitable solid salt to water is a convenient way of obtaining solutions in the pH range 2 – 12. Some household examples are:

- Use of sodium carbonate  $\text{Na}_2\text{CO}_3$  or potassium carbonate  $\text{K}_2\text{CO}_3$  in dishwasher powder.



The basic conditions resulting both react with acids in food to produce soluble salts and react with fats/oils to produce water soluble soaps.

- Use of sodium hydrogen sulfate  $\text{NaHSO}_4$  in powder toilet cleaner.



The acidic conditions resulting react with sticky materials reducing their adhesion to the toilet bowl.

The equations used to explain acidity or basicity of salt solutions are called hydrolysis equations. They show reaction of an ion with water.



Write a different hydrolysis equation for each to explain why:

1  $\text{NH}_4\text{Cl}$  solution is acidic

2  $\text{NaHCO}_3$  solution is basic

3  $\text{KH}_2\text{PO}_4$  solution is acidic

4  $\text{LiCH}_3\text{COO}$  solution is basic.

Check your answers

## Amphiprotic or amphoteric?

A molecule or ion which can behave as a proton donor or acceptor is called amphiprotic. Amphiprotic means protons on both sides, that is protons can be donated or accepted.

Water is an amphiprotic molecule.



Hydrogen carbonate ion is an amphiprotic ion.



Water solutions containing hydrogen carbonate ions are slightly basic showing that they behave more as a base than as an acid.

Hydrogen sulfate ion is an amphiprotic ion.



Water solutions containing hydrogen sulfate ions are acidic showing that they behave much more as an acid than as a base.

When an amphiprotic substance is placed into an acidic solution it behaves as a base. For example if sodium hydrogen carbonate, solid or solution, is added to an acid:  $\text{H}^+ + \text{HCO}_3^- \rightarrow \text{H}_2\text{CO}_3$

Similarly when an amphiprotic substance is placed into a basic solution it behaves as an acid.  $\text{OH}^- + \text{HCO}_3^- \rightarrow \text{H}_2\text{O} + \text{CO}_3^{2-}$



Write equations to show the behaviour of:

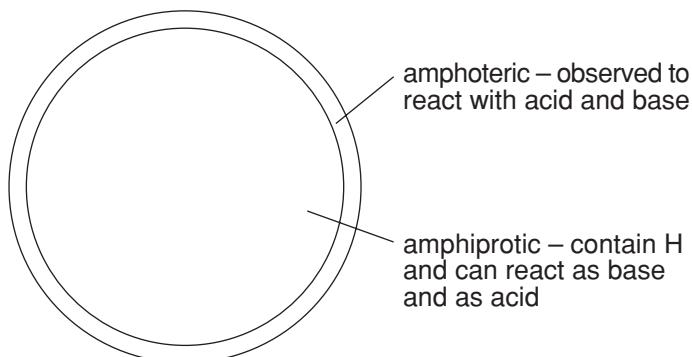
a)  $\text{HSO}_4^-$  in acidic solution

b)  $\text{HSO}_4^-$  in basic solution.

Check your answer.

Solvents such as water which can act as proton donors or proton acceptors are called amphiprotic solvents.

Another word which is often considered equivalent to amphiprotic is amphoteric. Amphoteric means 'can react with an acid and also react with a base'. Examples of amphoteric substances are the oxides of aluminium, zinc, chromium, tin and lead which react with both acid and base. However  $\text{Al}_2\text{O}_3$ ,  $\text{ZnO}$ ,  $\text{Cr}_2\text{O}_3$ ,  $\text{SnO}_2$  and  $\text{PbO}$  do not contain H and should not really be called amphoteric. It is probably best to think of amphoteric as the most common type of amphoteric substance.



In most texts amphoteric and amphiprotic are taken to have the same meaning.



- 1 Clarify (make clear or plain) which of the definitions amphiprotic or amphoteric is conceptual (based on micro level ideas about particles) rather than operational (based on macro level practical observations).

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- 2 Using equations justify (support an argument or conclusion) why the hydrogen phosphate ion is amphiprotic.

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- 3 Explain why aluminium oxide should be called amphoteric rather than amphiprotic.

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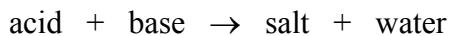
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Check your answers.

# Neutralisation reactions and equations

A neutralisation reaction is one in which an acid and base react, cancelling out the acidic and basic properties of the original acid and base. The products of the reaction are a salt and water.

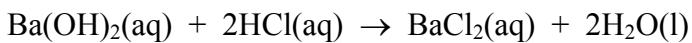


Writing balanced equations for neutralisation reactions is an important skill that underpins volumetric analysis of acids and bases.

A neutralisation reaction can be written in a number of ways:

- Full (sometimes called molecular or neutral formula) equation.  
All reactants are represented by neutral formulas. No ions are shown even though some reactants and products are ionic.

For the reaction of barium hydroxide solution with hydrochloric acid to form barium chloride salt and water the full equation is:



- Full ionic equation. All ions, including any which do not react (spectator ions), are shown.  
$$\text{Ba}^{2+}\text{(aq)} + 2\text{OH}^-\text{(aq)} + 2\text{H}^+\text{(aq)} + 2\text{Cl}^-\text{(aq)} \rightarrow \text{Ba}^{2+}\text{(aq)} + 2\text{Cl}^-\text{(aq)} + 2\text{H}_2\text{O(l)}$$
- Net ionic equation. Spectator ions which have not reacted are left out. In this example the  $\text{Ba}^{2+}\text{(aq)}$  and  $\text{Cl}^-\text{(aq)}$  which were present on the left hand side and are present unchanged on the right hand side are left out.  
$$\text{OH}^-\text{(aq)} + \text{H}^+\text{(aq)} \rightarrow \text{H}_2\text{O(l)}$$

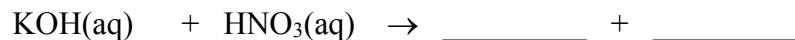
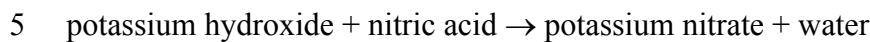
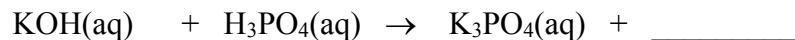
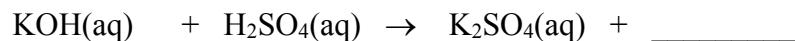
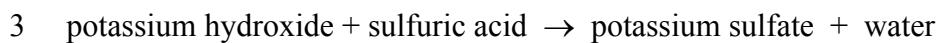
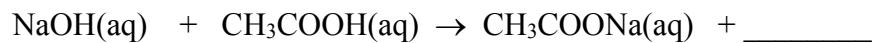
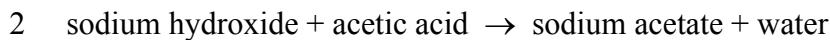
Note that:

- the full equation shows all the chemicals involved; this is the best sort of equation to use if you are going to do quantitative calculations such as for a **volumetric analysis**
- the full ionic equation shows all the ionic forms of the base, acid and salt solutions and the molecular nature of the water produced
- the net ionic equation focuses on what actually reacts (the hydroxide ions from the base and the hydrogen ions from the acid) and what is produced (water molecules).

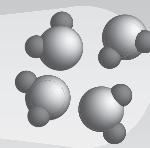
The salt formed, barium chloride, is not included in the net ionic equation because it is a soluble salt. If the salt formed was insoluble and thus precipitates out it would be included in the net ionic equation because its aqueous ions would have reacted together to form a solid. For insoluble salts the net ionic equation is the same as the full ionic equation.



Complete full (neutral formula) balanced equations for the following acid-base neutralisations. Assume reactions go to completion and all salts formed are soluble



Check your answers.

**MACRO**observe  
infer  
understand**SYMBOLIC** $H_2O$  formulas  
equations  
calculations**MICRO**particles  
energy  
interactions

# Volumetric analysis

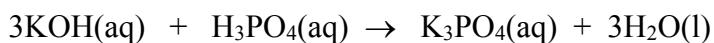
Volumetric analysis is quantitative analysis measuring volumes to find the concentration of a solution. This involves reacting a solution of known concentration with one of unknown concentration to determine the **equivalence point**.

The equivalence point is reached when the reactants have completely reacted together in the ratio of the mole quantities given in the balanced equation for the reaction.

The equivalence point for  $\text{NaOH(aq)} + \text{HCl(aq)} \rightarrow \text{NaCl(aq)} + \text{H}_2\text{O(l)}$  is when sodium hydroxide and hydrochloric acid have completely reacted in a 1:1 mole ratio.

The equivalence point for  $2\text{KOH(aq)} + \text{H}_2\text{SO}_4\text{(aq)} \rightarrow \text{K}_2\text{SO}_4\text{(aq)} + 2\text{H}_2\text{O(l)}$  is when potassium hydroxide and sulfuric acid have completely reacted in a 2:1 mole ratio.

Describe the equivalence point for the reaction:



Check your answer.

The equivalence point for a neutralisation reaction is measured by choosing an acid-base indicator that changes colour about the pH of the salt solution that was formed at the equivalence point.

strong acid + strong base  $\rightarrow$  neutral salt solution pH about 7  
bromothymol blue or litmus indicator solution change colour about pH 7

strong acid + weak base  $\rightarrow$  acidic salt solution pH below 7  
methyl orange indicator solution changes colour about pH 4

weak acid + strong base  $\rightarrow$  basic salt solution pH above 7  
phenolphthalein indicator solution changes colour about pH 9

The point at which the indicator changes colour is called the **end point**. An indicator is chosen with end point as close as possible to the equivalence point.

## Titration techniques

The technique in which one reactant in solution is slowly added to another reactant in solution until the end point is called **titration**.

### Titration technique using pipette and burette

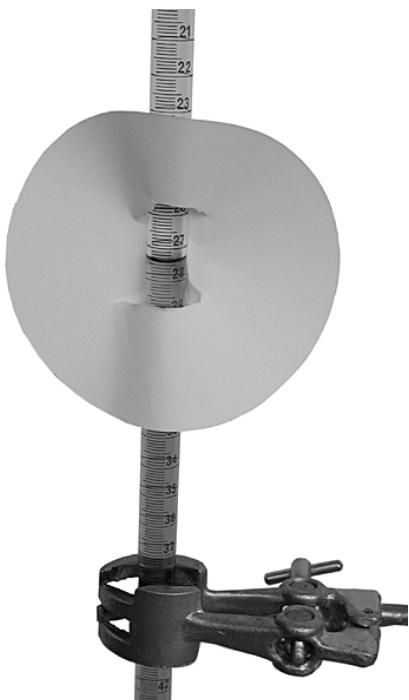
- 1 Pipette a fixed volume of one solution into a **conical flask**.
- 2 A pipette filler must be used. NEVER PIPETTE BY MOUTH.



- Attach the pipette to the pipette filler. Expel air by pressing the top filler valve between thumb and finger while squeezing the bulb with the other hand
- Hold the pipette so that the tip is well below the surface of liquid you wish to draw upwards. Press the middle valve between bulb and pipette so that liquid is drawn up into the pipette to above the filling mark.
- Carefully press the lowest valve so that air enters the top of the pipette and liquid drains slowly out until the bottom of the meniscus sits on the filling mark. The pipette should be vertical while doing this.
- Transfer the pipette to a conical flask. Press the lowest valve so that liquid drains into the conical flask. Hold the pipette tip against the inside of the flask for five seconds after draining is complete. The last drops inside the pipette must not be shaken or blown out.

- 2 Add at least two drops of a suitable indicator solution to the liquid in the conical flask. Choose an indicator that changes color about the pH of the salt solution formed at equivalence point.

3 Record the solution level in the burette.



Add this solution gradually from the **burette** while swirling the contents of the flask until the desired indicator colour change occurs. Control the flow of solution through the burette tap using left thumb in front of the tap and the left fingers behind.



Measurements and calculations should be to three significant figures using this type of equipment.

## Titration technique using syringes

Use a smaller syringe, say 3 mL, as the pipette, and a larger size, say 5 mL as the burette.

Syringes (without needles) can be purchased cheaply from pharmacies. The syringe sits in a container of its solution.

Draw the syringe plunger up slowly and eject the solution quickly to get rid of air bubbles inside the syringe.

Keep the tip of the syringe below the liquid surface while doing this.



- 1 Syringe a fixed volume of one solution into a **medicine measure**.
- 2 Add at least 2 drops of a suitable indicator solution to the liquid in the medicine measure. Choose an indicator that changes colour about the pH of the salt solution formed at equivalence point.
- 3 Record the solution level in the burette syringe. Always take readings for the same part of the syringe plunger tip. Add this solution gradually from the burette syringe while swirling the medicine measure until the desired colour change.



Measurements and calculations should be to two significant figures using this type of equipment.

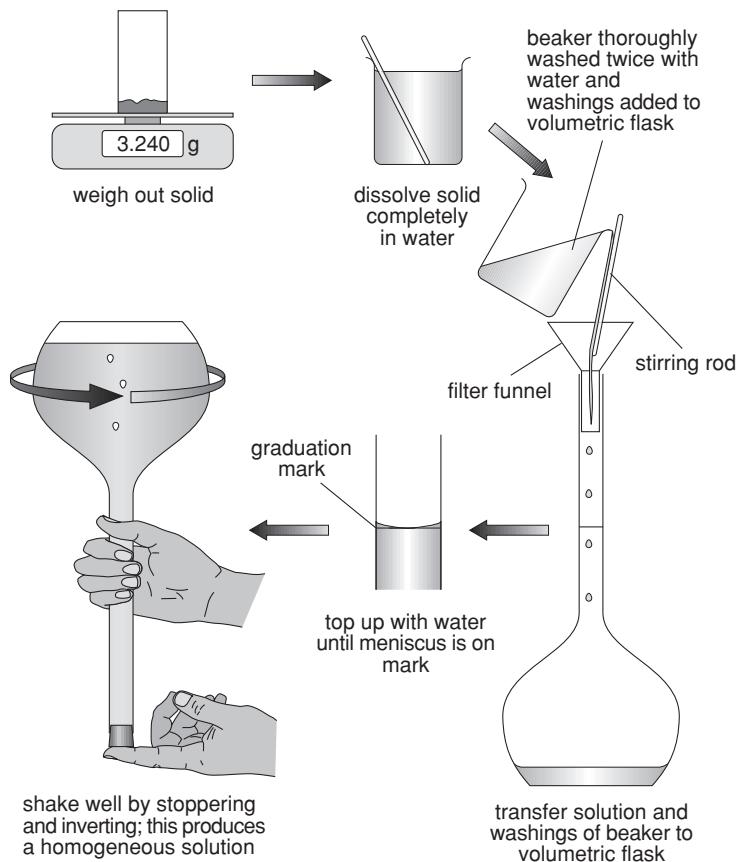
The solution of known concentration first prepared for a titration is called the **standard solution**. Preparing a standard solution in a **volumetric flask** is an important step carried out before titrations.

## Preparation of a standard solution

The chemical used to prepare a standard solution needs to be:

- a water soluble solid; it is easier to accurately measure and transfer a mass of solid than an amount of liquid or gas
- high purity; ideally **analytical reagent** (AR) grade, otherwise **laboratory reagent** (LR) grade rather than technical (TECH) grade
- of accurately known formula; this could include water of hydration such as in oxalic acid crystals  $(COOH)_2 \cdot 2H_2O$
- of stable mass on exposure to air; the solid must not lose or gain water, or react with oxygen or carbon dioxide.

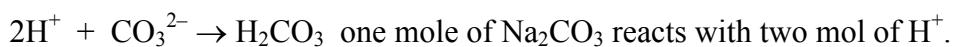
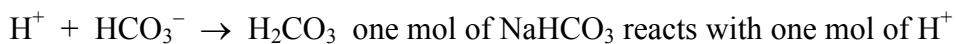
The solution is made up in a volumetric flask to a fixed total volume of solution. Water of the highest available purity such as distilled or de-ionised or demineralised water should be used.



<b>Preparation of a standard solution</b>	<b>Example</b>
<b>Procedure steps</b>	<b>250 mL of 0.100 M oxalic acid</b>
1 Calculate mol required = molarity x volume in L	$0.100 \text{ mol L}^{-1} \times 0.250\text{L} = 0.0250$
2 Calculate required mass mol x gmol $^{-1}$	$(\text{COOH})_2 \cdot 2\text{H}_2\text{O} = 126 \text{ g mol}^{-1}$ $0.0250 \times 126 = 3.15 \text{ g}$
3 Weigh accurately a mass close to the required mass in a beaker	eg. 3.24 g
4 Dissolve all of the measured mass in water, transfer to volumetric flask, wash inside beaker twice with water and transfer washings to volumetric flask.  It is essential that all of the weighed mass enters the volumetric flask.	250 mL volumetric flask used
5 Dilute the solution to the exact volume	mark on neck of volumetric flask
6 Stopper the flask, place a finger on stopper and the other hand on neck of flask. Invert flask and rotate to mix.	
7 Label: solution name, date, your name	oxalic acid, 15/1/01, Joe Blow
8 Calculate concentration	$(3.24/3.15) \times 0.100 = 0.103 \text{ M}$
9 Add concentration to the label	0.103 M

This oxalic acid solution of accurately known concentration is suitable for titration with strong bases of unknown concentration. Oxalic acid is a weak acid so phenolphthalein is usually used as indicator.

Sodium hydrogen carbonate  $\text{NaHCO}_3$  or anhydrous sodium carbonate  $\text{Na}_2\text{CO}_3$  are suitable for preparing standard solutions that can be titrated with strong acid solutions.



Methyl orange is a suitable indicator for these weak base-strong acid titrations.

# Calculations

## 1 Dilution of a solution using $cV = \text{constant}$

Suppose you wished to dilute some of a 0.103 M solution of oxalic acid to produce 500 mL of 0.0200 M solution.

### a) Calculation method explaining reasoning

The moles of oxalic acid required in the diluted solution is  $cV$  where  $c$  is molar concentration and  $V$  is volume in L.

$$cV = 0.0200 \text{ molL}^{-1} \times 0.500 \text{ L} = 0.0100 \text{ mol}$$

This no. of moles must be in the 0.103 M solution that is diluted

$$0.0100 \text{ mol} = 0.103 \text{ molL}^{-1} \times V \text{ measured in L}$$

$$V = 0.0100/0.103 = 0.0971 \text{ L} = 97.1 \text{ mL}$$

Thus water is added to 97.1 mL of 0.103 M solution to a total volume of 500 mL of solution to obtain 0.0200 M solution.

### b) Quick calculation method using $c_1V_1 = c_2V_2$

$$c_1V_1 = c_2V_2 \quad 0.103 \times V_1 = 0.0200 \times 500$$

$$V_1 = (0.0200 \times 500)/0.103 = 97.1$$

$V_1$  is in mL because  $V_2$  was in mL

97.1 mL is diluted to make the 500 mL solution.



- 1 Calculate the volume of 10 M hydrochloric acid needed to prepare 500 mL of 0.10 M hydrochloric acid solution.

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- 2 If 2.0 L of 2.0 M sodium hydroxide solution is diluted to a total volume of 10 L of solution what is the concentration of the diluted solution?

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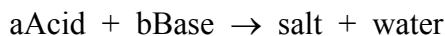
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Check your answers.

## 2. Calculating unknown concentration of a titration

solution using  $\frac{c_a V_a}{a} = \frac{c_b V_b}{b}$  for aAcid+bBase

If the balanced equation for an acid-base titration is



$$\frac{c_a V_a}{a} = \frac{c_b V_b}{b} \quad \text{where } C = \text{molar concentration}$$

V = volume (  $V_{\text{acid}}$  and  $V_{\text{base}}$  must be in same units eg both mL or both L)

a and b are the coefficients for acid and base in the balanced equation.

Note that it is essential that you have the correct formula for acid, base, salt and water and that the equation is correctly balanced before you can use this method.

In a titration  $V_{\text{acid}}$  and  $V_{\text{base}}$  are measured. An unknown concentration can be calculated if the other concentration – the standard solution concentration – is known.

Example: 25.0 mL of 0.0100 M barium hydroxide is neutralised by 34.3 mL of an unknown concentration hydrochloric acid solution.

*correct formulas:*  $\text{HCl} + \text{Ba}(\text{OH})_2 \rightarrow \text{BaCl}_2 + \text{H}_2\text{O}$

*balanced equation:*  $2\text{HCl} + \text{Ba}(\text{OH})_2 \rightarrow \text{BaCl}_2 + 2\text{H}_2\text{O}$

$$\frac{c_a \times 34.3}{2} = \frac{0.0100 \times 25.0}{1}$$

$$c_a = (0.0100 \times 25.0 \times 2) / 34.3 = 0.0146 \text{ M hydrochloric acid.}$$



25.0 mL of 0.124 M hydrochloric acid is neutralised by 0.0500 M sodium hydroxide solution in a titration. What volume of NaOH(aq) was used?

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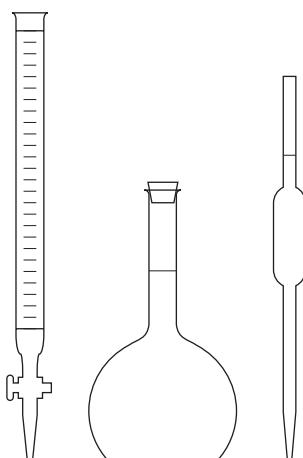
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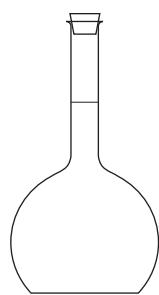
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Check your answer.

# Care and use of titration equipment



- eye level with the meniscus before reading
- equipment held or clamped vertically; line up burette with two corners of the room
- after use glassware should be emptied of solution and rinsed with water
- Note: concentrated basic solutions can dissolve glass; they should not be left in this equipment as they will alter internal volume of the volumetric glassware



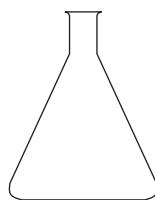
- wash with water before use
- all solid dissolved before adding last lot of water
- last amount of water added drop by drop
- stopper, invert and shake well
- label; solution name, concentration, date prepared, your name



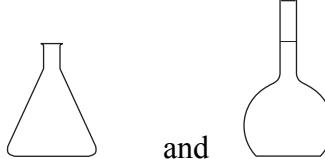
- rinse all of the internal volume with about 5 mL of solution, discard this washing
- do not allow solution into pipette bulb
- when filled hold vertically and check
  - bottom of meniscus sits on mark
  - no air bubbles near tip
  - no drops on outside of tip
- drain liquid into conical flask and hold tip against flask for 5 s after draining (last drops NOT shaken or blown out)

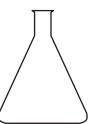


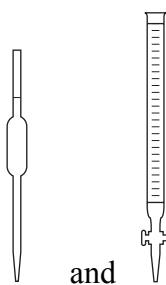
- rinse all of the internal volume with about 10 mL of solution while holding burette almost horizontal; drain the rinsing solution through the tap
- fill burette, open tap fully for 1-2 seconds to ensure no air bubbles are trapped below the tap
- if a funnel was used for filling remove funnel before taking a reading
- make sure burette is vertical by lining it up with two different corners of the room
- make sure all liquid leaving burette tip goes into the conical flask without touching neck or spraying out
- swirl the conical flask to mix contents while titrating

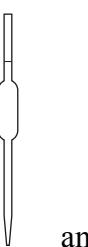
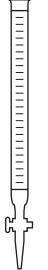


- wash with water
- at the end point contents are salt solution – only need to empty and rinse with water before using the conical flask for another titration



REMEMBER: Before using  and  wash inside



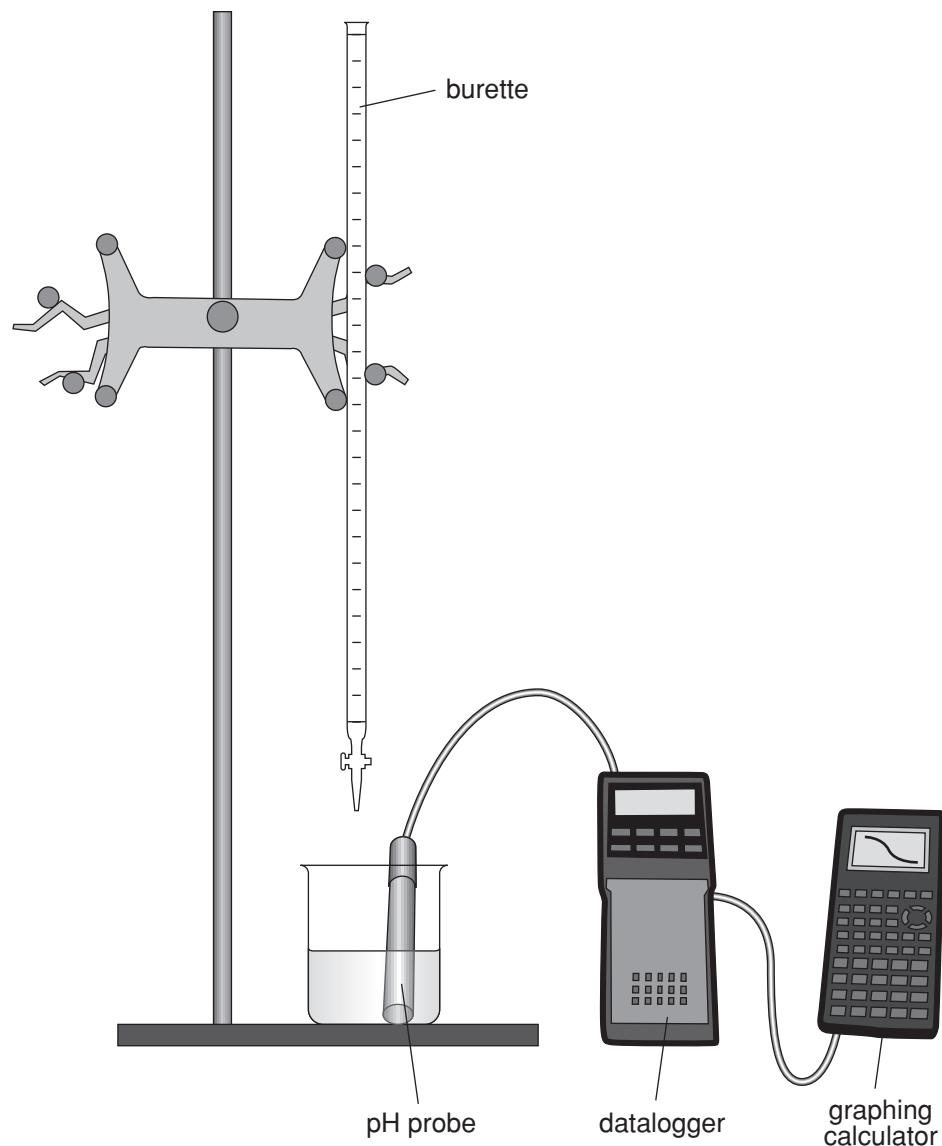
REMEMBER: Before using  and  wash inside with some of the filling solution, drain rinsings before filling.

Chemists spend a lot of time minimising errors and maintaining the accuracy of their titration equipment. All glassware should be emptied and washed with tap water at the end of the titrations. Ideally any water used for making solutions should be distilled. If not then de-ionised water is better than tap water. Tap water quality varies from location to location. Use tap water only if distilled or de-ionised is not available.

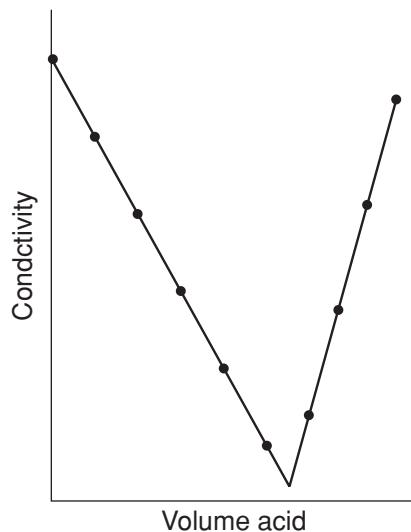
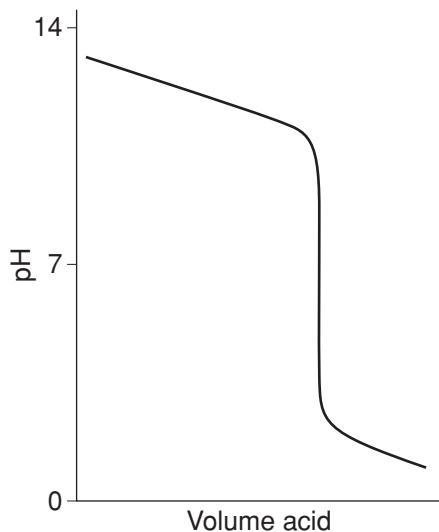
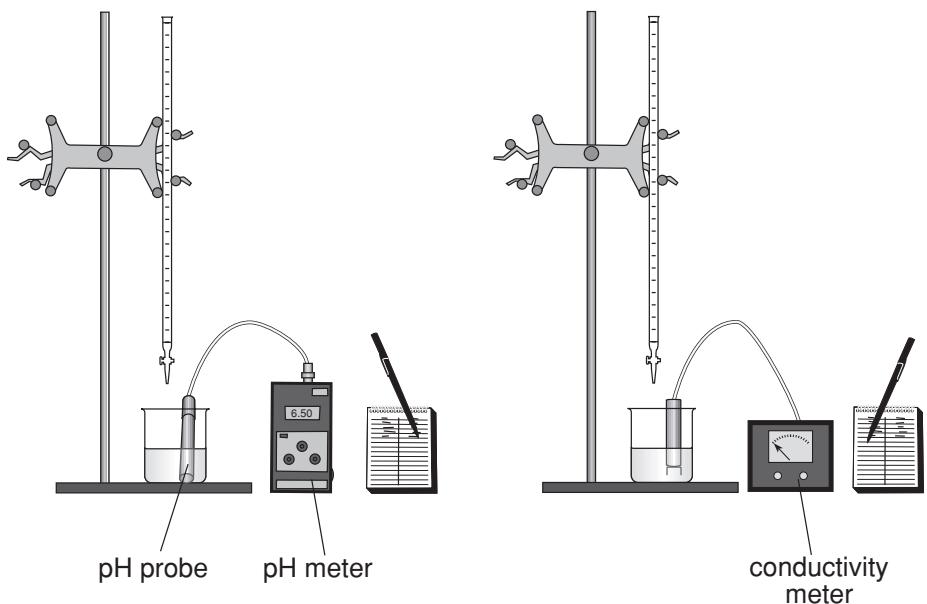
## Using a datalogger to measure the equivalence point

A datalogger stores data input electronically. The data can be displayed on a graphics screen or transferred to a computer.

Probes are available for attachment to many dataloggers to measure pH and electrical conductivity. The pH changes rapidly at equivalence while a graph of the electrical conductivity changes slope at equivalence.



If a datalogger is not available the measurements made using a pH meter or conductivity meter can be logged by recording the measurements in table format using pen and paper.



Note: this conductivity graph shape is obtained for a titration where the salt formed is water insoluble

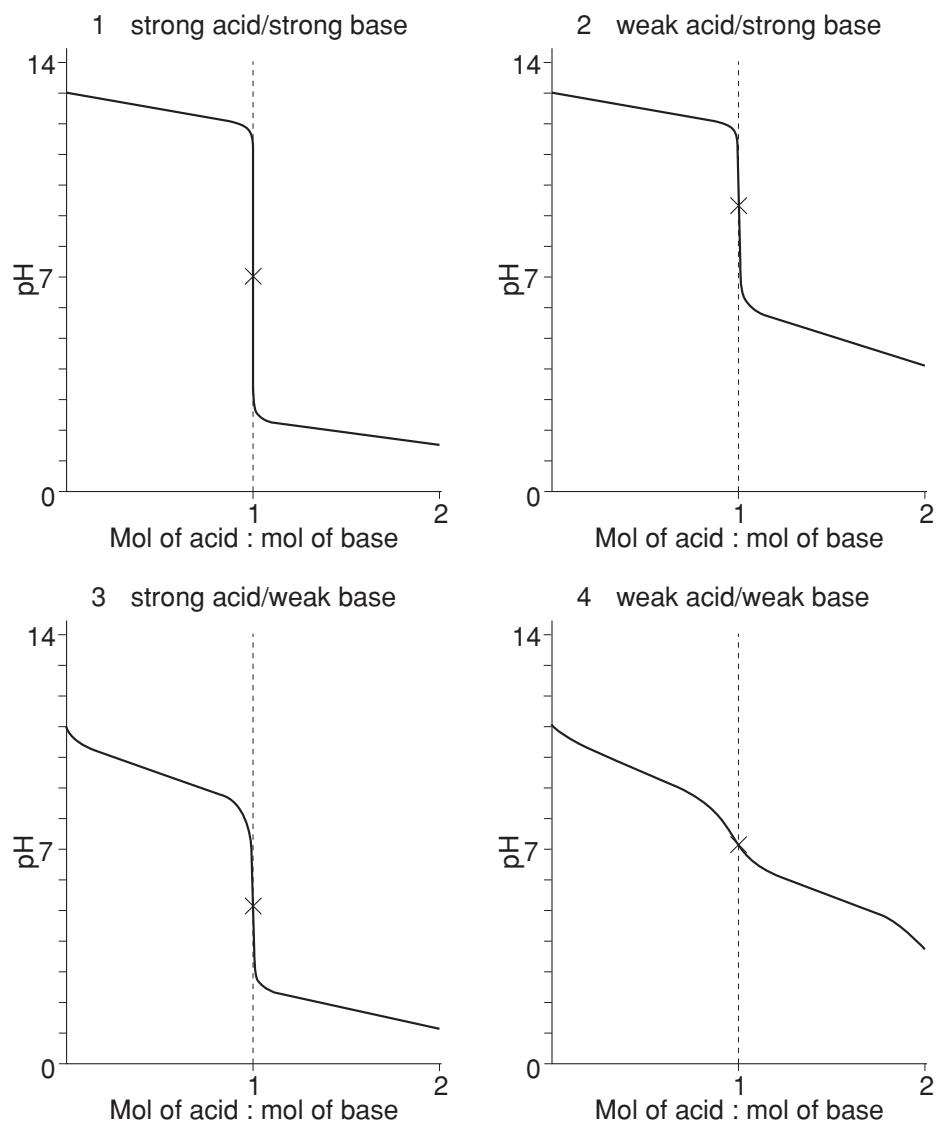
A magnetic stirrer is normally used to mix the acid and base solutions while the titration is carried out. A plastic coated magnet stirrer is placed in the beaker or conical flask.

The beaker or conical flask is placed on a platform above an electric motor. As the electric motor rotates a magnetic arm the magnetic stirrer in the beaker or conical flask rotates mixing the acid and base solutions.

# Titration pH graphs

The graphs below show changes in pH for various combinations of acid and base. All the examples shown are for:

- a 1:1 ratio of acid and base (one mole of acid neutralises one mole of base)
- acid in the burette added to base in the receiving flask ( the pH probe was in the receiving flask solution above a magnetic stirrer)
- all solutions titrated were 0.10 M concentration.



The slight change in pH at the equivalence point for the weak acid-weak base titration makes it very difficult to determine the equivalence point. If a weak acid is being titrated a strong base must be used. If a weak base is being titrated a strong acid must be used. This way the pH changes sufficiently for the measured end point to be detected by a rapid change in indicator colour or pH.



- 1 Which of the graphs on the previous page could represent a titration between:
  - a)  $\text{NaOH}(\text{aq})$  and  $\text{HCl}(\text{aq})$   
\_\_\_\_\_
  - b)  $\text{NaHCO}_3(\text{aq})$  and  $\text{CH}_3\text{COOH}$   
\_\_\_\_\_
  - c)  $\text{KOH}(\text{aq})$  and  $\text{CH}_3\text{COOH}(\text{aq})$   
\_\_\_\_\_
  - d)  $\text{NaHCO}_3(\text{aq})$  and  $\text{HNO}_3(\text{aq})$ ?  
\_\_\_\_\_

- 2 Using information from the graphs only (not your knowledge of which acids and which bases are strong and weak) justify your answers.

a) \_\_\_\_\_  
\_\_\_\_\_

b) \_\_\_\_\_  
\_\_\_\_\_

c) \_\_\_\_\_  
\_\_\_\_\_

d) \_\_\_\_\_  
\_\_\_\_\_

Check your answers.

## Determining the concentration of acetic acid in vinegar by titration with sodium hydroxide solution



This activity can be done:

- EITHER in a laboratory using volumetric flasks, pipette, burette, conical flask and phenolphthalein indicator
- OR at home using plastic storage bottles from supermarket products, 3 mL syringe as pipette, 5 to 10 mL syringe as burette, a small transparent glass or plastic container as the receiving flask and red cabbage solution indicator.

The principles are the same and techniques used similar. All measurements in the laboratory should be made to three significant figures (eg. 15.8 mL) while all measurements at home should be made to two significant figures (eg. 3.6 mL).

### What you will need:

- two volumetric flasks (250 mL and 500 mL) OR two plastic storage bottles from supermarket used to store acidic or basic liquids (mark the liquid levels in each plastic bottle where it holds 250 mL/250 g and 500 mL/500 g of water)
- 25 mL bulb pipette OR 3 mL syringe (without needle) from a pharmacy
- 50 mL burette and stand and clamp for holding the burette OR 5 to 10 mL syringe (without needle) from a pharmacy
- 250 mL conical flask OR a small transparent glass or plastic container such as a 40 mL conical medicine measure as receiving flask
- three small beakers (50 – 250 mL) or three small open transparent glass or plastic containers for holding solutions
- phenolphthalein indicator solution (can be made using laxatives containing phenolphthalein: crush laxative, swirl with methylated spirits, decant liquid and add an equal volume of water) OR red cabbage solution
- white vinegar
- sodium hydroxide OR a spray can of caustic oven cleaner that contains sodium hydroxide from a supermarket and a plastic bucket
- safety goggles or glasses
- 3 intact aspirin (acetylsalicylic acid) tablets; do not use soluble aspirin tablets which contain acetylsalicylate salts.



### What you will do:

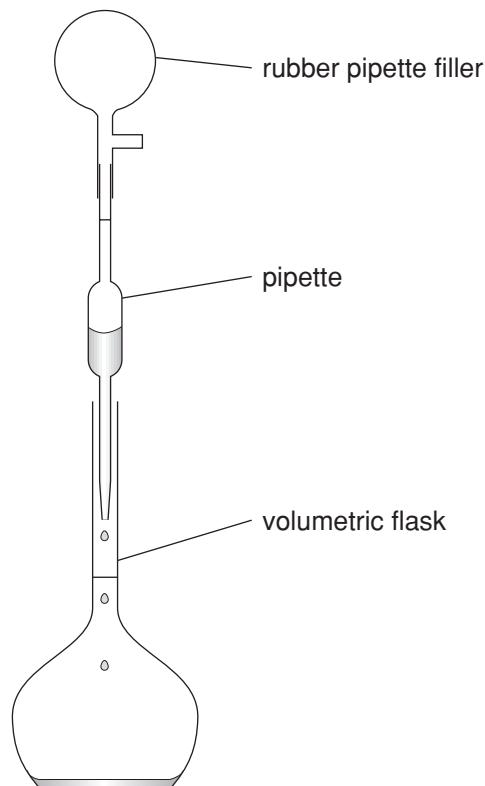
Wear eye protection throughout this activity involving sodium hydroxide.

#### A: Prepare an approximately 0.1 M sodium hydroxide solution

If you have solid sodium hydroxide then dissolve between one and two cubic centimetres volume of the solid in 400 mL of water.

OR

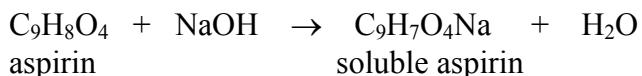
If you have a spray can of caustic oven cleaner that contains sodium hydroxide read the instructions carefully. Place about 400 mL of water in the plastic bucket. Spray four one second bursts of NaOH containing oven cleaner towards the water.



Make the solution up to a total volume of 500 mL. Store in a 500 mL volumetric flask or a plastic or glass (but not metal) container clearly labelled 0.1M sodium hydroxide and with your name and the date.

## B: Standardise the sodium hydroxide by titration against a standard

The standard you will use are aspirin tablets. Make sure the tablets are the less common type that contain molecular acetylsalicylic acid rather than the more common soluble aspirin tablets that contain ionic salts such as sodium acetylsalicylate. The tablets normally contain 300 mg of C<sub>9</sub>H<sub>8</sub>O<sub>4</sub> which has a formula mass of 180. The known amount of acid in a tablet can be titrated against the 0.1 M NaOH:



1 mol + 1 mol

$0.300/180 = 0.00167$  mol aspirin is neutralised by 0.00167 mol NaOH.

The tablet is broken up inside a container (eg. 100 mL beaker) containing five to ten mL of water. Add:

- Lab: 2 drops phenolphthalein solution or 1 mL red cabbage solution  
OR
- Home: 1 drop phenolphthalein solution or 0.5 mL red cabbage solution before the titration is commenced.

Don't forget to record the burette/syringe reading before commencing to add the NaOH solution as well as when the titration is complete.

Record the volume of NaOH solution added from a burette or syringe needed to change either the colourless phenolphthalein to a pink which remains 30 s after swirling or to change the red cabbage solution to a green that remains 30 s after swirling. The first titration is 'rough' giving you an idea of how much is required to reach the end point. The titration should be repeated obtaining two similar volumes (**titres**).

Titration	Initial volume (mL)	Final volume (mL)	Volume added (mL)
'rough'			
titration 2			
titration 3			

Use the average of two similar titres for your calculation.

Concentration of NaOH(aq) = 0.00167 mol/ average titre in L.

Calculate the concentration to three significant figures if using a burette eg. 0.103 M or to two significant figures if using a syringe eg 0.10 M.

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### C: Titrate standardised NaOH(aq) with diluted vinegar

When a titration is carried out the volume of acid and base used to reach the end point should be about the same. Vinegar can be up to 1 M acetic acid and will need to be diluted tenfold to 10% v/v vinegar before titrating.

Use a 25 mL bulb pipette or a 3 mL syringe to add 25 mL of vinegar to a 250 mL volumetric flask half filled with water. Add water to a total volume of 250 mL then label the container 10% v/v white vinegar with your name and the date.

Rinse the bulb pipette or syringe at least twice with the 10% vinegar solution being careful not to contaminate the bulk solution. Transfer either 25 mL (pipette used) or 2.5 mL (syringe used) to a conical or receiving flask.

Always add indicator solution before the titration is commenced. Don't forget to record the burette/syringe reading before commencing to add the NaOH solution as well as when the titration is complete.

Record the volume of NaOH solution added from a burette or syringe needed to change the colourless phenolphthalein to a pink which remains 30 s after swirling or the red cabbage solution to a green that remains 30 s after swirling.

Titration	Initial volume (mL)	Final volume (mL)	Volume added (mL)
'rough'			
titration 2			
titration 3			

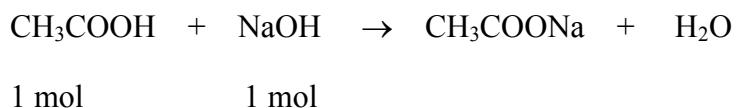
Use the average of two similar titres for your calculation.

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average titre =

Now use the average titre to calculate the concentration of the diluted vinegar and the undiluted vinegar.



Calculate the molar concentration of diluted vinegar using

$$\frac{c_a v_a}{a} = \frac{c_b v_b}{b}$$

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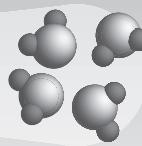
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Concentration of vinegar = 10 x concentration of diluted vinegar =

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**MACRO**observe  
infer  
understand**SYMBOLIC** $H_2O$  formulas  
equations  
calculations**MICRO**particles  
energy  
interactions

# Heat of neutralisation

When neutralisation occurs a proton  $H^+$  is transferred from acid to base and this H nucleus covalently bonds to another atom. Formation of the covalent bond releases energy so neutralisation is an exothermic reaction.

## Measuring and calculating molar heat of neutralisation

**What you will need:**

- 100 mL of white vinegar (undiluted) of known concentration calculated in the previous activity
- sodium hydroxide (pellets, flakes or granules)
- takeaway polystyrene cup and plastic lid with hole for a straw
- 0–110°C thermometer
- container for measuring out 100 mL of liquid



Wear eye protection throughout this activity involving sodium hydroxide.

**What you will do:**

- 1 Prepare 100 mL of NaOH solution of the same molar concentration as the white vinegar.

eg. If the vinegar was 0.50M acetic acid you need 100 mL of 0.50 M NaOH:  $\text{mol required} = \text{mol L}^{-1} \times \text{L} = 0.50 \times 0.100 = 0.050$   
 $0.050 \text{ mol} = 0.050 \text{ mol} \times (23 + 16 + 1) \text{ g mol}^{-1} = 0.050 \times 40 = 2.0 \text{ g}$

One  $\text{cm}^3$  (mL) volume of solid NaOH (pellets, flakes or granules) has a mass of about one gram. If you do not have access to a weighing balance measuring to at least 0.1 g then estimate  $x$  grams by measuring out  $x \text{ cm}^3$  ( $x$  mL) of the solid. Do this quickly as NaOH absorbs water and reacts with  $\text{CO}_2$  in the air.

Stir the NaOH with the 100 mL of water until it is all dissolved.  
Allow the solution to cool back to room temperature.

- 2 Place the 100 mL of white vinegar into the cup. Measure its temperature with the thermometer. \_\_\_\_\_
- 3 Measure the temperature of the 100 mL of sodium hydroxide solution. \_\_\_\_\_
- 4 Calculate the average temperature of the two solutions \_\_\_\_\_
- 5 Pour the 100 mL of sodium hydroxide solution into the 100 mL of white vinegar. Place the lid on the cup and insert the thermometer bulb to the centre of the liquid. Measure the maximum temperature reached by the thermometer. \_\_\_\_\_

Temperature increase = maximum temperature – average temperature



### **Calculations:**

Assume the 100 mL white vinegar and 100 mL sodium hydroxide solution which are mostly water have the same mass and specific heat as 100 mL of water.

$$\begin{aligned}\text{Heat released} &= \frac{\text{mass of water}}{\text{g}} \times 4.18 \times \frac{\text{temperature increase}}{\text{C}^{\circ}\text{-1}} \\ &= 200 \quad \times 4.18 \times \underline{\hspace{2cm}} \\ &= \underline{\hspace{2cm}} \text{ J}\end{aligned}$$

Molarity of acetic acid solution =  $\text{mol L}^{-1}$

Moles of acetic acid in 100 mL white vinegar =  $\underline{\hspace{2cm}}$

Heat released by  $\underline{\hspace{2cm}}$  mol acetic acid

Heat of reaction =  $\underline{\hspace{2cm}} \text{ J mol}^{-1}$  =  $\underline{\hspace{2cm}} \text{ kJ mol}^{-1}$

Write a balanced equation for the reaction.

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### **Conclusions:**

How does your result compare with published molar heats of neutralisation of about  $-57 \text{ kJ mol}^{-1}$ ?

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Suggest reasons why your result is different from  $-57 \text{ kJ mol}^{-1}$ .

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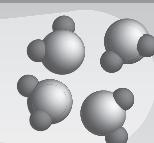
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Suggest ways of improving the accuracy of this activity.

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**MACRO**observe  
infer  
understand**SYMBOLIC** $H_2O$  formulas  
equations  
calculations**MICRO**particles  
energy  
interactions

# Buffers

Buffers control the level of acidity of a solution. A buffer is usually a mixture of a weak acid and its conjugate base. The mixture uses an equilibrium reaction to maintain the hydrogen ion concentration at a nearly constant level even when large amounts of strong acid or strong base are added.

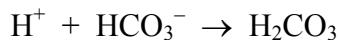
Buffer solutions are important in maintaining the pH of swimming pool water, sea water and blood. Important buffer solutions are shown in ***bold italics*** in the list below:

$HCl$	strong	$Cl^-$	extremely weak
$H_2SO_4$	strong	$HSO_4^-$	extremely weak
$HNO_3$	strong	$NO_3^-$	extremely weak
$H_3O^+$		$H_2O$	
$HSO_4^-$		$SO_4^{2-}$	
$H_2SO_3$		$HSO_3^-$	
$H_3PO_4$		$H_2PO_4^-$	
$CH_3COOH$		$CH_3COO^-$	
$H_2CO_3$		$HCO_3^-$	
$H_2S$		$HS^-$	
$HSO_3^-$		$SO_3^{2-}$	
$H_2PO_4^-$		$HPO_4^{2-}$	
$NH_4^+$		$NH_3$	
$HCO_3^-$		$CO_3^{2-}$	
$HPO_4^{2-}$		$PO_4^{3-}$	
$H_2O$	extremely weak	$OH^-$	strong
$HS^-$	extremely weak	$S^{2-}$	strong
$OH^-$	extremely weak	$O^{2-}$	strong

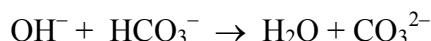
Note that the buffer solutions listed are all made up of a weak acid and its weak conjugate base.

The pH of sea water and blood are kept constant by equilibria involving  $\text{H}_2\text{CO}_3$ ,  $\text{HCO}_3^-$  and  $\text{CO}_3^{2-}$

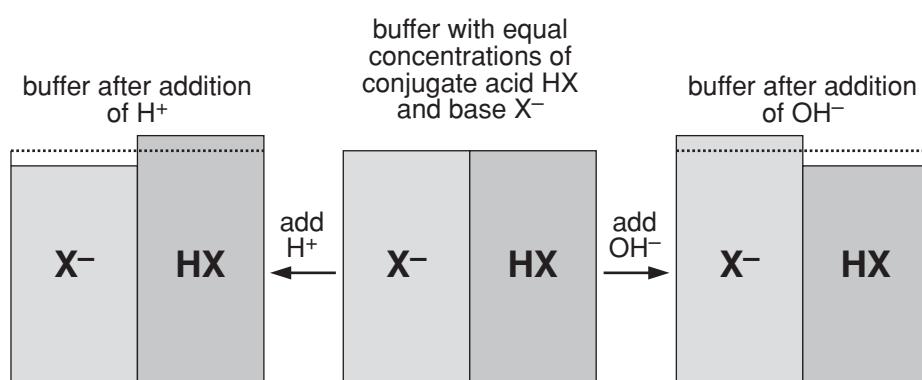
- added acid is removed by the reaction



- added base is removed by the reaction



A buffer solution that maintains pH about 4.5 is made up of acetic acid and acetate salt. The way in which it minimises change in pH is demonstrated in the diagram below.



Because HX is a weak acid, an increase or decrease in its amount produces only a very small change in the concentration of hydrogen ions. The pH hardly changes.



Another buffer solution is made up from an ammonium salt and ammonia solution.  $\text{NH}_4^+ \rightleftharpoons \text{NH}_3 + \text{H}^+$

- What is the acid and what is the conjugate base in this buffer solution?

- 
- Where does the equilibrium position shift to

a) when acid is added to this buffer solution? \_\_\_\_\_

b) when base is added to this buffer solution? \_\_\_\_\_

- Give an equation showing the reaction that occurs

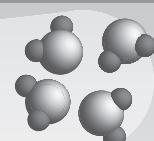
a) when acid is added to this buffer solution

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b) when base is added to this buffer solution.

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Check your answers.

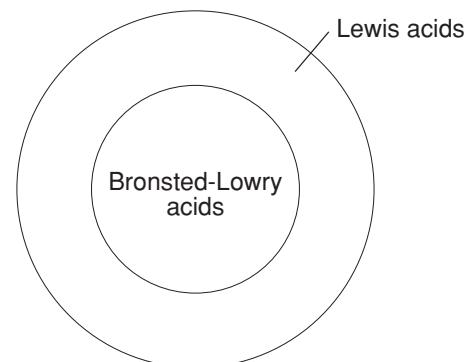
**MACRO**observe  
infer  
understand**SYMBOLIC** $H_2O$  formulas  
equations  
calculations**MICRO**particles  
energy  
interactions

# Lewis acid-base definition

The Lewis concepts of acid and base was developed about the same time as the Bronsted-Lowry concept but highlight the role of the electron pair.

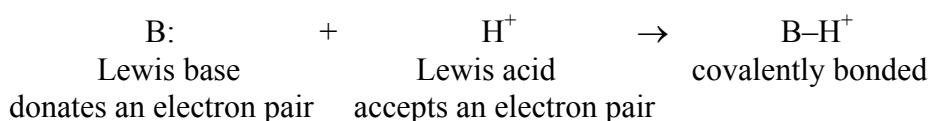
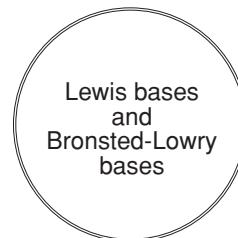
Lewis defined an acid as any species that accepts an electron pair.

The electron pair is used to form a covalent bond between the Lewis acid and the Lewis base. Species that do not contain hydrogen and therefore cannot be Bronsted-Lowry acids (proton donors) can be defined as Lewis acids eg  $CO_2$  and metal cations. Because all Bronsted-Lowry acids donate a proton which can accept an electron pair all Bronsted-Lowry acids are Lewis acids. Lewis's view expanded the concept of an acid to include many species beyond these proton donors.



A Lewis base is any species that donates an electron pair.

When a Bronsted-Lowry base accepts a proton it attaches the hydrogen using a donated electron pair to form a covalent bond. Lewis bases and Bronsted-Lowry bases are the same substances even though the definitions are different in how they describe what is happening.





The table below shows examples of Lewis acids. Tick the column headed Bronsted-Lowry acid if the example satisfies the Bronsted-Lowry definition of an acid; cross the column if the example doesn't. The large [ ] brackets indicate that the charge shown outside is distributed over the whole of the species.

Lewis acid-base reaction	Bronsted-Lowry acid?
$\text{MgO(s)} + \text{CO}_2\text{(g)} \rightarrow \text{MgCO}_3\text{(s)}$  $\text{Mg}^{2+} \cdot \ddot{\text{O}} \cdot \ddot{\text{O}}^{2-} + \ddot{\text{O}} = \ddot{\text{C}} = \ddot{\text{O}} \cdot \rightarrow \text{Mg}^{2+} \left[ \begin{array}{c} \ddot{\text{O}} \\   \\ \text{C} \\    \\ \ddot{\text{O}} \end{array} \right]^{2-}$	
$\text{H}_3\text{O}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow 2\text{H}_2\text{O(l)}$  $\left[ \begin{array}{c} \ddot{\text{O}} \\   \\ \text{H} & \text{H} & \text{H} \end{array} \right]^+ + \left[ \begin{array}{c} \ddot{\text{O}} \\   \\ \text{H} \end{array} \right]^- \rightarrow \text{H}_2\text{O} + \text{H}_2\text{O}$	
$\text{Cu}^{2+}(\text{aq}) + 4\text{H}_2\text{O(l)} \rightarrow \text{Cu}(\text{OH}_2)^{2+}(\text{aq})$  $\text{Cu}^{2+} + \left[ \begin{array}{c} \text{H} \\   \\ \ddot{\text{O}} \\   \\ \text{H} \\   \\ \ddot{\text{O}} \\   \\ \text{H} \\   \\ \ddot{\text{O}} \\   \\ \text{H} \end{array} \right] \rightarrow \left[ \begin{array}{c} \text{H} \\   \\ \ddot{\text{O}} \\   \\ \text{H} \\   \\ \ddot{\text{O}} \\   \\ \text{H} \\   \\ \ddot{\text{O}} \\   \\ \text{H} \end{array} \right]^{2+}$	

Check your answers.



Assess the importance of the Bronsted-Lowry definitions and Lewis definitions in understanding:

- a) what is happening during an acid-base titration

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- b) biologically active molecules requiring metal ions to function.

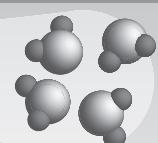
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Check your answers.



Now complete Exercises 4.1, 4.2 and 4.3.

**MACRO**observe  
infer  
understand**SYMBOLIC** $\text{H}_2\text{O}$  formulas  
equations  
calculations**MICRO**particles  
energy  
interactions

# Suggested answers

## Conjugate acid-base pairs

- 1    a)  $\text{NO}_3^-$    b)  $\text{HSO}_4^-$    c)  $\text{CO}_3^{2-}$    d)  $\text{SO}_4^{2-}$
- 2    a)  $\text{HS}^-$    b)  $\text{H}_2\text{S}$    c)  $\text{HCO}_3^-$    d)  $\text{HSO}_4^-$
- 3    The acid is listed first then the base for each conjugate pair:  
 b)  $\text{H}_2\text{O}$  and  $\text{OH}^-$ ,  $\text{HBr}$  and  $\text{Br}^-$   
 c)  $\text{H}_2\text{O}$  and  $\text{OH}^-$ ,  $\text{HNO}_3$  and  $\text{NO}_3^-$   
 d)  $\text{NH}_4^+$  and  $\text{NH}_3$ ,  $\text{HBr}$  and  $\text{Br}^-$   
 e)  $\text{H}_2\text{O}$  and  $\text{OH}^-$ ,  $\text{HSO}_4^-$  and  $\text{SO}_4^{2-}$   
 f)  $\text{H}_2\text{O}$  and  $\text{OH}^-$ ,  $\text{H}_2\text{S}$  and  $\text{HS}^-$

## Explaining acidic and basic salt solutions

- 1     $\text{NH}_4^+ + \text{H}_2\text{O} \rightarrow \text{NH}_3 + \text{H}_3\text{O}^+$
- 2     $\text{HCO}_3^- + \text{H}_2\text{O} \rightarrow \text{H}_2\text{CO}_3 + \text{OH}^-$
- 3     $\text{H}_2\text{PO}_4^- + \text{H}_2\text{O} \rightarrow \text{HPO}_4^{2-} + \text{H}_3\text{O}^+$
- 4     $\text{CH}_3\text{COO}^- + \text{H}_2\text{O} \rightarrow \text{CH}_3\text{COOH} + \text{OH}^-$

## Amphiprotic or amphoteric?

- a)  $\text{HSO}_4^- + \text{H}^+ \rightarrow \text{H}_2\text{SO}_4$
- b)  $\text{HSO}_4^- + \text{OH}^- \rightarrow \text{SO}_4^{2-} + \text{H}_2\text{O}$
- 1    Amphiprotic is a conceptual definition applied to species such as molecules or ions which can donate or accept a proton. Amphoteric is an operational definition applied to substances which are observed to react with an acid and to also react with a base.
- 2    The hydrogen phosphate ion  $\text{HPO}_4^-$  is amphiprotic because it can act as a proton donor (acid):  $\text{HPO}_4^- + \text{H}_2\text{O} \rightarrow \text{PO}_4^{3-} + \text{H}_3\text{O}^+$   
 and as a proton acceptor (base):  
 $\text{HPO}_4^- + \text{H}_2\text{O} \rightarrow \text{H}_2\text{PO}_4^{2-} + \text{OH}^-$

The second reaction occurs to the greatest extent so solutions of hydrogen phosphate salts are slightly basic.

- 3 Aluminium oxide  $\text{Al}_2\text{O}_3$  does not contain any H so it cannot act as a proton donor (acid) according to the Bronsted-Lowry theory and so on this definition cannot be amphiprotic.

## Neutralisation reactions and equations

- 1  $\text{NaOH(aq)} + \text{HCl(aq)} \rightarrow \text{NaCl(aq)} + \text{H}_2\text{O(l)}$
- 2  $\text{NaOH(aq)} + \text{CH}_3\text{COOH(aq)} \rightarrow \text{CH}_3\text{COONa(aq)} + \text{H}_2\text{O(l)}$
- 3  $2\text{KOH(aq)} + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{K}_2\text{SO}_4(\text{aq}) + 2\text{H}_2\text{O(l)}$
- 4  $3\text{KOH(aq)} + \text{H}_3\text{PO}_4(\text{aq}) \rightarrow \text{K}_3\text{PO}_4(\text{aq}) + 3\text{H}_2\text{O(l)}$
- 5  $\text{KOH(aq)} + \text{HNO}_3(\text{aq}) \rightarrow \text{KNO}_3(\text{aq}) + \text{H}_2\text{O(l)}$

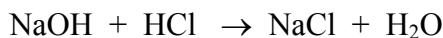
## Volumetric analysis

The equivalence point is when potassium hydroxide and phosphoric acid have completely reacted in a 3:1 mole ratio.

## Calculations

$$1 \quad c_1 V_1 = c_2 V_2 \quad 10 \times V_1 = 0.10 \times 500 \quad V_1 = (0.10 \times 500)/10 = 5.0 \text{ mL}$$

$$2 \quad c_1 V_1 = c_2 V_2 \quad 2.0 \times 2.0 = c_2 \times 10 \quad c_2 = (2.0 \times 2.0)/10 = 0.40 \text{ M}$$



$$\text{Using } \frac{c_a V_a}{a} = \frac{c_b V_b}{b}$$

$$\frac{0.124 \times 25.0}{1} = \frac{0.0500 \times V_b}{1}$$

$$V_b = (0.124 \times 25.0)/0.0500 = 62.0 \text{ mL NaOH(aq) used}$$

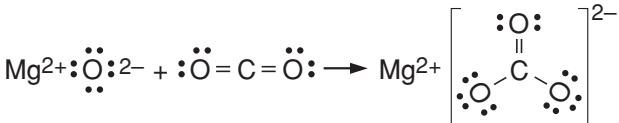
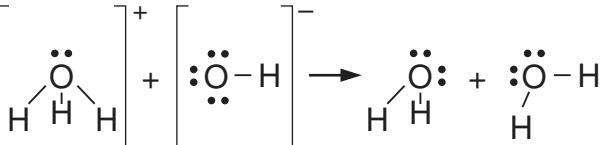
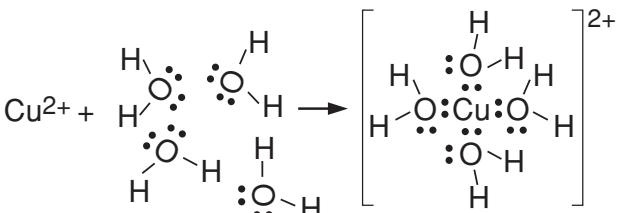
## Titration pH graphs

- 1 a) 1 b) 4 c) 2 d) 3
- 2 a) high pH start, low pH finish, large pH change around equivalence point  
b) pH 11 start, pH 3 finish, slow change in pH  
c) high pH start, pH 4 finish  
d) pH 11 start, low pH finish

## Buffers

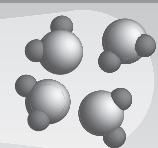
- 1  $\text{NH}_4^+$  acid,  $\text{NH}_3$  conjugate base
- 2 a) to the left  
b) to the right (because  $[\text{H}^+]$  is reduced)
- 3 a)  $\text{H}^+ + \text{NH}_3 \rightarrow \text{NH}_4^+$   
b)  $\text{NH}_4^+ \rightarrow \text{NH}_3 + \text{H}^+$  because  $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$

## Lewis acid-base definition

Lewis acid-base reaction	Bronsted-Lowry acid?
$\text{MgO(s)} + \text{CO}_2(\text{g}) \rightarrow \text{MgCO}_3(\text{s})$ 	$\times$
$\text{H}_3\text{O}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow 2\text{H}_2\text{O(l)}$ 	✓
$\text{Cu}^{2+}(\text{aq}) + 4\text{H}_2\text{O(l)} \rightarrow \text{Cu}(\text{OH}_2)^{2+}(\text{aq})$ 	$\times$

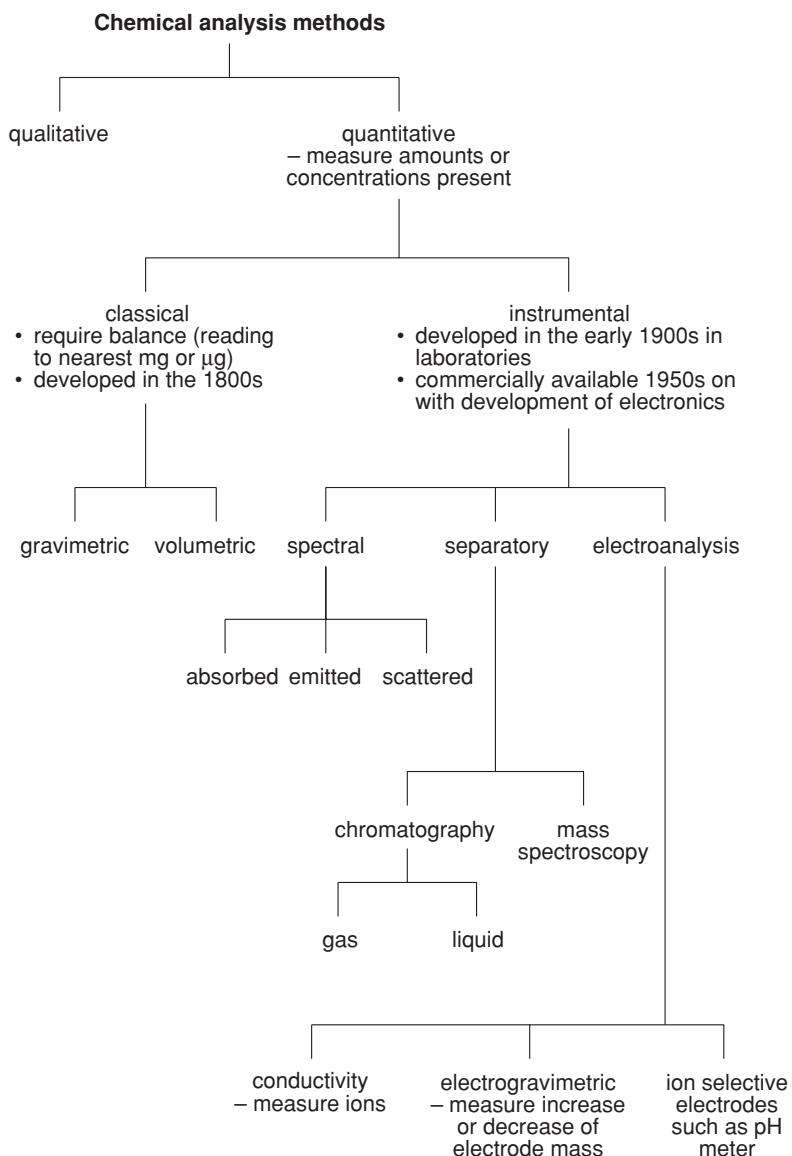
- a) Bronsted-Lowry definitions are used to explain what is happening when acid contains H than can be donated as  $\text{H}^+$  to react with a base. Focussing on the number of moles of  $\text{H}^+$  donated by each unit of acid ( eg.  $\text{HCl}$ ,  $\text{H}_2\text{SO}_4$ ) and accepted by each unit of base (eg.  $\text{KOH}$ ,  $\text{Ba}(\text{OH})_2$ ) enables quantitative calculations.
- b) Lewis definitions explain why metal ions interact as Lewis acids (electron pair acceptors) with non-bonding electron pairs (eg. on O or N atoms) in the biologically active molecule (electron pair donors).



**MACRO**observe  
infer  
understand**SYMBOLIC** $H_2O$  formulas  
equations  
calculations**MICRO**particles  
energy  
interactions**Exercises – Part 4**

Exercises 4.1 to 4.3

Name: \_\_\_\_\_

**Exercise 4.1: Chemical analysis methods**

Use the information in the previous diagram to answer the questions following.

- 1 Circle any quantitative method on the diagram you have studied so far in your course.
- 2 Label each of these quantitative methods used to accurately measure concentrations with a label from the previous page. Do not use the general labels quantitative, classical or instrumental.
  - a) A food technologist titrated lemon juice with sodium hydroxide to determine the citric acid concentration.  
\_\_\_\_\_
  - b) A swimming pool attendant used a pH meter to measure the pH of swimming pool water.  
\_\_\_\_\_
  - c) An archeologist arranged for carbon dating of a skeleton believed to be 30 000 years old.  
\_\_\_\_\_
  - d) A water supply treatment officer precipitated the sulfate ions in a sample using barium ions. She filtered off the barium sulfate precipitate, dried and weighed it then used her measurements to calculate the sulfate concentration of the water sample.  
\_\_\_\_\_
  - e) A research meteorologist used measurements of the intensity of solar radiation transmitted to the earth's surface to study growth of an ozone hole in the atmosphere over Tasmania.  
\_\_\_\_\_
  - f) A chemical engineer studying the efficiency of a process weighed electrodes before and after electrolysis.  
\_\_\_\_\_
  - g) A botanist trying to measure the amount of citric acid in a plant liquid used paper chromatography to separate the citric acid.  
\_\_\_\_\_
  - h) An automotive engineer measuring oxide pollutants emitted by a hot engine used chromatography equipment connected to the exhaust system.  
\_\_\_\_\_

## Exercise 4.2: Positioning of a burette

The diagram below shows two burettes containing solutions for a titration. If only an ordinary clamp is available for holding the burette, how can you make sure the burette is vertical?

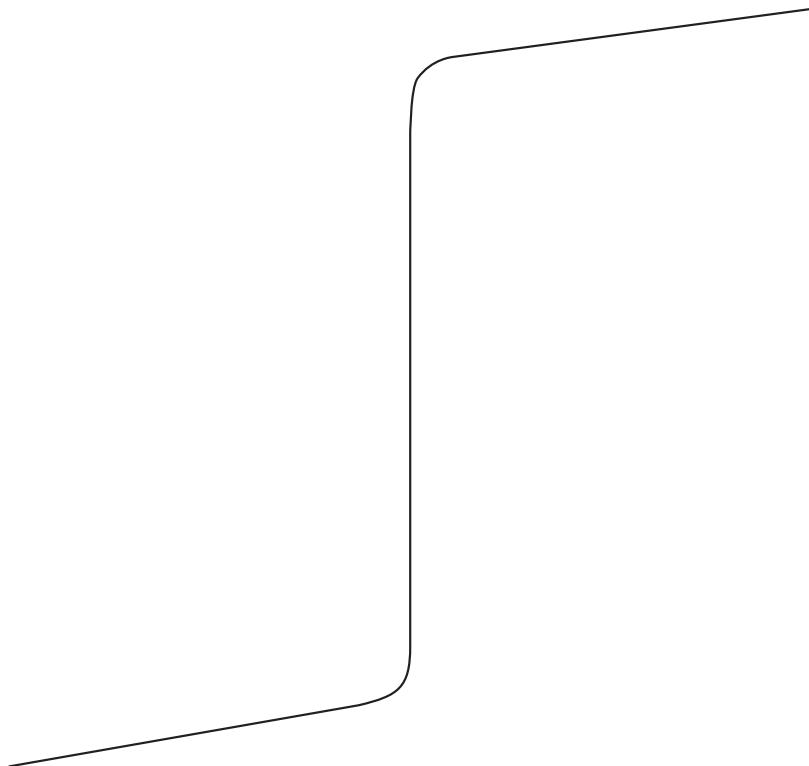
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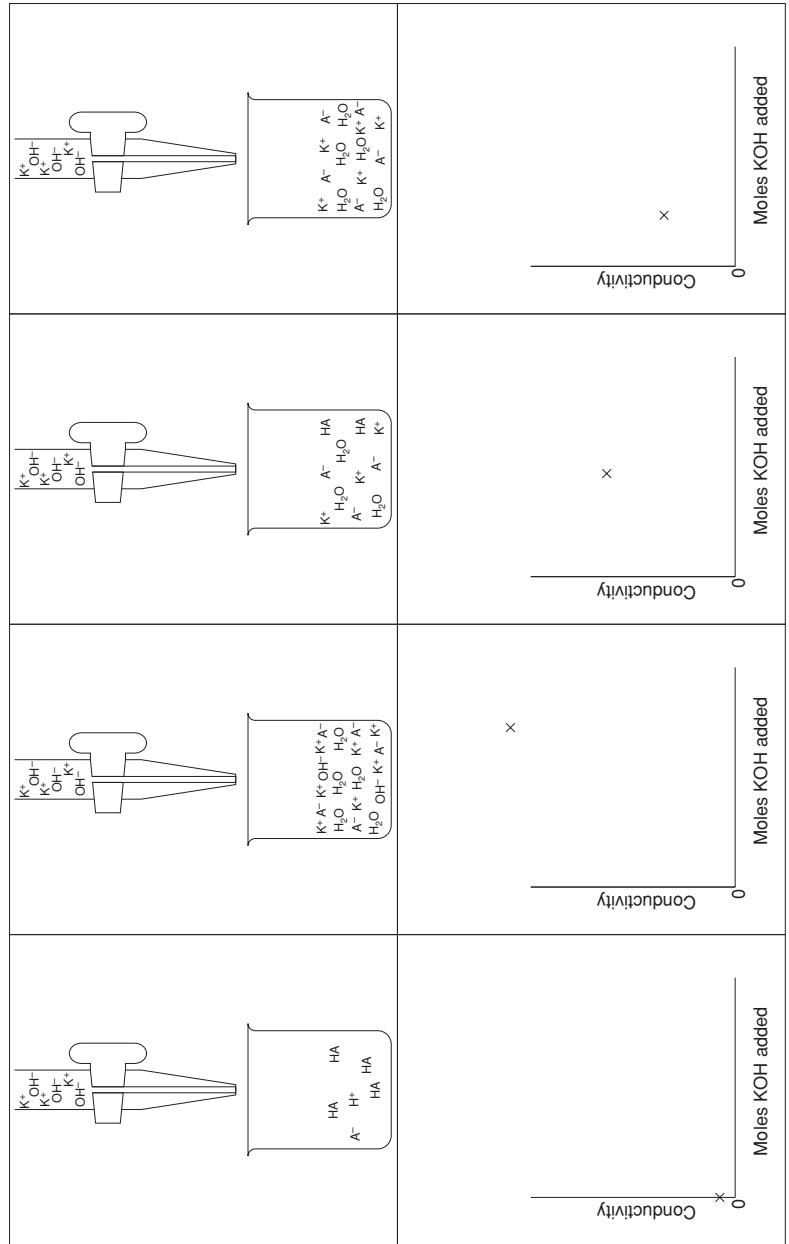
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### **Exercise 4.3: Visualising the particles during a titration**

The eight diagrams on the next page show different stages in a titration. Cut the diagrams out and arrange them at appropriate levels of the titration curve. Each of the burette and beaker diagrams matches up with one of the graph diagrams. Add axes to the diagram and label them.

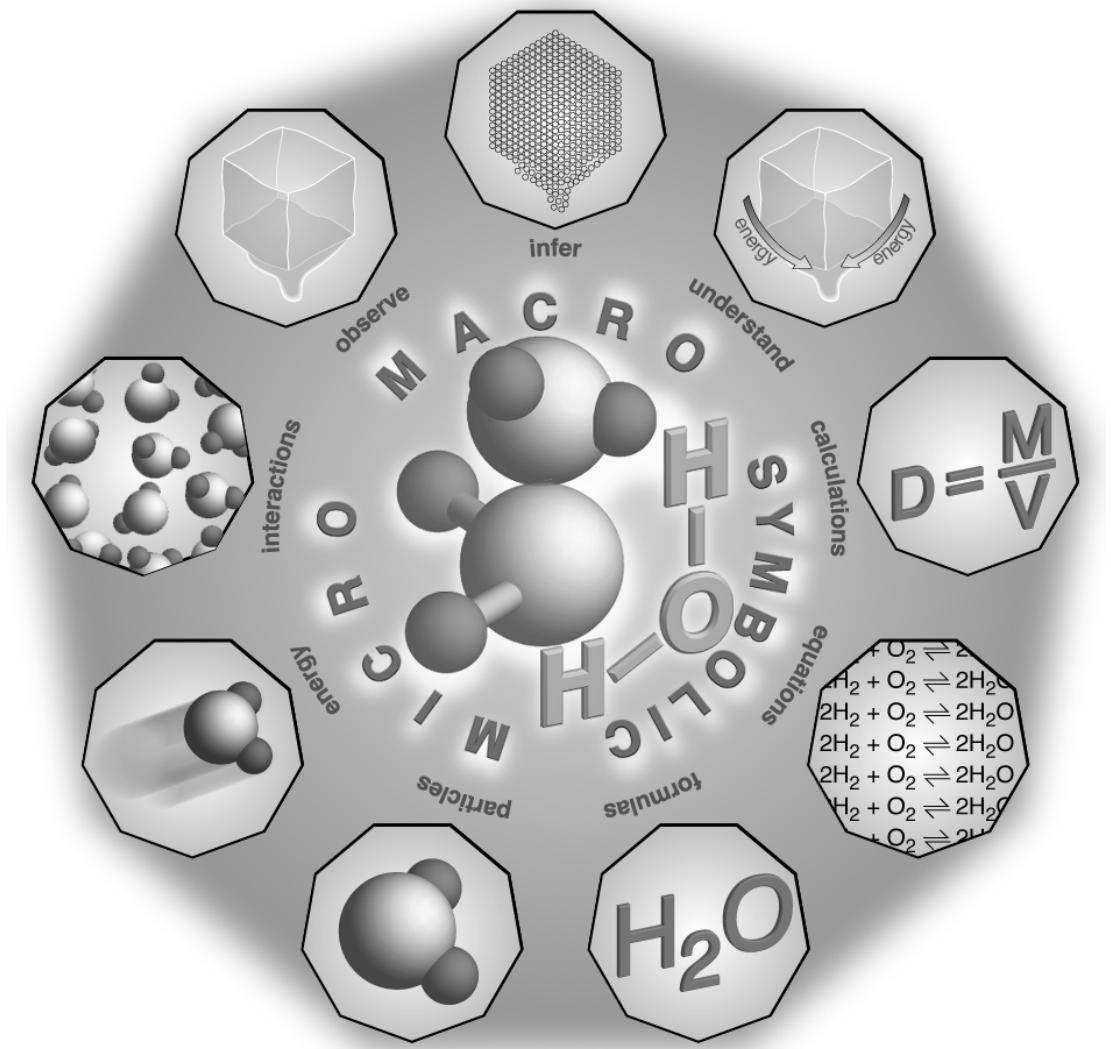


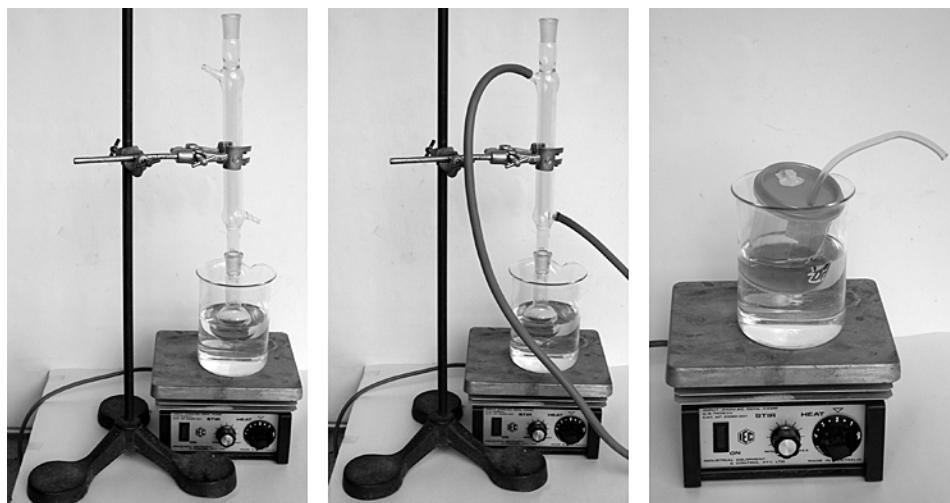
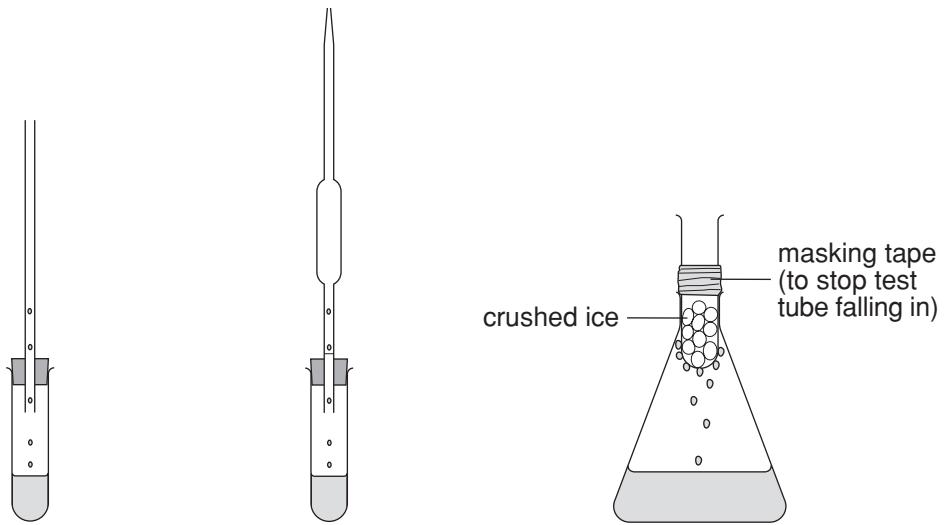


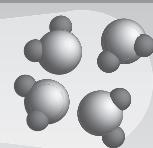


## The acidic environment

## Part 5: Esterification





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understand**SYMBOLIC** $H_2O$  formulas  
equations  
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energy  
interactions

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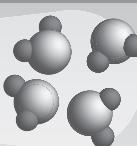
## MACRO

observe  
infer  
understand



## SYMBOLIC

$H_2O$  formulas  
equations  
calculations



## MICRO

particles  
energy  
interactions

# Introduction

The **flavours** of fruits, the **fragrances** of flowers, the sex attractants of insects, solid animal fats and liquid vegetable oils are all naturally occurring esters. Naturally occurring and synthetic esters are used extensively by the fragrance and cosmetic industries and in medicines.

Esterification – the reaction between an organic acid and an alkanol to form an ester and water – can be carried out and modelled in the laboratory.

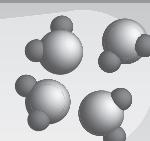
In Part 5 you will be given opportunities to learn to:

- describe the differences between the alkanol and alkanoic acid functional groups in carbon compounds
- explain the difference in melting point and boiling point caused by the alkanoic acid and alkanol functional groups
- identify esterification as the reaction between an acid and an alkanol and describe, using equations, examples of esterification
- describe the purpose of using concentrated sulfuric acid in esterification for catalysis and absorption of water
- explain the need for refluxing during esterification
- outline some examples of the occurrence, production and uses of esters.

In Part 5 you will be given opportunities to:

- identify data, plan, select equipment and perform a first-hand investigation to prepare an ester using reflux
- process information from secondary sources to identify and describe the uses of esters as flavours and perfumes in processed foods and cosmetics.

Extracts from *Chemistry Stage 6 Syllabus* © Board of Studies NSW, originally issued 1999. The most up-to-date version is to be found at [http://www.boardofstudies.nsw.edu.au/syllabus99/syllabus2000\\_list.html](http://www.boardofstudies.nsw.edu.au/syllabus99/syllabus2000_list.html)

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# Carbon compound functional groups

A **functional group** is an atom or group of atoms that reacts in a characteristic way in different carbon compounds.

The **hydroxy** functional group  $-OH$  provides alkanols with their characteristic properties. Alkanols have higher melting points and boiling points than other carbon compounds of similar molecular mass because of hydrogen bonding between the hydroxy groups.



The O in one hydroxy group can hydrogen bond to the H in a OH group of a nearby molecule. The  $-OH$  group also enables alkanols to dissolve polar molecules and some salts.

The electronegative O in OH attracts electrons from the attached carbon  $C^{\delta+}-O^{\delta-}-H$  making the carbon attached to the OH susceptible to attack by electron rich groups. Electron rich groups can react replacing the OH group. Such a reaction is called a substitution reaction.

Alkanols are named by removing the *-e* end of the alkane name and adding *-ol*. If the hydroxy group is on an end carbon the alkanol is a 1-alkanol; if on the second carbon from the end a 2-alkanol, and so on.



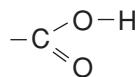
Complete naming the alkanes and alkanols in the table following.  
The *n-* prefix stands for normal meaning straight chained.

Structural formula	Systematic name	Common name
$CH_3CH_3$	ethane	ethane
$CH_3CH_2OH$	ethanol	ethyl alcohol
$CH_3CH_2CH_3$		propane

$\text{CH}_3\text{CH}_2\text{CH}_2\text{OH}$	1-propanol	n-propyl alcohol
$\text{CH}_3\text{CHOHCH}_3$		isopropyl alcohol
$\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_3$	butane	n-butane
$\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{OH}$		n-butyl alcohol
$\text{CH}_3\text{CH}_2\text{CHOHCH}_3$		sec-butyl alcohol
$\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_3$		n-pentane
$\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{OH}$		n-pentyl alcohol
$\text{CH}_3\text{CH}_2\text{CH}_2\text{CHOHCH}_3$		isopentyl alcohol

Check your answers.

Most naturally occurring acids are carbon compounds that contain the **carboxylic acid** functional group –COOH.

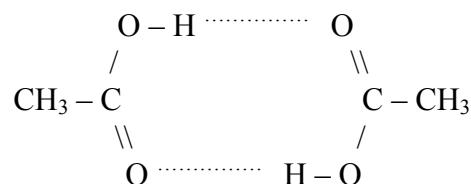


Unlike the – OH group which can be found on different parts of a hydrocarbon chain the – COOH group is only ever found at the end of a carbon chain.

The shared electron pair (covalent bond) between the H and its attached O atom in COOH is weak. The electrons in this bond are attracted towards the two electronegative (electron-attracting) oxygen atoms. When these organic acids come in contact with water the hydrogen in the –COOH can be attracted away as a  $\text{H}^+$  to a polar water molecule. This leaves a  $-\text{COO}^-$  behind on the acid, changing the acid molecule to a negative ion.



The alkanoic acids have higher boiling points than similar sized alkanols. Because an acid contains two O atoms two neighbouring alkanoic acid molecules can form two hydrogen bonds between one another:



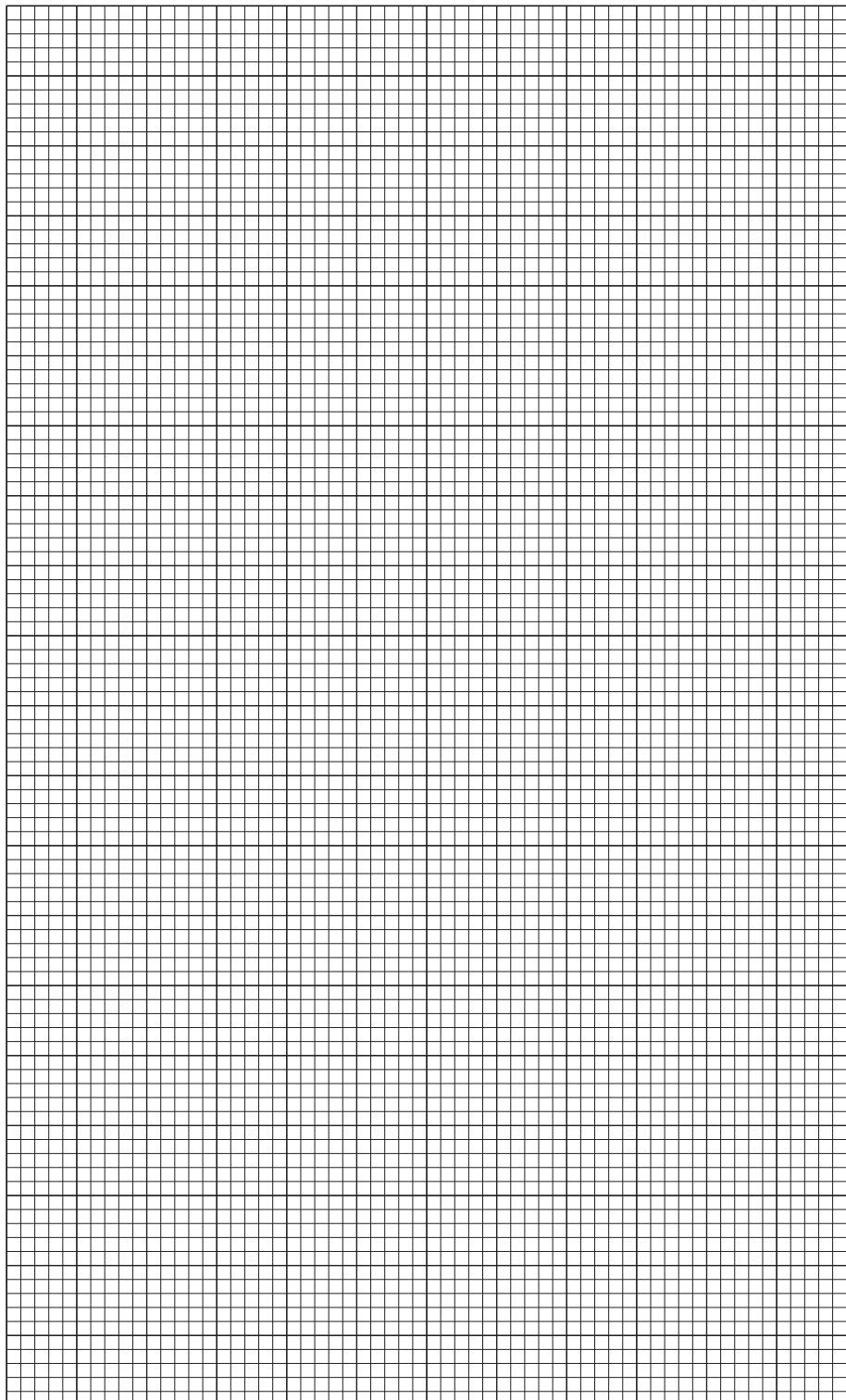


Use the supplied data to graph the melting points and boiling points of alkanoic acids and 1-alkanols on the next page. Explain the symbols you use in a **legend**. Then complete the written description and explanation below.

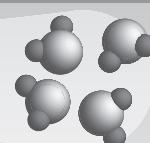
1-alkanol	MP (°C)	BP (°C)	alkanoic acid	MP (°C)	BP (°C)
methanol	-98	65	methanoic acid	8	101
ethanol	-114	78	ethanoic acid	17	118
1-propanol	-126	97	propanoic acid	-21	141
1-butanol	-90	118	butanoic acid	-5	163
1-pentanol	-79	138	pentanoic acid	-34	186
1-hexanol	-52	157	hexanoic acid	-3	205
1-heptanol	-34	176	heptanoic acid	-8	223
1-octanol	-16	195	octanoic acid	16	239



- Boiling point is a measure of change from \_\_\_\_\_ state to \_\_\_\_\_ state. The BP is a measure of the energy needed to overcome \_\_\_\_\_ bonding and dispersion forces between the molecules. The extent of hydrogen bonding is the same for all alkanols but as the hydrocarbon chain gets larger \_\_\_\_\_ forces become more significant. Similarly the extent of hydrogen bonding is the same for all alkanols but as the hydrocarbon chain gets larger \_\_\_\_\_ forces become more significant.
- Melting point does not change as regularly as BP. The MP is a measure of energy to change from \_\_\_\_\_ state to \_\_\_\_\_ state. In the solid state the molecules are closely packed in a regular arrangement. How the molecules pack together depends on their shape. MPs for the alkanols follow a regular pattern from \_\_\_\_\_ to \_\_\_\_\_. The first two alkanols – \_\_\_\_\_ and \_\_\_\_\_ – have higher MPs than expected. Hydrogen bonding must be particularly effective in their packing arrangement. The zigzag pattern for MPs of alkanoic acids where there can be up to \_\_\_\_\_ hydrogen bonds per molecule indicates varying effectiveness of hydrogen bonding according to packing of the molecules.

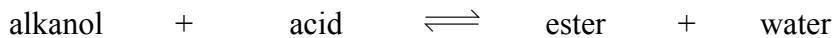


Check your answers.

**MACRO**observe  
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understand**SYMBOLIC** $H_2O$  formulas  
equations  
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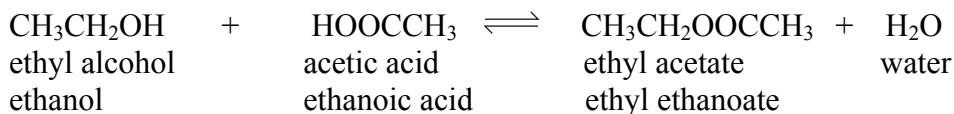
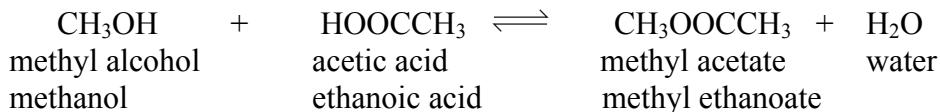
# Esterification

An acid containing the carboxylic acid group  $-COOH$  can react with an alkanol containing the  $-OH$  group to form an ester and water.

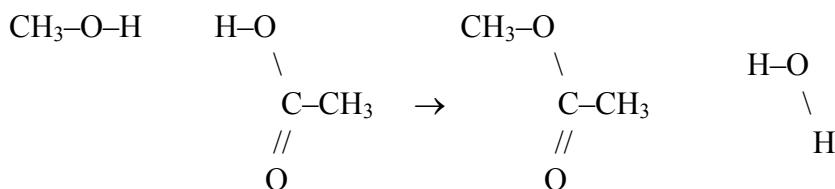


This acid catalysed, reversible reaction is called **esterification**.

Common names of the alkanol and acid are normally used to obtain the ester name. The first part of the ester name comes from the alkanol common name.



Esterification can also be called a **condensation** reaction because a water molecule condenses out.



In 1938 chemists using methanol containing  $^{18}\text{O}$  isotope found that the  $^{18}\text{O}$  isotope appeared in the ester but not the water. Use this evidence to predict where the atoms in the water came from.

Check your answer.

A small amount of acid is added to catalyse the reaction. At equilibrium significant amounts of the two reactants and two products are present. The equilibrium can be shifted significantly to the right, increasing the yield of ester if a relatively large amount (say 5 to 10% of the reactant volume) of concentrated sulfuric acid is used. The concentrated sulfuric acid acts as a dehydrating agent as well as an acid catalyst. By absorbing the water and lowering the water concentration the equilibrium shifts to the right increasing the yield of ester.

## Preparing esters

### A: Using laboratory chemicals



#### What you will need:

- a selection of alcohols such as ethyl, propyl, butyl and pentyl
  - acetic (ethanoic) acid and butyric (butanoic) acid
  - concentrated sulfuric acid in a dropper bottle
  - test tubes to hold the mixtures of alcohol, acid and conc.  $\text{H}_2\text{SO}_4$
  - a means of heating the test tubes such as a container of heated water on an electric hotplate
- 1 Methyl alcohol should not be used as its vapour is much more toxic than other alcohol vapours and it is more volatile than other alcohols.
  - 2 Keep all alcohols and organic acids away from flames or sparking.
  - 3 Concentrated sulfuric acid is highly corrosive to skin and eyes. Use eye and skin protection while working near a cold water tap.



#### What you will do:

- 1 Add equal volumes (eg. 2 mL) of an acid and an alcohol to a clean, dry test tube.
- 2 Add 2 drops of conc.  $\text{H}_2\text{SO}_4$  to the mixture then shake the test tube.
- 3 Place the test tube in the hot water.
- 4 Smell the vapour emitted from the test tube every ten minutes and note any change in odour. Enter ester names into the table below which includes *odours* noted by a class of students:

	ethyl alcohol -'metho'	propyl alcohol -sharp fruity	butyl alcohol -glue solvent	pentyl alcohol -ink solvent
ethanoic acid-vinegar	-solvent	-pear	-apple	-banana
butanoic acid -rancid	-pineapple	-monstera fruit	-passionfruit	-apricot

## B: Using household chemicals

Even if you do not carry out this activity read through the information and attempt the questions. This will give you a better understanding of:

- the steps organic chemists carry out when they are reacting and investigating carbon compounds
- separation techniques
- safe handling of chemicals.

### What you will need:

- 10 aspirin (not soluble aspirin) tablets; each tablet has a mass of about 0.3 g aspirin and contains starch binder and cellulose filler
- 15 mL methylated spirits
- 10 mL concentrated hydrochloric acid; often sold as muriatic acid at hardware stores for cleaning mortar off bricks or lowering the pH of swimming pools
- 1 g of boric (boracic) acid  $\text{B}(\text{OH})_3$ ; a solid acid sold in pharmacies and used in eyebaths, wound ointments and ant poisons
- two plastic spoons
- means of measuring or estimating liquid volumes to 0.5 mL
- filter and filter papers
- paper towel or filter papers for drying product
- 3 test tubes and 3 small (100 to 250 mL) beakers
- a source of hot water and a container in which test tubes can be placed
- thermometer (0 to 100°C)
- crushed ice or small iceblocks.

- 1 Methylated spirits is flammable. Keep it well away from flames or sparks.
- 2 Concentrated hydrochloric acid is highly corrosive to skin and eyes. Use eye and skin protection and work near a cold water tap.



## What you will do:

### Separation of pure acetylsalicylic acid (ASA) crystals from tablets

1 Add 10 aspirin tablets to 10 mL of methylated spirit in a test tube. Place in a container of water at about 60-70°C. Shake occasionally so that the tablets break up and the ASA dissolves.



2 Filter the solution and wash the residue with about 2 mL of methylated spirits collecting the washing with the filtrate. The residue is mostly starch and cellulose. Describe the residue:

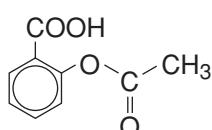
- 
- 3 Add about 50 mL of water to the methylated spirits filtrate so that acetylsalicylic acid crystals form. Cool the mixture by surrounding its container in a container of crushed ice for at least 5 min. Filter off the crystals of acetylsalicylic acid (ASA). Discard the filtrate down a sink hole to a sewer.
- 4 Carefully remove the moist filter paper holding the ASA crystals and place it on another filter paper. Using a spoon scrape the crystals on to another filter paper or paper towel. After a minute scrape the crystals on to another dry filter paper or paper towel and spread out the ASA crystals to encourage drying.

Compare the ASA crystals with the residue of starch and cellulose previously separated from the aspirin tablets.

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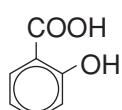
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The structural formula for acetylsalicylic acid is:



This carbon compound contains the acid functional group –COOH and an ester functional group –COOC (the last C is at a vertex of the hexagon representing a six carbon benzene ring). Because the properties of the –COOH are more prominent than the properties of the –COOC in this compound the molecule is named as an acid.

Your next task is to remove the acetyl ester functional group to make salicylic acid:



- What are the two functional groups in salicylic acid?  
\_\_\_\_\_
- From the name, salicylic acid, which of the two functional groups would be most reactive?  
\_\_\_\_\_

Check your answers.

### Changing acetylsalicylic acid (ASA) to salicylic acid (SA)



The first step requires you to dilute the 10 mL of concentrated hydrochloric acid by adding it carefully to 10 mL of water. **ALWAYS ADD THE ACID TO THE WATER.** Make sure you are wearing protective clothing, eye protection and are near a cold water tap to flood any acid affected body parts with water.

- 1 Carefully pour 10 mL of concentrated hydrochloric acid into 10 mL of water in a test tube. Gently shake the test tube to produce a solution that is about 15% HCl (w/w).
- 2 Add the acetylsalicylic acid to the hydrochloric acid solution and place the test tube in a container of water kept at 50–60°C. Swirl and shake the mixture every five minutes for 30 minutes. When the mixture thickens carefully smell air from above the top of the test tube. Can you detect a familiar odour? \_\_\_\_\_

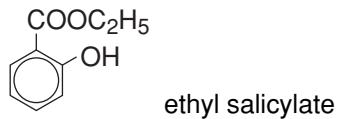


Check your answer.

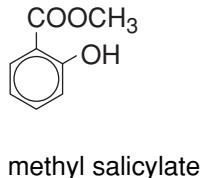
- 3 Cool the test tube by placing it in a beaker of cooling water for 10 minutes then add 20 mL of cold water. This should separate out the salicylic acid (SA).
- 4 Filter the salicylic acid residue from the filtrate. Wash the SA twice with 2 mL of cold water to remove acetic acid. Discard the filtrate.
- 5 Carefully remove the moist filter paper holding the SA crystals and place it on another filter paper. Using a spoon scrape the crystals on to another filter paper or paper towel. After a minute scrape the crystals on to another dry filter paper or paper towel and spread out the SA crystals to encourage drying. Leave overnight to dry in a dry warm place away from the reach of children and animals. It is important that the SA is as dry as possible for its esterification.

## Changing salicylic acid (SA) to the ester ethyl salicylate

The ethyl alcohol of methylated spirits can react with salicylic acid to form ethyl salicylate ester which has a distinctive odour.



Any methanol in the methylated spirits will also react forming methyl salicylate which has a similar odour. Methyl salicylate, oil of wintergreen, is rubbed onto skin to relieve the pain of rheumatism and strained muscles.



- 1 Add the SA and a few rice grain volumes of boric acid to a test tube. Then add about 3 mL of methylated spirits.
- 2 Place the test tube in a container of hot water maintained at 80-90°C. Shake the test tube every five minutes for at least 30 min. Can you detect a new, possibly familiar odour? If not, heat without shaking for another 30 min.
- 3 When you have smelt the new odour for at least 10 min pour the test tube contents into a container of water. Most esters float in water but this ester is denser than water. Tilt the container to see ester at the bottom corner.

- 1 Write an equation using structural formulas to show how salicylic acid and ethanol react to form ethyl salicylate and water:



- 2 It was important that the SA was dried thoroughly for the esterification. If it was not dry what effect would this have on the yield of ethyl salicylate?

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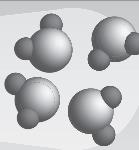
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- 3 The second stage of this activity – **ester hydrolysis** was the reverse of esterification. The reaction of the ester acetylsalicylic acid with water was catalysed by the hydrochloric acid. Salicylic acid was the alcohol product and acetic acid was the acid product of this reaction. Write an equation using structural formulas for this hydrolysis.



- 4 Most aspirin sold is soluble aspirin containing a sodium or calcium acetylsalicylate salt. When the acetylsalicylate ions reach the stomach containing hydrochloric acid the acetylsalicylate ions change to molecular acetylsalicylic acid. Write an equation using structural formula for this acid-base reaction and label the acid and the base which react.

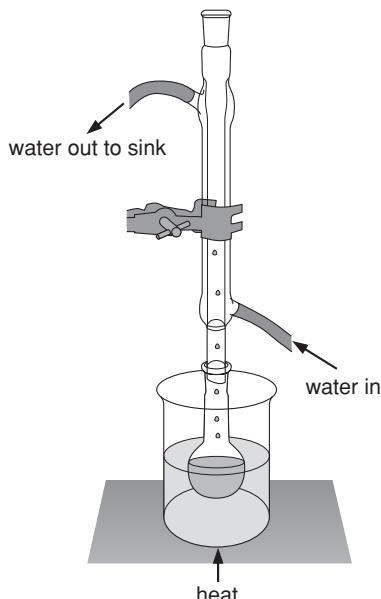


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# Refluxing

**Refluxing** is a process which enables vapour escaping from a reaction mixture to condense and return to the reaction mixture. This prevents loss of volatile reactants or volatile products from the reaction mixture.

Preparing an ester using **reflux** prevents loss of the volatile alkanol and volatile ester.

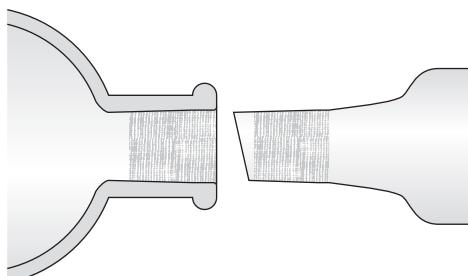


Using reflux to react and retain chemicals in a flask.

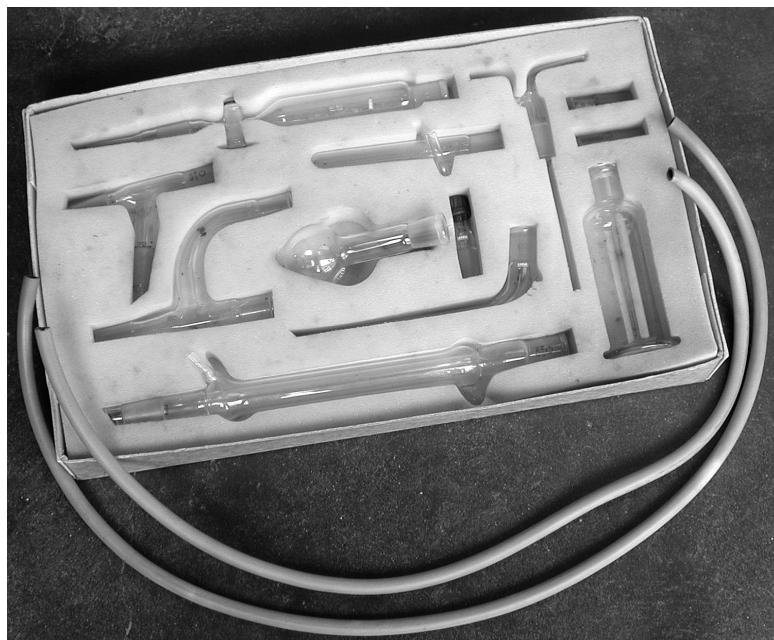
The condenser is normally cooled by cooling water flowing in the bottom of the outer jacket and leaving with absorbed heat by the top. If the reaction mixture is not heated to a high temperature an air cooled condenser could be used where air moving around the condenser condenses vapours.

Heat resistant components with ground glass joints can be joined to make up refluxing apparatus. The 'male' joint with an outer ground glass surface is inserted into a 'female' joint with an inner ground glass surface.

When the male joint is fully inside the female joint it is slightly rotated about five to ten degrees of angle to provide a gas tight seal. When separating the male and female part the two parts are held firmly and pulled apart and twisted at the same time.



Forming a ground glass joint.



Ground glass joint component kit with rubber tubing for the condenser

#### Procedure for using reflux to prepare an ester:

- 1 Place the alkanol, alkanoic acid and drops of concentrated sulfuric acid into a flask.
- 2 Add two or three boiling chips made of porous ceramic material to the reaction mixture in the flask; these will promote steady boiling of the reaction mixture when it is heated.
3. Attach a condenser to the flask and clamp the condenser vertically. If ground glass joints are used the flask will be supported by attachment to the **condenser**. Turn on the water supply slowly until a steady water flow passes in the bottom and out the top of the condenser.

- 4 Heat the flask by immersion in a beaker of hot water preferably heated by an electric hot plate or heating mantle. The flask can be heated with a burner flame but electric heating by water is safer.
- 5 The mixture should boil for about half an hour to reach an equilibrium mixture of alcohol, acid, ester and water.
- 6 Remove the heating source and allow the mixture to cool for 10 minutes. Using a towel hold either side of the joint. Separate the condenser and flask by pulling apart and twisting in one movement.
- 7 Transfer the mixture to a separating funnel containing water of about the same volume as the mixture. Shake the contents and allow the layers to separate. The lower layer contains water soluble acid and alkanol and can be run off and discarded. Esters cannot hydrogen bond and so have low water solubility.
- 8 Add solid  $\text{Na}_2\text{CO}_3$  or  $\text{NaHCO}_3$  to the remaining mixture in the separating funnel until no more reaction with acid occurs. Add an equal volume of water to dissolve any salts formed. Shake, stand and let layers separate then discard the lower aqueous layer.
- 9 Place the remaining layer in a flask, add two or three boiling chips then distill the mixture collecting different fractions noting the temperature range at which they boiled. Detect the ester fraction(s) by smell or by looking up the boiling points of the alcanoic acid, alkanol and ester.

## Selecting equipment



Select the equipment that is needed to carry out steps 3, 7 and 9 by looking at the diagrams of ground glass joint components in the *Appendix*. Draw how you would arrange the components, and any additional equipment needed, to carry out. Alternatively you could cut out, arrange and paste parts from the kit diagrams in the *Appendix*.

- a) refluxing in step 3

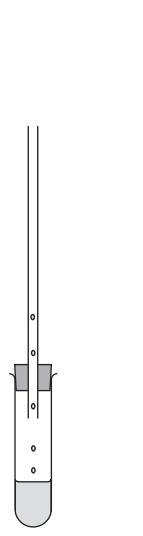


b) separation in step 7

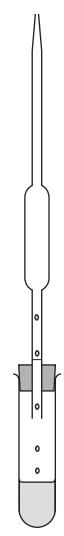
c) distillation in step 9.

Check your answers.

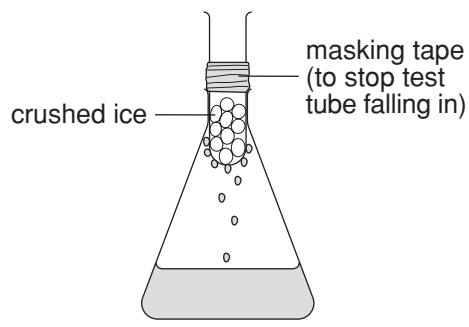
# Reflux methods



tube condenser

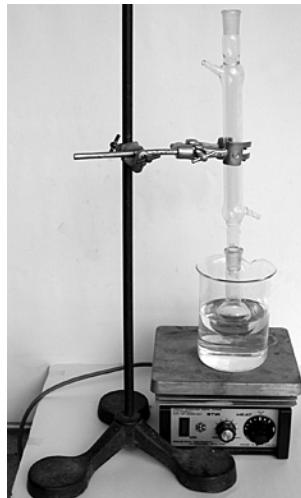


pipette condenser



cold finger condenser

The masking tape on the test tube sits loosely on, but is not attached to, the top of the conical flask.



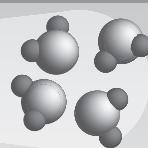
air cooled condenser



water cooled condenser



air cooled condenser  
using the plastic tube of  
a child's drink bottle (the  
air hole of lid is sealed)

**MACRO**observe  
infer  
understand**SYMBOLIC** $H_2O$  formulas  
equations  
calculations**MICRO**particles  
energy  
interactions

# Suggested answers

## Carbon compound functional groups

Structural formula	Systematic name	Common name
$CH_3CH_3$	ethane	ethane
$CH_3CH_2OH$	ethanol	ethyl alcohol
$CH_3CH_2CH_3$	propane	propane
$CH_3CH_2CH_2OH$	1-propanol	n-propyl alcohol
$CH_3CHOHCH_3$	2-propanol	isopropyl alcohol
$CH_3CH_2CH_2CH_3$	butane	n-butane
$CH_3CH_2CH_2CH_2OH$	1-butanol	n-butyl alcohol
$CH_3CH_2CHOHCH_3$	2-butanol	sec-butyl alcohol
$CH_3CH_2CH_2CH_2CH_3$	pentane	n-pentane
$CH_3CH_2CH_2CH_2CH_2OH$	1-pentanol	n-pentyl alcohol
$CH_3CH_2CH_2CHOHCH_3$	2-pentanol	isopentyl alcohol

- 1 Boiling point is a measure of change from *liquid* state to *gas* state. The BP is a measure of the energy needed to overcome *hydrogen* bonding and dispersion forces between the molecules. The extent of hydrogen bonding is the same for all alkanols but as the hydrocarbon chain gets larger *dispersion* forces become more significant. Similarly the extent of hydrogen bonding is the same for all alkanols but as the hydrocarbon chain gets larger *dispersion* forces become more significant.

- 2 Melting point does not change as regularly as BP. The MP is a measure of energy to change from *solid* state to *liquid* state. In the solid state the molecules are closely packed in a regular arrangement. How the molecules pack together depends on their shape. MPs for the alkanols follow a regular pattern from *1-propanol* to *1-octanol*. The first two alkanols – *methanol* and *ethanol* – have higher MPs than expected. Hydrogen bonding must be particularly effective in their packing arrangement. The zigzag pattern for MPs of alkanoic acids where there can be up to *two* hydrogen bonds per molecule indicates varying effectiveness of hydrogen bonding according to packing of the molecules.

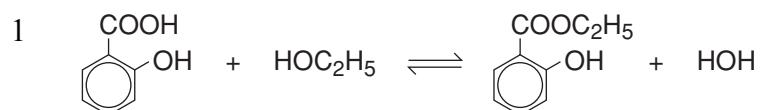
## Esterification

One hydrogen in water came from the alcohol OH group while the oxygen and other hydrogen came from the OH of the acid COOH.

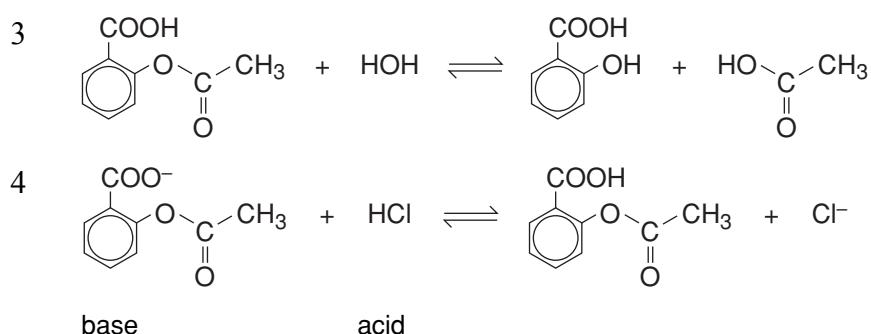
## Preparing esters B: Using household chemicals

- 2 A white mass without crystals
  - 4 The acetylsalicylic acid (ASA) crystals are finer and whiter than the starch and cellulose residue.
  - The two functional groups in salicylic acid are the hydroxy OH group and the carboxylic acid COOH group.
  - Because the compound is named as an acid the acid group would be expected to be more reactive than the hydroxy group.

The odour of vinegar, acetic acid, should be detectable.

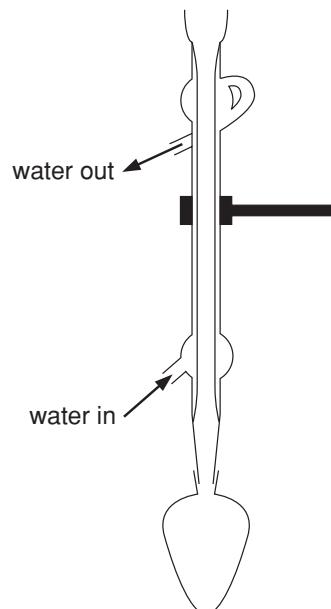


- 2 If water was present with the SA crystals the yield of ester would be lower. Water is a product of the reaction and, by Le Chatelier's principle, shifts the equilibrium position to the reactant side reducing yield of products.

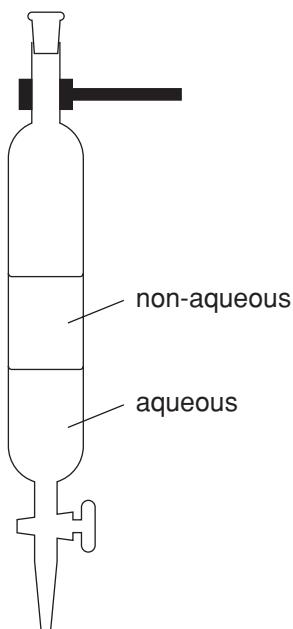


## Selecting equipment

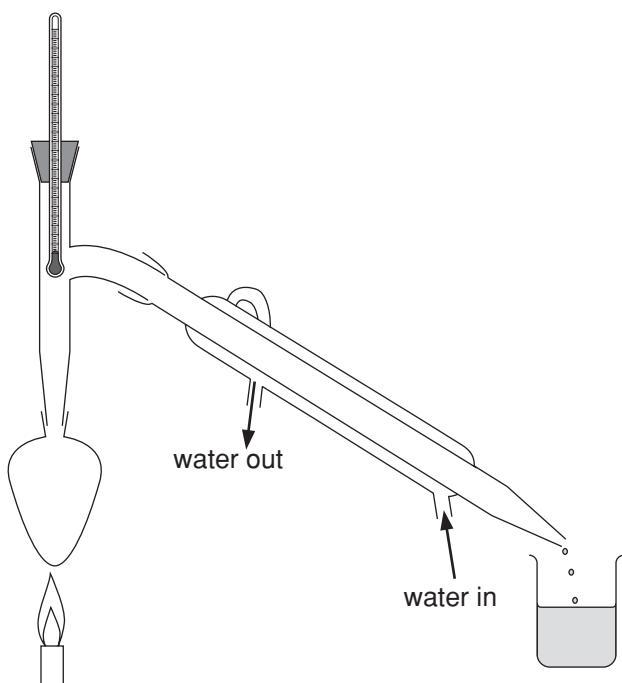
a) refluxing



b) separation



c) distillation



**MACRO**

observe  
infer  
understand



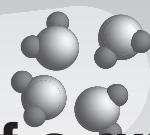
**SYMBOLIC**

$H_2O$  formulas  
equations  
calculations

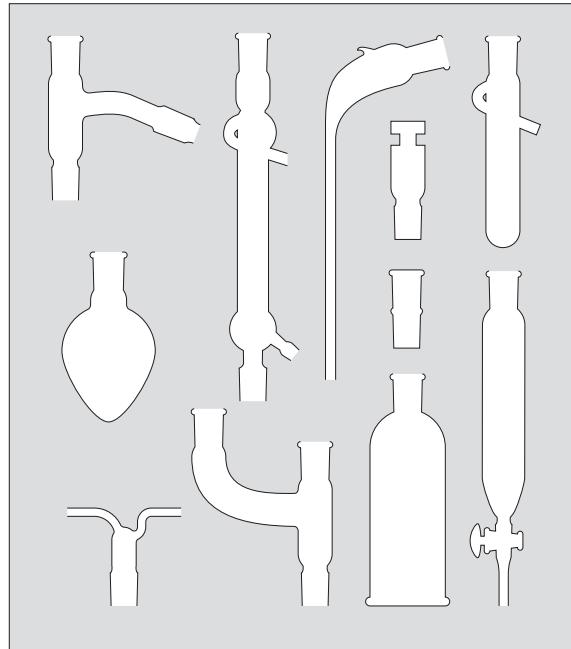
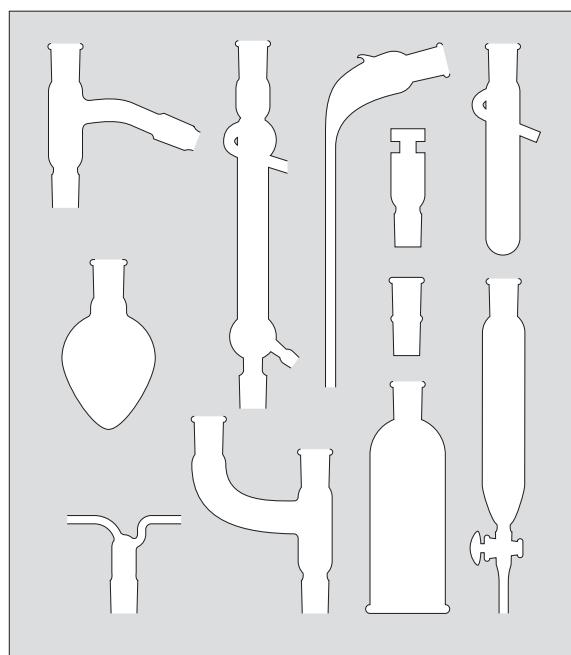


**MICRO**

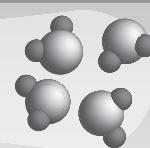
particles  
energy  
interactions



## Appendix: Components of a ground glass joint kit





**MACRO**observe  
infer  
understand**SYMBOLIC** $H_2O$  formulas  
equations  
calculations**MICRO**particles  
energy  
interactions**Exercises – Part 5**

Exercises 5.1 to 5.2

Name: \_\_\_\_\_

**Exercise 5.1: Variety of esters in the supermarket**

Every supermarket has many products which contain esters. Many of the esters are extracted from natural resources such as vegetable oils, animal fats and natural flavouring essences. Others are made from chemicals in animal or plant products or from petrochemicals. The table below gives a variety of esters found in a supermarket's products. Complete the table.

Ester name	Alcohol name	Acid name	Use	Product
butyl phthalate	butyl	phthalic	plastic softener	flexible child's toy
propyl hydroxybenzoate			preservative	toothpaste
octyl methoxycinnamate			UV absorber	sunscreen
triethyl citrate	ethyl		thickener	dessert
octyl gallate			antioxidant	processed food
poly(methyl 2-cyanoacrylate)		2-cyano acrylic	strong adhesive	superglue
2-propyl myristate			flow agent	lipstick
methyl salicylate				
dimethyl phthalate				insect repellent

## **Exercise 5.2: Which artificial sweetener is an ester?**

Here is an opportunity for you to work out which artificial sweetener in the supermarket is an ester. You will use your knowledge that esters undergo hydrolysis and can break down when heated with water. An artificial sweetener which is an ester would therefore not be used in cooking where ingredients are mixed with water and heated. The label may recommend that the sweetener only be added after cooking is complete.

Go to the artificial sweetener section of a supermarket – usually close to the cane sugar (sucrose) section. Here is a list of artificial sweeteners approved by the ANZFA (Australian New Zealand Food Authority):

<b>Number</b>	<b>Name</b>	<b>Relative sweetness (w/w) relative to sucrose</b>
950	Acesulphame potassium	150
951	Aspartame	150
952	Cyclamates	50
954	Saccharins	400
955	Sucratose (trichlorogalactosucrose)	650
956	Alitame	2000
957	Thaumatin	1000

- a) Can you work out which of these is an ester? \_\_\_\_\_  
b) What information have you found to support your conclusion?

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# Chemistry

HSC Course

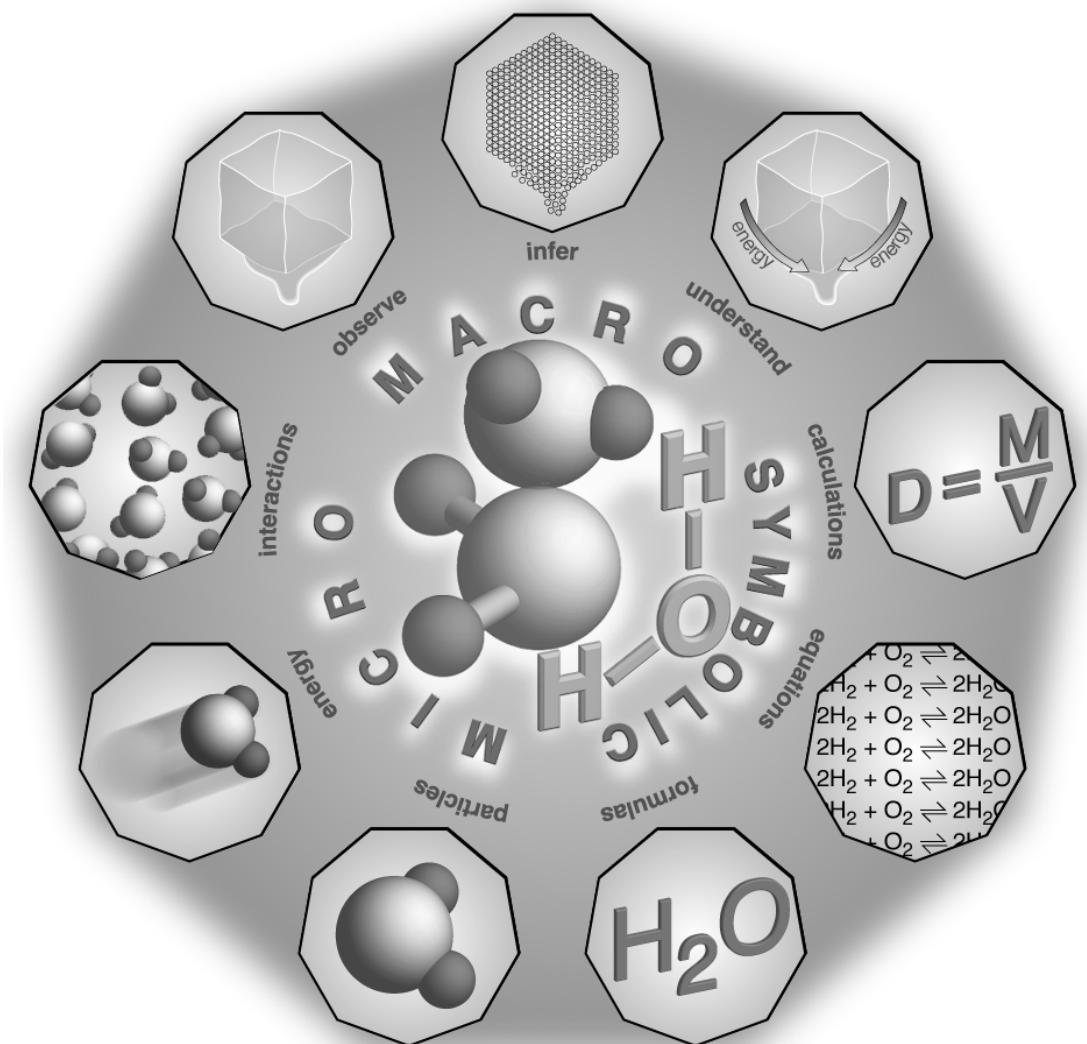
Stage 6

NEW SOUTH WALES  
DEPARTMENT  
OF EDUCATION  
AND TRAINING



## The acidic environment

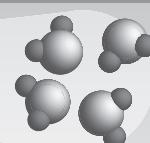
Part 6: Key ideas, mole calculations and skill outcomes



Name	Atomic weight	Symbol	Atomic number
Hydrogen	1.008	H	1
Lithium	6.94	Li	3
Beryllium	9.01	Be	4

2	He	4.003										
	Helium											
5	6	7	8	9	F	10						
B	C	N	O	Fluorine	Ne							
10.81	12.01	14.01	16.00	19.00	20.18							
Baron	Carbon	Nitrogen	Oxygen	Chlorine	Neon							
13	14	15	16	17								
Al	Si	Phosphorus	Sulfur	Cl	Ar							
26.98	28.09	30.97	32.07	35.45	39.95							
	Silicon											
29	30	31	32	33	34	35	36					
Cu	Zn	Ga	Ge	As	Se	Br	Kr					
63.55	65.39	69.72	72.61	74.92	78.96	79.90	83.80					
Copper	Zinc	Gallium	Germarium	Arsenic	Selenium	Bromine	Krypton					
47	48	In	Sn	Sb	Te	I						
Ag	Cd	Indium	Tin	Antimony	Tellurium	Iodine						
107.9	112.4	114.8	118.7	121.8	127.6	126.9						
Silver	Cadmium											
79	80	81	82	83	84	85	86					
Au	Hg	Tl	Pb	Bi	Po	At	Rn					
197.0	200.6	204.4	207.2	209.0	[210.0]	[210.0]	[222.0]					
Gold	Mercury	Thallium	Lead	Bismuth	Poliobrium	Astatine	Radon					
—	Uuu	—	Uuq	—	Ununquadium	—	Ununhexium					
—	Uuu	—	—	—	Ununtrium	—	Ununpentium					
—	—	—	—	—	Ununtrium	—	Ununpentium					

57	La	58	Ce	59	Pr	60	Nd	61	Pm	62	Sm	63	Eu	64	Gd	65	Tb	66	Dy	67	Ho	68	Er	69	Tm	70	Yb	71	Lu	75.0	Lutetium
138.9	Cerium	140.1	Praseodymium	140.9	Neodymium	144.2	Neodymium	144.2	Promethium	150.4	Samarium	150.4	Europium	157.3	Gadolinium	158.9	Terbium	162.5	Dysprosium	164.9	Thulium	167.3	Erbium	168.9	Ytterbium	173.0	Ytterbium	175.0	Lutetium		
89	Ac	90	Th	91	Pa	92	U	93	Pu	94	Cm	95	Am	96	Bk	97	Cf	98	Es	99	Fm	100	Md	101	No	102	Lr	103	[262.1]		
[227.0]	Actinium	[232.0]	Thorium	[231.0]	Protactinium	[238.0]	Uranium	[237.0]	Neptunium	[239.1]	Plutonium	[241.1]	Americium	[244.1]	Curium	[249.1]	Berkelium	[252.1]	Einsteinium	[252.1]	Californium	[257.1]	Fermium	[258.1]	Mendelevium	[259.1]	Nobelium	[259.1]	Lawrencium		

**MACRO**observe  
infer  
understand**SYMBOLIC** $H_2O$  formulas  
equations  
calculations**MICRO**particles  
energy  
interactions

# Contents

Introduction .....	2
Key ideas .....	3
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## MACRO

observe  
infer  
understand



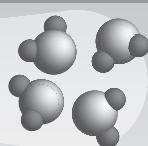
## SYMBOLIC

$H_2O$  formulas  
equations  
calculations



## MICRO

particles  
energy  
interactions



# Introduction

In this part you will review the first five parts of this module and revise quantitative calculations using the mole concept. The exercises will give you further practice in mole calculations, and enable you to assess your chemistry skills for the H11 to H15 HSC course outcomes.

## MACRO

observe  
infer  
understand



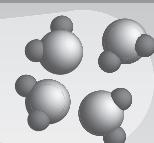
## SYMBOLIC

$H_2O$  formulas  
equations  
calculations



## MICRO

particles  
energy  
interactions



# Key ideas

Assess your understanding of this module by answering these true/false statements one part at a time. Write T or F at the end of each statement. Approximately half the statements are true. If you decide the statement is false, rewrite the statement so that it is true.

## Indicators

- Common substances can be classified as acidic, basic or neutral according to their pH in water.
- All acid-base indicators come from plants.
- Water always has a pH of 7.
- Water insoluble barium sulfate powder is part of most soil pH kits.
- If acid or base is added to an indicator solution a shift in equilibrium occurs.
- Adding acid to a blue litmus solution can produce a colour change.
- If red litmus does not change colour in a solution then the solution must be acidic.
- $NaHCO_3$  produces a slightly basic solution in water.
- An indicator is a dye sensitive to acid or base.
- Sea water is usually acidic.

## Acid in the atmosphere

- Most non-metal oxides are acidic.
- Dry  $SO_2$  gas will cause dry litmus paper to turn red.
- A solution of  $CO_2$  in water is acidic.

- The molar volume of most gases at 0°C and 101.3 kPa is about 24.5 litres.
- Rain water with a pH of 5.5 would be called acid rain.
- The solubility of carbon dioxide in water decreases as the temperature rises.
- When ionic substances are added to water the solubility of molecules of dissolved gas decreases.
- Lightning causes nitrogen and oxygen in the air to react together directly forming nitrogen dioxide.
- Bases are water soluble alkalis.
- Basicity of metal oxides increases going down a group of the periodic table.

## Acids in water

- Most hydrogen ions are protons.
- $[H^+] \times [OH^-] = 10^{-14}$  at 25°C for aqueous acid solutions.
- $[H^+] \times [OH^-] = 10^{-14}$  at 25°C for all aqueous salt solutions.
- A solution with pH > 7 has a  $[H^+] < 10^{-7}$
- A weak acid solution has a lower concentration of ions compared with a strong acid solution of the same concentration of acid.
- A concentrated acid solution always contains a high concentration of ions.
- Acetic acid CH<sub>3</sub>COOH is a polyprotic acid.
- Some ions can be acids.
- Basic substances are added to foods as preservatives.
- K<sub>2</sub>HPO<sub>4</sub> is called potassium dihydrogen phosphate.

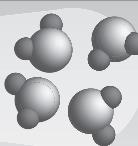
## Acid concepts and titrations

- The Arrhenius acid definition is the same as the Bronsted-Lowry acid definition.
- Carbonic acid is the conjugate acid of the hydrogen carbonate ion.
- Neutralisation is an exothermic reaction.
- Buffers are usually made up of a weak acid and its conjugate base.

- A volumetric flask must be dry before it can be used to make up a standard solution.
- A pipette filler should always be used to fill a pipette.
- If titration of a strong acid and a strong base produces an insoluble salt the conductivity of the solution will be a minimum at the equivalence point.
- A Lewis acid is an electron pair donor
- All Lewis acids are Bronsted-Lowry acids
- 'Have pair (of electrons), will share' could define a Lewis base.

## Esterification

- There can be twice as much hydrogen bonding between two alkanoic acid molecules compared with between two alkanol molecules.
- The relatively high MPs of methanol and ethanol indicate effective hydrogen bonding in the solid state.
- Boiling chips are used to catalyse esterification.
- Ester hydrolysis is the reaction of an ester with hydrogen.
- When using a condenser for reflux the cooling water should flow out of the bottom.
- Esters cannot hydrogen bond and so have very low water solubility.
- In a separating funnel most ester layers float on top of the aqueous layer.
- A condenser is arranged vertically for refluxing and distillation.
- Adding carbonate to a mixture containing organic acid and water molecules changes the acid to a more soluble ionic form.
- Care should be taken when carrying out esterification using a flame because esters are flammable.

**MACRO**observe  
infer  
understand**SYMBOLIC** $\text{H}_2\text{O}$  formulas  
equations  
calculations**MICRO**particles  
energy  
interactions

# Mole calculations

GAS VOLUME in L  
at 101.3 kPa

at 298 K (25°C)

at 273 K (0°C)

$$\begin{array}{ccccccc} \boxed{n = \frac{m}{M}} & \text{MASS IN g} & \xrightarrow[\div]{\text{molar mass } M} & \text{MOLES } n & \xrightarrow[\div]{N_A 6 \times 10^{23}} & \text{NUMBER OF PARTICLES } N & \boxed{n = \frac{N}{N_A}} \\ & m & & & & & \end{array}$$

$$\times \left| \begin{array}{c} v \\ \text{in litres} \end{array} \right| \div$$

SOLUTION MOLARITY IN molL<sup>-1</sup>  
c

$$\boxed{c = \frac{n}{V}}$$



Making sure that you have the correct formula for each chemical, using atomic masses from the periodic table on the inside cover of this part and giving answers to the appropriate numbers of significant figures, use the mole calculations diagram to calculate the following:

1 In one mole of water what is the number of:

a) water molecules

---

b) oxygen atoms

---

c) hydrogen atoms

---

d) atoms?

---

2 Calculate the molarity of the following:

- a) 2 moles of  $\text{CaCl}_2$  in 5 L of solution
- 

- b)  $\text{Ca}^{2+}$  in the solution in part a)
- 

- c)  $\text{Cl}^-$  in the solution in part a)
- 

3 Calculate the mass of:

- a) 2.000 moles of copper
- 

- b) 0.200 moles of sulfur dioxide
- 

- c) 1.20 moles of anhydrous copper sulfate
- 

4 What is the gas volume of:

- a) 0.50 moles of nitrogen monoxide at  $25^\circ\text{C}$  and 101.3 kPa
- 

- b) 0.50 moles of nitrogen dioxide at  $25^\circ\text{C}$  and 101.3 kPa
- 

- c) 0.50 moles of carbon dioxide at  $0^\circ\text{C}$  and 101.3 kPa?
- 

5 How many moles is:

- a)  $6 \times 10^{21}$  molecules of carbon dioxide
- 
- 

- b) there of solute in 600 mL of  $0.15 \text{ molL}^{-1}$  solution
- 
- 

- c) in 5.85 g of sodium chloride
- 
-

- d) 5.85 g of hydrogen gas at 273 K and 1.00 atm

---

---

- e) 5.85 g of oxygen gas at 298 K and 101.3 kPa?

---

---

The next question requires two step calculations. The first step involves calculating the number of moles. The second step involves converting this number of moles to another quantity. The mole is the link between the two quantities. Exercise 6.1 consists of calculations of this type.

- 6 a) How many grams of hydrogen chloride gas can be released by boiling 1.5 L of 0.60 M HCl(aq)?

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- b) What volume is occupied by 5.1 g of ammonia at 0°C and 101.3 kPa pressure?

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- c) What is the molarity of 100 g of ammonium chloride dissolved in 560 mL of water?

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d) What is the mass of  $5.0 \times 10^{22}$  sulfuric acid molecules?

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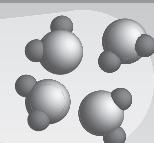
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Check your answers.



Do Exercises 6.1 and 6.2 now.



**MACRO**observe  
infer  
understand**SYMBOLIC** $H_2O$  formulas  
equations  
calculations**MICRO**particles  
energy  
interactions

# Suggested answers

## Indicators

- Common substances can be classified as acidic, basic or neutral according to their pH in water. T
- All acid-base indicators come from plants. F  
*Some acid-base indicators are synthesised from other chemicals.*
- Water always has a pH of 7. F  
*Water which has absorbed carbon dioxide from air and rain water have a pH < 7. Sea water has a pH above 7.*
- Water insoluble barium sulfate powder is part of most soil pH kits. T
- If acid or base is added to an indicator solution a shift in equilibrium occurs. T
- Adding acid to a blue litmus solution can produce a colour change. T
- If red litmus does not change colour in a solution then the solution must be acidic. F  
*The solution could be neutral and the red litmus would not change colour.*
- $NaHCO_3$  produces a slightly basic solution in water. T
- An indicator is a dye sensitive to acid or base. T
- Sea water is usually acidic. F  
*Sea water is slightly basic with a pH about 8.5.*

## Acid in the atmosphere

- Most non-metal oxides are acidic. T
- Dry  $SO_2$  gas will cause dry litmus paper to turn red. F  
*Moisture is needed to dissolve the  $SO_2$  to form  $H^+$  ions.*
- A solution of  $CO_2$  in water is acidic. T
- The molar volume of most gases at  $0^\circ C$  and 101.3 kPa is about 24.5 litres. F

*The molar volume of most gases at 0°C and 101.3 kPa is about 22.4 litres.*

- Rain water with a pH of 5.5 would be called acid rain. F

*In Australia unpolluted rainwater has a pH between 5 and 6 due to dissolved CO<sub>2</sub> from the atmosphere.*

- The solubility of carbon dioxide in water decreases as the temperature rises. T
- When ionic substances are added to water the solubility of molecules of dissolved gas decreases. T
- Lightning causes nitrogen and oxygen in the air to react together directly forming nitrogen dioxide. F

*Nitrogen and oxygen react to form nitrogen monoxide which is oxidised by oxygen in air to nitrogen dioxide*

- Bases are water soluble alkalis. F
- *Alkalis are water soluble bases.*
- Basicity of metal oxides increases going down a group of the periodic table. T

## Acids in water

- Most hydrogen ions are protons. T
- $[H^+] \times [OH^-] = 10^{-14}$  at 25°C for aqueous acid solutions. T
- $[H^+] \times [OH^-] = 10^{-14}$  at 25°C for all aqueous salt solutions. T
- A solution with pH > 7 has a  $[H^+] > 10^{-7}$  F
- *A solution with pH > 7 has a  $[H^+] < 10^{-7}$*
- A weak acid solution has a lower concentration of ions compared with a strong acid solution of the same concentration of acid. T
- A concentrated acid solution always contains a high concentration of ions. F

*If the concentrated acid solution was of a very weak acid there would not be a high concentration of ions.*

- Acetic acid CH<sub>3</sub>COOH is a polyprotic acid. F
- *Acetic acid is monoprotic as only the H in COOH can ionise.*
- Some ions can be acids. T
- Basic substances are added to foods as preservatives. F
- *Acidic substances are added to foods as preservatives.*
- K<sub>2</sub>HPO<sub>4</sub> is called potassium dihydrogen phosphate. F
- *K<sub>2</sub>HPO<sub>4</sub> is called dipotassium hydrogen phosphate.*

## Acid concepts and titrations

- The Arrhenius acid definition is the same as the Bronsted-Lowry acid definition. F

*The Arrhenius definition applies to water only while the Bronsted-Lowry definition is not restricted to water.*

- Carbonic acid is the conjugate acid of the hydrogen carbonate ion. T
- Neutralisation is an exothermic reaction. T
- Buffers are usually made up of a weak acid and its conjugate base. T
- A volumetric flask must be dry before it can be used to make up a standard solution. F

*The volumetric flask can contain some solvent as solvent will be added to make up the solution.*

- A pipette filler should always be used to fill a pipette. T
- If titration of a strong acid and a strong base produces an insoluble salt the conductivity of the solution will be a minimum at the equivalence point. T
- A Lewis acid is an electron pair donor. F  
*A Lewis acid is an electron pair acceptor.*
- All Lewis acids are Bronsted-Lowry acids. F  
*Bronsted-Lowry acids need to contain H to act as proton donors; Lewis acids do not have to contain H as they are defined as electron pair acceptors.*
- 'Have pair (of electrons), will share' could define a Lewis base. T

## Esterification

- There can be twice as much hydrogen bonding between two alkanoic acid molecules compared with between two alkanol molecules. T
- The relatively high MPs of methanol and ethanol indicate effective hydrogen bonding in the solid state. T
- Boiling chips are used to catalyse esterification. F  
*Boiling chips are used to promote steady boiling; an acid is needed to catalyse esterification.*
- Ester hydrolysis is the reaction of an ester with hydrogen. F  
*Ester hydrolysis is the reaction of an ester with water to form an alkanoic acid and an alkanol.*
- When using a condenser for reflux the cooling water should flow out of the bottom. F

*Cooling water flows in the bottom of a condenser.*

- Esters cannot hydrogen bond and so have very low water solubility. T
- In a separating funnel most ester layers float on top of the aqueous layer. T
- A condenser is arranged vertically for refluxing and distillation. F  
*A condenser is arranged vertically for refluxing but on an angle in distillation.*
- Adding carbonate to a mixture containing organic acid and water molecules changes the acid to a more soluble ionic form. T
- Care should be taken when carrying out esterification using a flame because esters are flammable. T

## Mole calculations

- 1 In one mole of water there are:
  - a)  $1 \text{ mole} = 6 \times 10^{23}$  of the units making up water =  $6 \times 10^{23}$  water molecules.
  - b)  $6 \times 10^{23}$  O atoms as each H<sub>2</sub>O contains one O
  - c)  $2 \times 6 \times 10^{23} = 1.2 \times 10^{24}$  as each H<sub>2</sub>O contains two H atoms
  - d)  $3 \times 6 \times 10^{23} = 1.8 \times 10^{24}$  atoms as each H<sub>2</sub>O contains three atoms.
- 2 a) The molarity of 2 moles of CaCl<sub>2</sub> in 5 L of solution is:  
$$c = \frac{n}{V} = 2/5 = 0.4 \text{ M}$$
  
b) Each unit of CaCl<sub>2</sub> contains one Ca<sup>2+</sup> so concentration of Ca<sup>2+</sup> is the same as the concentration of CaCl<sub>2</sub> = 0.4 M  
c) Each unit of CaCl<sub>2</sub> contains two Cl<sup>-</sup> so concentration of Cl<sup>-</sup> is double the concentration of CaCl<sub>2</sub> =  $2 \times 0.4 = 0.8 \text{ M}$
- 3 The mass of:
  - a) 2.000 moles of copper Cu  
 $2.000 \times 63.55 = 127.1 \text{ g}$
  - b) 0.200 moles of sulfur dioxide SO<sub>2</sub>  
 $0.200 \times (32.07 + 2 \times 16.00) = 0.200 \times 64.07 = 12.8 \text{ g}$
  - c) 1.20 moles of anhydrous copper sulfate CuSO<sub>4</sub>  
 $1.20 \times (63.55 + 32.07 + 4 \times 16.00) = 1.2 \times 159.62 = 192 \text{ g}$

4 The gas volume of:

- a) 0.50 moles of nitrogen monoxide at 25°C and 101.3 kPa

$$0.50 \times 24.5 \text{ L} = 12 \text{ L of NO}$$

- b) 0.50 moles of nitrogen dioxide at 25°C and 101.3 kPa

$$0.50 \times 24.5 \text{ L} = 12 \text{ L of NO}_2$$

- c) 0.50 moles of carbon dioxide at 0°C and 101.3 kPa?

$$0.50 \times 22.4 \text{ L} = 11 \text{ L of CO}_2$$

5 The number of moles is:

- a)  $6 \times 10^{21}$  molecules of carbon dioxide CO<sub>2</sub>

$$n = \frac{N}{N_A} = 6 \times 10^{21} / 6 \times 10^{23} = 10^{-2} = 0.01$$

- b) there of solute in 600 mL of 0.15 molL<sup>-1</sup> solution

$$c = \frac{n}{V} \quad \text{thus } n = cV = 0.15 \text{ molL}^{-1} \times 0.6 \text{ L} = 0.090$$

- c) in 5.85 g of sodium chloride NaCl

$$n = \frac{m}{M} = 5.85 / (22.99 + 35.45) = 0.100$$

- d) 5.85 g of hydrogen gas at 273 K and 1.00 atm

$$n = \frac{m}{M} = 5.85 / (2 \times 1.008) = 2.90 \text{ mol} = 2.90 \times 22.4 \text{ L} = 65.0 \text{ L}$$

of H<sub>2</sub>

- e) 5.85 g of oxygen gas at 298 K and 101.3 kPa?

$$n = \frac{m}{M} = 5.85 / (2 \times 16.00) = 0.183 \text{ mol} = 0.183 \times 24.5 = 4.48 \text{ L}$$

of O<sub>2</sub>

6 a) How many grams of hydrogen chloride gas can be released by boiling 1.5 L of 0.60 M HCl(aq)?

$$c = \frac{n}{V}$$

$$\text{thus } n = cV = 0.60 \times 1.5 = 0.90 \text{ mol} = 0.90 \times (1.008 + 35.45) = 33 \text{ g}$$

- b) What volume is occupied by 5.1 g of ammonia at 0°C and 101.3 kPa pressure?

$$n = \frac{m}{M} = 5.1 / (14.01 + 3 \times 1.008) = 0.30 \text{ mol} = 0.30 \times 22.4 \text{ L}$$
$$= 6.7 \text{ L of NH}_3$$

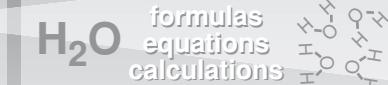
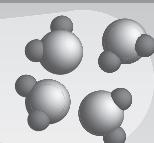
- c) What is the molarity of 100 g of ammonium chloride dissolved in 560 mL of water?

$$n = \frac{m}{M} = 100/(14.01 + [4 \times 1.008] + 35.45)$$
$$= 100/53.5 = 1.87 \text{ mol}$$

$$c = \frac{n}{V} = 1.87/0.560 = 3.33 \text{ M}$$

- d) What is the mass of  $5.0 \times 10^{22}$  sulfuric acid molecules?

$$n = \frac{N}{N_A} = 5.0 \times 10^{22}/6 \times 10^{23} = 0.083 \text{ mol}$$
$$= 0.083 \times (2 \times 1.008 + 32.07 + 4 \times 16.00) = 0.083 \times 98.086$$
$$= 8.2 \text{ g}$$

**MACRO**observe  
infer  
understand**SYMBOLIC** $H_2O$  formulas  
equations  
calculations**MICRO**particles  
energy  
interactions

# Exercises – Part 6

Exercises 6.1 to 6.2

Name: \_\_\_\_\_

## Exercise 6.1: Mole calculations

1 Calculate the number of:

- a) atoms in 3.00 moles of lithium metal

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- b) molecules in 4.0 moles of methane

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- c) ions in 5 moles of sodium chloride

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- d) ions in 5 moles of calcium chloride

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2 Calculate the molarity of 10 L of solution containing 5.85 g of NaCl.

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3 What volume of CO<sub>2</sub> gas is released at 25°C and 101.3 kPa when 10.0 g of CaCO<sub>3</sub> completely dissolves in hydrochloric acid?

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## **Exercise 6.2: Chemistry skills check list**

Now that you have completed half of the HSC course it is time for you to check that you have had opportunities to develop skills in:

- 1 planning investigations
- 2 conducting investigations
- 3 communicating information and understanding
- 4 developing scientific thinking and problem solving techniques
- 5 working individually and in teams.

Place a tick  to the right hand side of any skill you have covered in studying your chemistry course. Place a dash – to the right hand side if you cannot tick the skill.

### **1 Planning investigations**

Have you:

- a) carried out an open-ended investigation?

What was the title of the open-ended investigation you carried out?  
\_\_\_\_\_

- b) collected

i) qualitative data such as colour, change in appearance, and so on?

ii) quantitative data, that is, made measurements?

- c) identified the order of magnitude and uncertainty in a measurement?  
eg. a measurement of 984 g to the nearest g has an order of magnitude of  $10^3$  g and uncertainty of  $\pm 0.5$  g

- d) used and recorded the correct units for measurements?

- e) organised gathered information so that it can be efficiently analysed?  
eg. into a table or on a graph

- f) understood the difference between a dependent variable and an independent variable?

The independent variable is what you are interested in measuring.

The dependent variables are quantities that could affect the value of the independent variable measurement. For example, consider measuring the effect of temperature on colour of a gas mixture.

Colour is the independent variable and temperature, concentration, pressure, for example, would be dependent variables.

- g) identified variables eg. temperature, amounts of chemical, that need to be kept constant,  
developed strategies eg. carry out experiment at a constant temperature, to ensure variables are kept constant,  
demonstrated the use of a control eg. compared result between experiments that differ by one factor only?
- h) aimed for validity (leading to effective results and worthwhile conclusions) and reliable (trustworthy) data?
  - i) designed and reported on appropriate procedures?
  - j) predicted issues eg. safety, and acted accordingly eg. used safety procedures?
  - k) identified and set up appropriate equipment for investigations?
  - l) carried out a risk assessment of intended experimental procedures eg. by referring to a MSDS, reading a chemical label carefully, mentally thinking your way through a procedure and possible risks before physically carrying out the procedure.  
identified and addressed potential hazards eg. flammable liquid use near flame or electrical sparks?
  - m) identified technology that could be used eg. indicator or pH meter or pH probe connected to a cataloger, and determined its suitability?
  - n) recognise the difference between destructive and non-destructive testing of a material?  
eg. using an indicator solution or a pH meter to measure the pH of a solution  
eg. measuring the amount of  $\text{Na}^+\text{Cl}^-$  salt in water by titration with standard  $\text{Ag}^+$  solution precipitating  $\text{AgCl}$  or measuring the amount of salt by conductivity measurements.
  - o) analysed potentially different results of these two procedures?  
eg. consider if the results of adding an indicator (usually a solution of a weak acid and its conjugate base) to a solution could give a different result compared to a measurement made using a calibrated pH meter  
eg. consider whether other salts in the water sample could affect the reliability of the measurements of salt content.

**Could you satisfy HSC Course Outcome H11?**

*justifies the appropriateness of a particular investigation plan?*

## **2 Conducting investigations aims to produce valid and reliable data**

Have you:

- a) modified procedures and analysed the effect of these modifications?
- b) minimised hazards and waste? eg. by risk assessment and using small quantities of chemicals
- c) disposed of waste materials carefully and safely? eg. by considering whether an acid or base should be neutralised before disposal down a sink hole and how this can be done safely
- d) identified and used safe work practices? eg. whether protective eye wear or clothing is needed when working with a particular acid or concentration of acid
- e) used appropriate data collection techniques and used appropriate sensors? eg. use a thermometer and notebook to record temperature or use a temperature probe and a datalogger
- f) repeated trials of experiments as appropriate? eg. if the 'rough' titre and next two titres are not sufficiently close in value
- g) accessed information from a range of resources including the Internet?
- h) practised efficient data collection techniques eg. used search engines on the Internet?
- i) extracted information from graphs, tables, written and spoken material?
- j) summarised and collated (gathered together) information from a range of resources?
- k) identified practising male and female Australian scientists, areas in which they are working and information about their research eg. from TV programs, newspapers, magazines or internet sites such as [www.abc.net.au/science](http://www.abc.net.au/science) and [www.lmpc.edu.au/science](http://www.lmpc.edu.au/science)
- l) assessed (judged the value of) the accuracy of measurements and calculations and their relative importance?
- m) identified and applied mathematical formulas

$$\text{eg. density} = \frac{\text{mass}}{\text{volume}}$$

$$n = \frac{m}{M}$$

$$c = \frac{n}{V}$$

$$cV = \text{constant} \quad \text{or} \quad c_1V_1 = c_2V_2$$

$$\Delta H = m C \Delta T$$

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

$$[\text{H}^+] \times [\text{OH}^-] = 1.0 \times 10^{-14} \text{ at } 25^\circ\text{C}$$

$$\frac{c_a V_a}{a} = \frac{c_b V_b}{b}$$

- n) illustrated trends and patterns by selecting and using appropriate methods, eg. graphing and including computer assisted analysis eg. using a computer spreadsheet program?
- o) evaluated the relevance of information?
- p) assessed the reliability of information by considering information from various sources?
- q) compared accuracy of scientific information from mass media  
eg. radio, television, newspapers, magazines by comparison with similar information from scientific journals  
eg. Scientific American [www.sciam.com](http://www.sciam.com)  
New Scientist [www.newscientist.com](http://www.newscientist.com)  
Nature [www.nature.com](http://www.nature.com)

### Could you satisfy HSC Course Outcome H12?

*evaluates ways in which accuracy and reliability could be improved in investigations?*

### 3 Communicating information and understanding

Have you:

- a) used different text types. For example:
  - set out procedure steps in your report for the open-ended investigation?
  - offered explanations in answering questions?
  - argued a point of view (exposition)?
  - described an experiment using aim, method, results, conclusion?
- b) selected and used appropriate media. For example:
  - writing
  - drawing
  - photography
  - video?
- c) acknowledged sources of information?
- d) used symbols, formulas and appropriate units?

- e) used a variety of pictorial representations to show relationships and present information clearly and succinctly eg.
  - made molecular models
  - drawn chemical formulas
  - written chemical equations in symbols
  - used flow charting
  - used a key
  - used a table
  - used a Venn diagram?
- f) selected and drawn appropriate graphs?
- g) identified situations where use of curve of best fit rather than a straight line of best fit is appropriate on a graph?

**Could you satisfy HSC Course Outcome H13?**

*uses terminology and reporting styles appropriately and successfully to communicate information and understanding?*

**4 Developing scientific thinking and problem-solving techniques to draw valid conclusions from gathered data and information**

Have you analysed information to:

- a) identify trends, patterns and relationships as well as contradictions?  
eg. MP and BP of CH<sub>4</sub> NH<sub>3</sub> H<sub>2</sub>O HF Ne
- b) justify inferences (what you first think happens) and conclusions (what you are more certain happened)?
- c) identify and explain how data (observed information) supports or refutes an hypothesis (satisfactory explanation for a variety of observations), a prediction (what you think will happen) or a proposed solution to a problem?
- d) predict outcomes and generate plausible explanations related to observations?
- e) make and justify generalisations? eg. ‘like dissolves like’
- f) use models to explain and/or make predictions? eg. particle theory
- g) use cause and effect relationships to explain? eg. acid rain and the death of land plants or fish and shellfish in lakes
- h) identify examples where scientific ideas interconnect?  
eg. biogeochemical cycles and explaining the effects of acid rain

Have you solved problems by:

- a) identifying and explaining the nature of the problem?
- b) choosing a strategy to solve a problem?
- c) developing a range of possible solutions?
- d) evaluating the appropriateness of different strategies?

Have you used available evidence to:

- a) design and produce creative solutions to problems?
- b) propose coherent and logical ideas using scientific principles?
- c) apply critical thinking in considering predictions, hypotheses and results?
- d) formulate cause and effect relationships?

**Could you satisfy HSC Course Outcome H14?**

*assesses the validity of conclusions from gathered data and information?*

**5 Have you worked on some activities in chemistry:**

- **by yourself**
- **with another person or persons?**

**Could you satisfy HSC Course Outcome H15:**

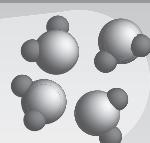
*explains why an investigation is best undertaken individually or by a team?*

**If you ticked most of these skills – congratulations!!!**

You are well on the way towards completing the HSC course chemistry skills (9.1 pp 44 – 46 *Chemistry Stage 6 Syllabus* © Board of Studies NSW, 1999).

When you have completed the last HSC core module *Chemical Monitoring and Management* check through this list again concentrating on the skills you have just put a – next to.



**MACRO**observe  
infer  
understand**SYMBOLIC** $H_2O$  formulas  
equations  
calculations**MICRO**particles  
energy  
interactions

# Student evaluation of module

We need your input! Can you please complete this short evaluation to provide us with information about this module. This information will help us to improve the design of these materials for future publications.

1 Name: \_\_\_\_\_

2 Location: \_\_\_\_\_

3 Did you find the information in the module easy to understand?  
\_\_\_\_\_

4 Were the instructions clear? \_\_\_\_\_

5 What did you most like learning about? Why?  
\_\_\_\_\_  
\_\_\_\_\_  
\_\_\_\_\_

6 Which sort of learning activity did you enjoy the most? Why?  
\_\_\_\_\_  
\_\_\_\_\_  
\_\_\_\_\_

7 Did you complete the module within 30 hours? (Please indicate the approximate length of time spent on the module.)  
\_\_\_\_\_  
\_\_\_\_\_

- 8 Do you have access to the appropriate resources? eg a computer, the internet, scientific equipment, chemicals, people that can provide information and help with understanding science

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Please return this information to your teacher, who will pass it along to the materials developers at OTEN – DE.

