

# Areas of Study

## AOS 1 Matter is composed of atoms

Chapter 1 Atomic structure  
The beginnings of chemistry  
Where do elements come from?  
Elements of life\*

## AOS 2 Structure, bonding and properties of materials

Chapter 2 Elements and compounds  
Atmospheres  
Metals by design  
Flavours and odours  
Crystals\*  
Chemistry and the home\*

## AOS 3 Chemical quantities

Chapter 3 Stoichiometry  
Chapter 4 Chemical solutions  
Chapter 5 Gases  
The air we breathe  
Consumer chemistry  
Water quality  
Soil chemistry\*

## AOS 4 Chemical reactions

Chapter 6 Precipitation reactions  
Chapter 7 Redox reactions  
Chapter 8 Acid–base reactions  
Chapter 9 Organic chemistry  
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Corrosion at a cost  
Polymers  
Soaps and detergents  
Mining in the North  
Chemicals from the sea  
Electrochemistry  
Food analysis  
Acids in the environment\*  
Metals and the modern lifestyle\*  
Chemistry in archeology\*

## AOS 5 Energy changes

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## AOS 6 Qualitative and quantitative techniques

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## AOS 7 Chemical equilibrium

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Biotechnology \*

\*These contexts are on the *eChemistry* second edition Student CD.

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## Contexts

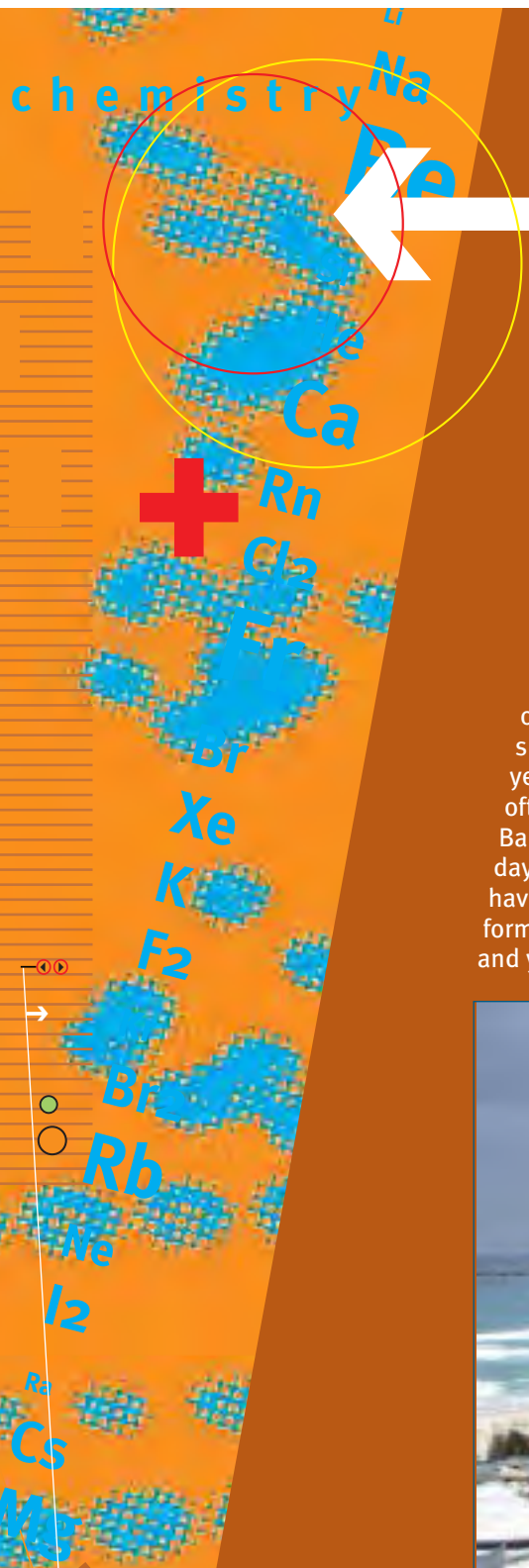
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# corrosion

at a cost

Have you ever considered living in paradise in a home by the beach? Many of us dream about it but the reality brings a new perspective. Have you ever considered just how many items around your home are made of metals?

When you live by the beach it becomes very obvious because anything made of metal soon shows signs of corrosion—light fittings, ceiling fans, steel door surrounds, door locks and window catches. Many electrical items only last a few years and are almost treated as disposable. Stereo systems and television sets often stop working after only a couple of years and there is little hope of repair. Backing up the data on your computer is essential since you may come home one day to find that it just won't start anymore. The family car is a similar story. You may have learned in the past of the wonders of aluminium and how a protective coating forms on its surface to stop corrosion. Beside the ocean even aluminium corrodes and you will see the affects of it as door frames and windows need replacing.



figure co.1 The joys of living beside the beach—sun, sand, saltwater and corrosion!



Metals are used for these types of items because they are strong, easily shaped and cost-effective compared to the alternatives. As we learn of the causes of corrosion and how it works we are better prepared to find ways of preventing it from happening.

## •• The corrosion process

Corrosion is the degradation of a metal as it reacts with other elements, loses electrons and forms other compounds. The process of corrosion is a redox reaction. Oxidation has become known as a process in which one element loses electrons to another element, not just oxygen.



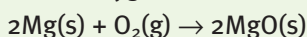
Section 7.1 **Oxidation and reduction** (page 338) and Section 7.2 **Redox reactions** (page 341) describe the background knowledge needed to understand corrosion reactions.

### Dry corrosion

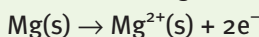
Dry corrosion is a direct reaction between a metal and oxygen in the air. It involves a metal reacting with oxygen, resulting in a metal oxide. It is a relatively slow process for most metals at room temperature.

At times you will have used strips of magnesium in science experiments. Before performing the experiment it is normally necessary to use sandpaper or steel wool to remove the dull powdery coating and expose fresh metal.

On the laboratory shelf magnesium ribbon will react with oxygen in the air to form an oxide coating.



During this process magnesium loses two electrons according to the half reaction:

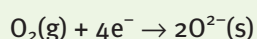


**figure co.2** The magnesium oxide coating needs to be removed when performing experiments with magnesium ribbon.



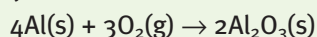
**figure co.3** Aluminium is used for window frames because it is strong, light and corrosion resistant.

These electrons are transferred to the oxygen atoms.



Magnesium has been oxidised and oxygen has been reduced, resulting in the production of magnesium oxide.

Aluminium undergoes a similar process to form a layer of aluminium oxide.



Aluminium oxide is a tough impenetrable coating that protects the metal underneath from further corrosion. If this coating is scratched, fresh aluminium oxide forms restoring the protection. However, if this process is continual under extreme conditions (for example near the ocean) aluminium will also eventually be completely oxidised.

### Wet corrosion

As implied by the name, wet corrosion occurs in the presence of water whether that be moisture in the air or total immersion in water. In the presence of water, metals don't have to directly lose their electrons to another element. The most common experience of corrosion for many people is that of rust, the corrosion of iron. The formation of rust is outlined in table co.1.

The red-brown substance  $\text{Fe}_2\text{O}_3 \cdot x\text{H}_2\text{O}$  is rust. Unlike the tough aluminium oxide coating that forms on aluminium, this iron(III) oxide (hydrated), flakes off the surface of iron exposing the metal underneath, which can then corrode.

A droplet of water on the surface of a piece of iron sets up a tiny electrochemical cell similar to those described in section 7.4 Ease of oxidation.



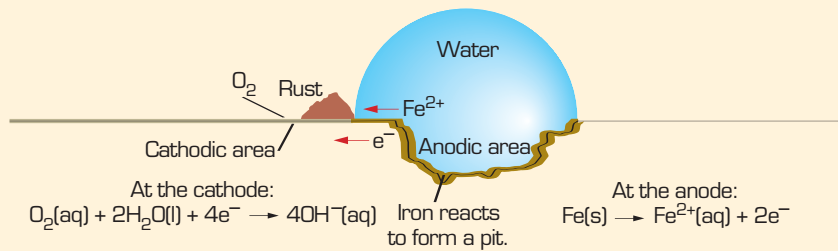
Section 7.4 **Ease of oxidation** describes electrochemical cells (page 348).



**table co. 1** The formation of rust from iron.

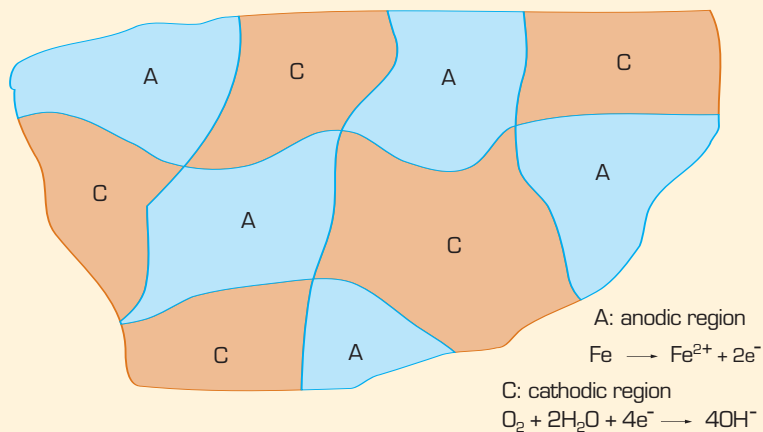
Step	Description	Reactions
1	The first step in the formation of rust is the loss of electrons by iron atoms. Oxygen dissolved in the water gains the electrons lost by the iron.	Anodic reaction $\text{Fe(s)} \rightarrow \text{Fe}^{2+}(\text{aq}) + 2\text{e}^{-}$ Cathodic reaction $\text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) + 4\text{e}^{-} \rightarrow 4\text{OH}^{-}(\text{aq})$ Overall reaction $2\text{Fe(s)} + \text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow 2\text{Fe}^{2+}(\text{aq}) + 4\text{OH}^{-}(\text{aq})$
2	Iron(II) ions react with hydroxide ions forming the green precipitate iron(II) hydroxide.	$\text{Fe}^{2+}(\text{aq}) + 2\text{OH}^{-}(\text{aq}) \rightarrow \text{Fe}(\text{OH})_2(\text{s})$
3	The iron(II) hydroxide in turn reacts with oxygen and water to produce iron(III) hydroxide, a red-brown precipitate.	$4\text{Fe}(\text{OH})_2(\text{s}) + \text{O}_2(\text{aq}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow 4\text{Fe}(\text{OH})_3(\text{s})$
4	The iron(III) hydroxide can lose water to form $\text{Fe}_2\text{O}_3 \cdot x\text{H}_2\text{O}$ , a red-brown substance.	For example $2\text{Fe}(\text{OH})_3(\text{s}) \rightarrow \text{Fe}_2\text{O}_3 \cdot 2\text{H}_2\text{O}(\text{s}) + \text{H}_2\text{O}(\text{l})$

**figure co.4** A drop of water on a piece of steel can create an electrochemical cell. Electrons flow from the site of oxidation to the site of reduction.



The edge of the droplet, where oxygen concentration is high, acts as the cathodic area of the electrochemical cell. The region beneath the water droplet, where the oxygen concentration is low, forms the anodic area. Electrons flow from the anodic area to the cathodic area as the initial redox reactions of the corrosion process takes place (figures co.4 and co.5).

**\*** Section 7.3 **Oxidation numbers and half equations** describes one method for writing balanced chemical reactions for corrosion (page 343).



**figure co.5** The surface of a piece of corroding steel develops anodic regions where oxidation occurs and cathodic regions where reduction occurs.

## Experiment co.1

### Observing corrosion

#### Purpose

To investigate how several different treatments affect the corrosion of iron.

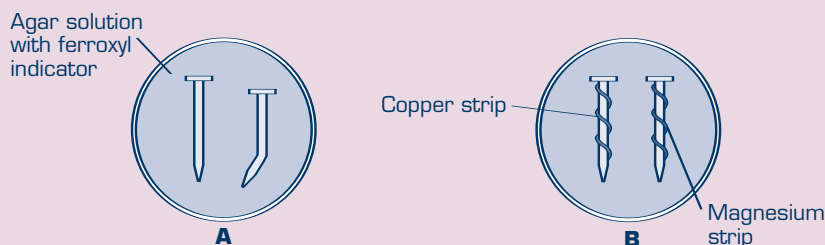


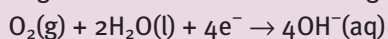
figure co.6

#### Method

- 1 Clean all nails with emery paper or steel wool.
- 2 Treat the nails as shown in figure co.6.
- 3 Place the nails in the Petri dishes.
- 4 Pour warm agar solution with ferroxyl indicator over the nails until they are covered.
- 5 Allow the agar to cool then cover with a lid, and label.
- 6 Draw diagrams of the nails and use labels and sketches to note any changes observed after:
  - a 20 to 30 minutes
  - b 24 hours
  - c 1 week.

#### Some information

**Pink areas:** Phenolphthalein indicates the presence of hydroxide ions. This is the region where reduction is taking place according to the equation



**Blue areas:** Iron is oxidised to  $\text{Fe}^{2+}$  and these ions react with the hexacyanoferrate(III) ions to produce the insoluble blue substance iron(II) hexacyanoferrate(III).

#### Discussion

- 1 a Describe the observable changes that have taken place to indicate that the iron nail has corroded.  
b Identify the ions present in these areas.
- 2 Observe Petri dish A. Which regions of the iron nail have corroded the most? Give an explanation.
- 3 Observe Petri dish B. Describe the effects of having the nails in contact with each metal. Include half equations to show what happens at the anode and cathode.
- 4 Describe how corrosion has been accelerated or prevented by the conditions in this experiment.

#### materials

- (per group)
- 4 iron nails
  - 2 Petri dishes with lids
  - emery paper or steel wool
  - 250 mL agar solution with ferroxyl indicator (10 g sodium chloride, 1 g potassium ferricyanide, 1 g phenolphthalein, 500 mL distilled water)
  - 2 mm strips of copper and magnesium

#### safety

- Wear safety glasses and a laboratory coat for this experiment.
- Phenolphthalein is a powerful laxative. Do not ingest.
- Potassium ferricyanide is slightly toxic if ingested and a skin irritant.
- Take care when transferring hot agar.
- Wash hands thoroughly.



## ❖ The rate of corrosion

The conditions in which a metal is placed determine the rate of corrosion. Wet corrosion is much faster than dry corrosion because the metal can lose its electrons to form ions in water and doesn't have to lose electrons directly to another element. Cars used in inland country areas tend to show fewer signs of corrosion than those in coastal areas. By the ocean, dissolved salt tends to accelerate the rate of corrosion but other impurities such as acidic pollutants from the air also cause corrosion to occur more rapidly.

In an effort to counteract the effects of sea water, pieces of zinc are attached below the water line of steel-hulled boats. This zinc acts as a sacrificial anode—the zinc is oxidised and the steel is protected. An electrochemical cell has been constructed with the zinc as the anode and the steel as the cathode. When a more reactive metal is attached to a less reactive metal the less reactive one is protected against corrosion. Boat builders need to be careful of the screws, bolts and other metal fittings they use. If a less reactive metal is attached to the steel, corrosion of the steel will be accelerated.

Some impurities in metals can cause them to corrode more rapidly. Other impurities, such as those added to produce various alloys, can make metals extremely resistant to corrosion. Stainless steel (iron combined with small amounts of molybdenum, manganese, nickel and chromium) is valued for its corrosion resistance.



**figure co.7** This 1500-year-old pillar in Delhi shows little signs of rust due to the high purity of iron from which it is made.



**figure co.8** Stainless steel sinks rarely show any signs of corrosion despite being in almost constant contact with water.

One way of preventing corrosion of a metal is to coat it in a layer that is impermeable to oxygen. The paint on a car is one example of this. If the paint on a car is chipped, a cell can be formed similar to the drop of water described earlier. The exposed area is the cathodic area and the corrosion actually occurs at an anodic area below a section of undamaged paint. Food cans are steel coated with a thin layer of tin. The tin protects the steel from oxygen in the air, but if the tin is scratched it actually accelerates corrosion as it is a less reactive metal than the steel.

If a metal has been stressed by bending, working or a collision it is more susceptible to corrosion. If cars have been left out in a hail storm, they are often sold cheaply by the dealers to panel beaters who repair them and resell them. Even though they may be competently repaired, these cars are susceptible to corrosion at the previously dented sites.

Table co.2 describes the main conditions under which corrosion is accelerated and inhibited.



**figure co.9** The zinc coating of galvanised steel house frames protects the steel in two ways: it forms a barrier that prevents oxygen from reaching the surface of the steel and also it is oxidised in preference to the steel due to its greater reactivity.



**table co.2** Preventing and accelerating corrosion.

Corrosion acceleration	Corrosion prevention
The presence of moisture allows ions to form and react together.	The reactions of corrosion occur at a greater rate in solution. Removing moisture can slow or even stop corrosion.
Impurities in water such as salt and acidic pollutants accelerate corrosion.	A sacrificial anode can be used to minimise corrosion. A more reactive metal attached to a less reactive one protects the less reactive metal.
Some impurities within a metal can cause it to corrode faster.	Alloys contain beneficial impurities. Some alloys of steel are highly resistant to corrosion due to the metals with which they are combined.
When the surface protection on a metal (such as paint on a car) is chipped or damaged a cell is formed as in the water drop of figure co.4. Rust begins to form below the paint surface adjacent to the chip.	Paint and other protective coatings keep oxygen away from the metallic surface. The zinc surface of galvanised iron simultaneously protects the surface from oxygen and acts as a sacrificial anode.
The sites of stress on a metal such as those caused by bending or hammering are prone to corrosion.	Car panels after accidents and iron roofs after hail storms are often replaced as even after repair they are still susceptible to corrosion.

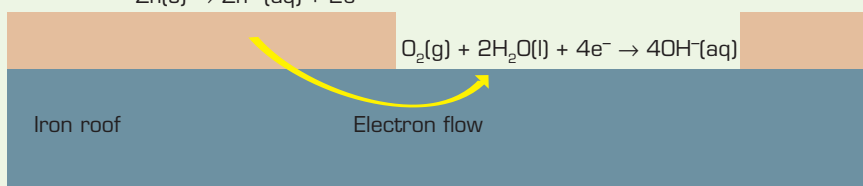
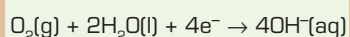
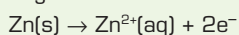


**figure co.10** The Sydney Harbour Bridge is a 53 000-tonne steel structure protected from corrosion by paint. This is a continuous process that uses 90 000 litres of paint per coat.

**figure co.11** (a) Galvanised iron roofs are common throughout Australia. (b) The iron is protected as the zinc coating gives up its electrons.



(b) Anodic region Cathodic region



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Chemical reactions





## Experimental investigations

There are a variety of ways of approaching an investigation into corrosion. Experiment co.1 may give you a starting point. There are many variables that you can change to observe the extent of corrosion or the rate at which it occurs under different conditions. Some of these variables are listed below.

- 1 The metal**—iron nails are easy to obtain and fairly uniform in size and shape but you may have access to other forms of iron such as lathe turnings from an engineering works. Samples of metals such as copper or aluminium may also be available.
- 2 Atmospheric conditions**—samples of a metal could be observed in water, humid air and dry air. You may like to observe a sample that is partially immersed in water to see the effects.
- 3 Impurities in water**—solutions containing different substances could be compared for their ability to accelerate corrosion, for example salt water, acids or bases. You could also compare the effects of different concentrations of these.
- 4 Oxygen content of water**—dissolved oxygen can be removed from water by boiling and kept out with a layer of oil. A layer of oil will also maintain the amount of oxygen in water. The effects of different concentrations of oxygen on corrosion may be observed.
- 5 Barrier protection**—you might like to test the ability of different substances to protect a metal. Different commercial paints could be compared with oils or greases.
- 6 Metal combinations**—investigate the effects of combining metals on the rate of corrosion. You might like to extend the investigation of this context using other metals besides copper, iron and magnesium.
- 7 Applied voltage**—this is another method sometimes used to protect metals from corrosion. You could test whether it is the positive or negative electrode of a power supply that protects the metal. A comparison of different voltages may also prove interesting.
- 8 Physical treatment**—a variety of metals could be treated (or mistreated) by hammering, scratching, rolling or some other method that you devise. See how these treatments affect the ability of the metal to resist corrosion.
- 9** You may even decide to combine several of these treatments together; for example, you could physically treat samples of metals by hammering and then paint them with rustproof paint. Library and Internet research could also provide ideas for further investigation. There may be a local business that works in metals and consulting with their staff may prove fruitful.



# Soaps and detergents

Personal hygiene means keeping our bodies and clothes clean. To keep clean you need to regularly wash to remove sweat, excess oils and dirt from yourself and the clothes you wear. Cleaning is important for the long-term health of your skin. It reduces bacteria and associated odours, keeps pore ducts open and removes dried-up surface skin cells.

The practice of good personal hygiene includes:

- frequently washing your hands
- having a daily shower or bath
- regularly washing your hair
- daily brushing and flossing your teeth
- using a deodorant to reduce body odour
- wearing clean clothes, especially a daily change of socks and underwear
- keeping your fingernails and toenails trimmed.

Chemistry helps us to maximise the benefits of good personal hygiene, to look good and to feel good.

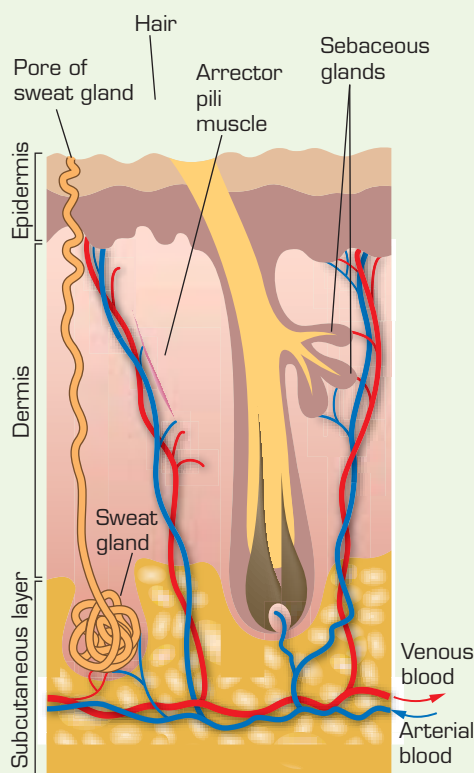


figure sd.1 Wash your hands often.

A major part of our body is the skin. In fact the skin is the largest organ in the body, forming a complete external cover.

Glands in our skin secrete oils and sweat. These materials are important. The oils keep our skin flexible, while sweat helps us to maintain a constant body temperature. Sweat is mainly water with about 1% sodium chloride (salt). Sweat glands concentrated in some areas of the body, for example the armpit, secrete oils as well as sweat. Bacteria attack these oils and can cause offensive body odour. To reduce bacterial action and body odour wash frequently. This cleans the sweat, excess oils and dirt from the skin.

**figure sd.2** To keep your skin healthy eat properly, keep clean and avoid getting too much sun.



When you spend too long in the swimming pool or bathtub your skin becomes wrinkled. Have you ever wondered why? The outer layer of your skin, the epidermis, produces an oily substance called sebum that keeps water out. If you spend too long in the water most of the sebum is washed off. The epidermis then starts to absorb water and swells, causing the skin to wrinkle.

## •• Surfactants

Cleaning agents—soaps and synthetic substitutes—were collectively called detergents. After the 1950s the synthetic cleaning agents became more widely used and people referred to them as ‘detergents’ as distinct from ‘soap’. Cleaning agents in the bathroom include shampoos and soaps. In the kitchen and laundry you will find dishwashing liquid and washing powders.

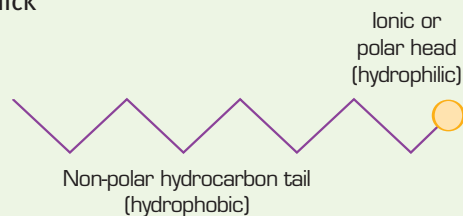
Cleaning agents remove dirt, oils and grease by dissolving them. Their key ingredient is a surfactant—a surface active agent.

Surfactant molecules have both fat-soluble and water-soluble parts. This is shown in figure sd.3. The fat-soluble part of the molecule is referred to as hydrophobic (water-hating). This part is non-polar and can also be described as lipophilic (fat-loving). It sticks to oil and dirt. The water-soluble part of the molecule is hydrophilic (water-loving) and sticks to the water molecules. The molecule’s dual nature allows the almost water-insoluble dirt and grime to be pulled away from the surface being cleaned and carried away as small particles in the water (see figure sd.4). Even though washing-up detergents and shampoos work in the same way, it is not advisable to wash your hair with dishwashing liquid. The dishwashing liquid is formulated to remove thick grime from pots and pans, not to gently clean your hair.



Section 2.3 Covalent bonding describes intermolecular forces and polar bonding (page 233).

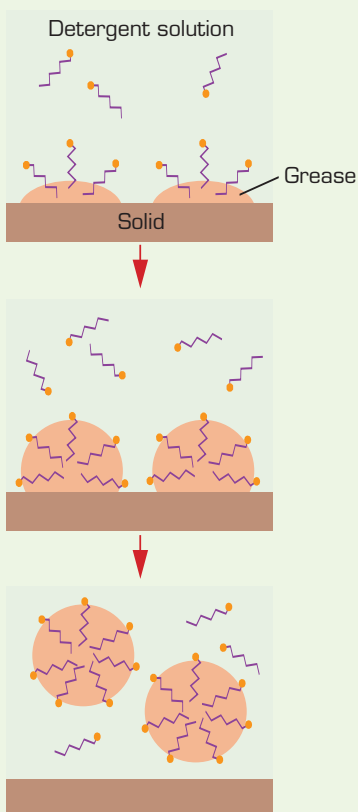
A surfactant is a surface active agent and is the key ingredient in detergents.



**figure sd.3**  $\text{CH}_3(\text{CH}_2)_{16}\text{COO}^-\text{Na}^+$ , sodium stearate soap. Surfactant molecules have a hydrophilic head and a hydrophobic tail.





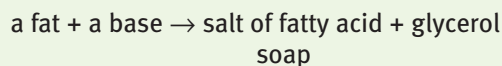


**figure sd.4** Surfactants in action. The fat-soluble part of the molecule adheres to oil and dirt, while the water-soluble part of the molecule adheres to the water molecules. Agitation causes the grime to be cleaned from the surface and carried away in the water.

## Soap making

Throughout history, soap has been valued for its ability to clean clothes and cooking utensils, and for bathing. It is no wonder that soap-making is one of the oldest chemical manufacturing processes. In the eighteenth century soap making was a profitable industry in Italy.

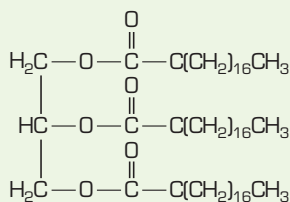
The cheapest and oldest way of making soap is by blending and boiling the basic ingredients of animal fat or a vegetable oil with water and sodium hydroxide. The chemical change is called saponification.



**figure sd.5** Soaps can be scented with herbs and fragrant oils to enhance their appeal.

Plants and animals use fats and oils as energy-storage compounds. Fats and oils are organic molecules called esters and most are triesters (see figure sd.6). They belong to the group of compounds called lipids. Lipids are soluble in organic solvents but insoluble in water. This means that they are difficult to wash from our clothing and our body without the aid of soaps and detergents. The marvels of chemistry allow the use of the material causing the problem to produce the product needed to solve the problem.

Soap is the end product of the saponification of a fat molecule. From figure sd.6 you can see that the structure of a fat is quite complex. The earlier equation for saponification showed no molecular formulae. To help you understand soap-making at the molecular level, examine the structures in the general reaction below and the sodium stearate example.

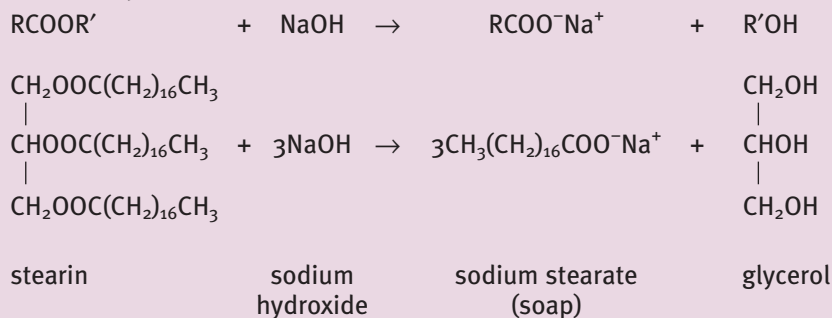


**figure sd.6** Beef tallow contains the fat **stearin**, triglyceride glyceryl tristearate (systematic name: propane-1,2,3-triyl trioctadecanoate) and makes the most common of soap, sodium stearate.



Section 8.3 **Defining acids and bases** describes the chemical features of acids and bases (page 370).

General equation:



Sodium stearate soap is the sodium salt of stearic acid, found mainly in animal fat. A list of fatty acids and their sources is provided in table sd.1.

**table sd. 1** Some common fatty acids.

Found in	Common name	Structure
coconut oil	lauric acid	$\text{CH}_3(\text{CH}_2)_{10}\text{COOH}$
palm oil and most fats	palmitic acid	$\text{CH}_3(\text{CH}_2)_{14}\text{COOH}$
mainly in animal fat	stearic acid	$\text{CH}_3(\text{CH}_2)_{16}\text{COOH}$
olive oil and most fats	oleic acid	$\text{CH}_3(\text{CH}_2)_7\text{CH}=\text{CH}(\text{CH}_2)_7\text{COOH}$

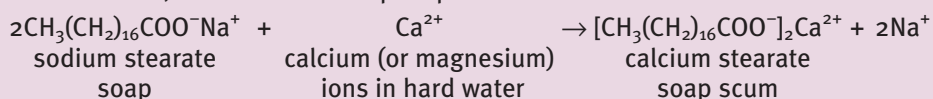
**\*** Section 9.5 **Reactions of organic acids** describes the structure and properties of esters (page 401). Section 2.5 **Organic compounds** describes carboxylic acids (page 246).



**figure sd. 7** Soap will only lather up in soft water.

## •• The decline of soaps

Soaps do not lather in all types of water. Differences between water sources are due mainly to the type and quantity of dissolved ions. Sea water, for example, contains a very large concentration of sodium and chloride ions and a relatively high concentration of calcium and magnesium ions. It is very difficult to get soap to lather and hence work properly in sea water. Water in which soap does not lather is called hard water. The calcium and magnesium ions in hard water combine with the long hydrocarbon part of soap to form an insoluble solid, or scum. This is a precipitation reaction.



**\*** Section 6.2 **Precipitation reactions** describes these types of reactions (page 331).



Soap scum is an insoluble deposit that forms in hard water rendering a soap ineffective.



## Experiment sd.1

### The saponification reaction

#### Introduction

You will set up reflux apparatus and perform a saponification reaction using olive oil and solid sodium hydroxide, then test the prepared soap for its emulsifying properties.

On completion of this investigation, you will be able to:

- set up reflux apparatus
- identify processes occurring in the reaction flask
- demonstrate the emulsifying properties of produced soap.

#### Procedure

- 1 Set up the quick-fit reflux apparatus as shown in figure sd.8. To the reaction flask add 5 mL olive oil, 10 g sodium hydroxide pellets and 10 mL water.
- 2 Add a boiling chip to the mixture and reflux until the oil layer is no longer visible.
- 3 Remove the condenser and pour the mixture into a beaker containing 100 mL ice-cold brine.
- 4 Filter off the solid residue formed through the cheesecloth. (Filter paper is too slow.)
- 5 Wash the residue in a small amount of ice-cold distilled water.
- 6 Place 2 drops of olive oil into two separate test-tubes containing 10 mL water. To one of them add a piece of the soap produced in the reaction.
- 7 Set up a third test-tube containing 10 mL water and a small amount of soap.
- 8 Shake all three test-tubes vigorously and compare for lathering ability and emulsifying properties.

#### Discussion

- 1 To what group of substances does the solid product belong?
- 2 The principal acid derived from olive oil is oleic acid ( $C_{17}H_{33}COOH$ ). Write a balanced equation for the saponification of olive oil (assuming this is the only acid present).
- 3 Briefly explain the function of the sodium chloride solution in this investigation.
- 4 What evidence was there to show that the soap has emulsifying properties?

#### Extension

Compare the conditions in the laboratory preparation of soap with an industrial soap-making process.

#### Disposal of waste

All water-soluble chemicals can be washed down the sink. Reuse olive oil or place on perlite, wrap up in paper and dispose of in the garbage bin.

#### materials

- 5 mL olive oil
- 10 g sodium hydroxide (NaOH)
- 50 mL ice-cold brine (saturated NaCl solution)
- quick-fit apparatus with 50 mL flask and condenser
- heating mantle
- cheesecloth or equivalent
- 400 mL beaker
- test-tubes
- boiling chips

#### safety

- Sodium hydroxide solid is highly corrosive to skin and eyes.
- Wear eye and skin protection.

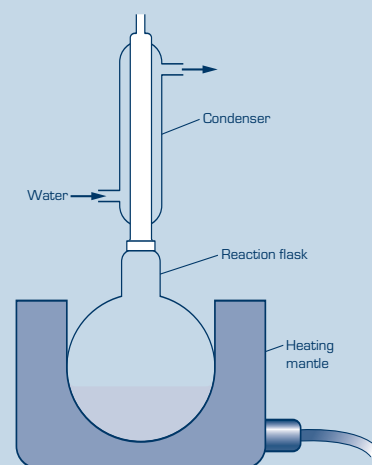






figure sd.8 Reflux apparatus set-up





**table sd.2** Structure and common uses of soap and the three different types of synthetic detergents.

Type	Non-polar portion	Polar portion	Common uses
soap		$\begin{array}{l} \text{O} \\ \parallel \\ \text{C} \\   \\ \text{O}^- \text{Na}^+ \end{array}$	personal hygiene
anionic		$\begin{array}{l} \text{O} \\ \parallel \\ \text{O} - \text{S} - \text{O}^- \text{Na}^+ \\ \parallel \\ \text{O} \end{array}$	laundry detergents and dishwashing liquids
cationic		$\begin{array}{l} \text{CH}_3 \\   \\ \text{N}^+ \text{CH}_3 \text{Br}^- \\   \\ \text{CH}_3 \end{array}$	hair conditioner, fabric softener
non-ionic		$\begin{array}{l} \text{CH}_2 - \text{OH} \\   \\ \text{C} - \text{CH}_2 - \text{OH} \\   \\ \text{CH}_2 - \text{OH} \end{array}$	wetting agents (paint, pesticides), automatic dishwasher powders, cosmetics



## Experimental investigations

### 1 Soaps from different oils

Using the method described in experiment sd.1 The saponification reaction, you could produce a range of soaps from different types of oil. For example, in place of olive oil you could try peanut, canola or sunflower oil. Test and compare the properties and effectiveness of each of the different soaps produced. The following investigations may assist you in devising these tests.

### 2 Action of a detergent on grease

The following procedure can be used to analyse the effectiveness of a detergent. You may like to collect a range of different brands and types of detergents and use this method to make a comparison.

#### Method

- Place 2 small lumps of margarine on opposite ends of a white tile. Sprinkle a small amount of red, oil-soluble dye on each and mix with a stirring rod until each lump of margarine turns red.
- Add 2 drops of water to one lump and mix. Add 2 drops of detergent solution to the other lump and mix. Record your observations.
- Wash the margarine from the surface of the tile and record your observations.

#### Discussion

- Account for the differences you observed in the behaviour of the water and the detergent solution when they were mixed with margarine.



- b** Use labelled diagrams to explain how the detergent removes the margarine from the surface of the tile.

### 3 Soap and synthetic detergent in hard and soft water

The following procedure could be used to compare the effectiveness of soaps and detergents in hard and soft water. You may decide to use this procedure to compare different brands of soaps and detergents or to compare the quality of water from different sources.

#### Method

- a** Pour 5 mL of deionised water into 2 test tubes and shake well. Record your observations.
- b** Add 0.5 mL of soap solution to the first test tube and shake well. Record your observations.
- c** Add 0.5 mL of synthetic detergent solution to the second test tube and shake well. Record your observations.

#### Discussion

- a** Explain why synthetic detergents are preferable to soap for washing clothes in regions that have hard water.
- b** When soap is used in hard water a scum may form. Write an equation for a reaction which may occur to produce this scum.

### 4 Water hardness

Water hardness can be classified using the equivalent quantity of calcium carbonate, as shown in table sd.3.

**table sd.3** Classification of water hardness.

Concentration of $\text{CaCO}_3$ (ppm)	Classification
less than 70	soft
75–150	moderately hard
150–300	hard
more than 300	very hard

#### Method

- a** Plan and conduct an investigation to determine the hardness of a water sample.
- b** Write an article for a popular magazine that explains to the reader the differences between saturated and unsaturated fats, and between butter and margarine. Your article will also discuss these materials in terms of healthy eating.
- c** Examine the environmental impact of soaps and detergents and relate community concerns and scientific research to changes made in the production and use of detergents.





## Consider this

Use the information presented in the context and related chapters to answer the following questions.

- 1 A soap contains the salt potassium palmitate,  $\text{CH}_3(\text{CH}_2)_{10}\text{COO}^-\text{K}^+$ .
  - a Identify the reactants used to make this soap.
  - b Identify the lipophilic end of the molecule.
  - c Identify the hydrophilic end of the molecule.
  - d Describe with the aid of an equation how the salt dissolves when placed in water.
  - e Describe its cleaning actions using appropriate scientific language and diagrams.
- 2 Discuss similarities and differences between soaps and detergents. Present your information as a PowerPoint presentation or as a poster designed to inform the general public.
- 3 Sea water is a solution of salts and contains more than 70 dissolved elements. Salinity is the amount (in grams) of total dissolved salts present in 1 kilogram of water. Normal sea water has a salinity of 35 grams per kilogram (or litre) of water, which can also be expressed as 35 000 ppm (parts per million).
  - a Table sd.4 lists the six most abundant ions in sea water. Complete the table by calculating (where required):
    - the percentage mass  $\left(\frac{\text{mass of ion}}{\text{mass of solution}} \times 100\%\right)$
    - parts per million  $\left(\frac{\text{mass of ion in mg}}{\text{mass of solution in kg}}\right)$

**table sd.4** Concentrations of the six most abundant ions in sea water.

Ion	% mass in 1 kg of sea water	ppm
chloride ( $\text{Cl}^-$ )	55.0	19 250
sodium ( $\text{Na}^+$ )	30.6	
sulfate ( $\text{SO}_4^{2-}$ )	7.7	
magnesium ( $\text{Mg}^{2+}$ )		1291
calcium ( $\text{Ca}^{2+}$ )	1.2	
potassium ( $\text{K}^+$ )		385

- b Calculate the mass of soap that would be precipitated as scum by 1 litre of normal sea water.



## Further investigations

Research the types of hard water and methods of water softening. Present your findings as either a poster or a PowerPoint presentation.

The international standard for seawater composition is determined using sea water from Wormly in southern England.

# 8

## Acid-base reactions

### 8.1 Introducing acids and bases

**Acids** and **bases** are very common substances. They are widely used throughout industry and can be found in most households. Acids can be found in rain as a byproduct of air pollution and they can be used to clean concrete and flavour foods. They are even produced by our muscles when our muscles fatigue. Bases are used to make powerful cleaning products. They can also be found in cement and used to treat indigestion. Bases that are soluble in water are referred to as **alkalis**.

There are many acids and bases that are commonly found in and around your home (see tables 8.1 and 8.2).

An alkali is a base that dissolves in water to form an alkaline solution.

**table 8.1** Common acids and some everyday uses.

Name	Formula	Uses
hydrochloric acid	HCl	present in stomach acid; used as a cleaning agent for brickwork
sulfuric acid	H <sub>2</sub> SO <sub>4</sub>	one of the most common chemicals manufactured; used in car batteries, paints, fertilisers and detergents
nitric acid	HNO <sub>3</sub>	used in the manufacture of fertilisers, dyes and explosives
ethanoic acid (acetic acid)	CH <sub>3</sub> COOH	key ingredient in vinegar; used as a preservative
carbonic acid	H <sub>2</sub> CO <sub>3</sub>	used in carbonated soft drinks and beer
phosphoric acid	H <sub>3</sub> PO <sub>4</sub>	used in some soft drinks and the manufacture of fertilisers
citric acid	C <sub>6</sub> H <sub>8</sub> O <sub>7</sub>	occurs in citrus fruits
lactic acid	C <sub>3</sub> H <sub>6</sub> O <sub>3</sub>	occurs in milk products; produced in muscles during strenuous exercise
ascorbic acid	C <sub>6</sub> H <sub>8</sub> O <sub>7</sub>	vitamin C—found in citrus fruits
formic acid (methanoic acid)	HCOOH	gives the bite of the green ant its sting



table 8.2 Common bases and some everyday uses.

Name	Formula	Uses
sodium hydroxide (caustic soda)	NaOH	used in drain and oven cleaners; soap making
ammonia	NH <sub>3</sub>	used in household cleaners, fertilisers and explosives
calcium hydroxide	Ca(OH) <sub>2</sub>	found in cement and mortar; in garden lime to adjust soil pH
magnesium hydroxide	Mg(OH) <sub>2</sub>	key ingredient in some antacids such as milk of magnesia—to overcome indigestion
sodium carbonate	Na <sub>2</sub> CO <sub>3</sub>	used in manufacture of washing powder and glass

## Properties of acids and bases

Acids and bases have characteristic properties, some of which are listed in table 8.3. Some properties are common to both acids and bases.

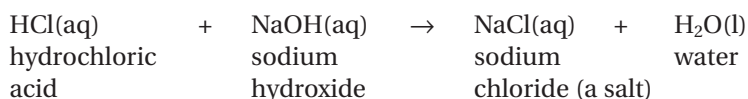
table 8.3 Properties of acids and bases.

Properties of acids	Properties of bases
sour tasting	bitter tasting
cause indicators to change colour (e.g. turn litmus red)	cause indicators to change colour (e.g. turn litmus blue)
react with bases to produce a salt and water	react with acids to produce a salt and water
can be strong or weak	can be strong or weak
can be dilute or concentrated	can be dilute or concentrated
can be corrosive	can be corrosive

## Neutralisation reactions

Acids and bases can be defined in many different ways. In simple terms, an acid is a substance that can neutralise a base, and a base is a substance that can neutralise an acid. When an acid and a base neutralise each other they always produce a salt and water.

For example:



This is called a **neutralisation reaction**.

A neutralised solution is neither an acid nor a base and consists of a salt and water.



A **salt** is the compound formed when some or all of the hydrogen in an acid is replaced by a metal ion.



The neutralisation reaction can be summarised by the word equation:  
acid + base → salt + water

### Questions

- 1 Define an acid in simple terms.
- 2 Describe what the term neutralisation means when referring to acids and bases.
- 3 Give two examples of where you might find a base in your home.
- 4 Identify four properties acids and bases have in common.
- 5 From the neutralisation reaction described below, write a chemical equation.  
Calcium carbonate and water was produced when carbonic acid reacted with calcium hydroxide.



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Acid–base reactions



## 8.2 Reactions involving acids

Acids react with a number of elements and compounds in a highly predictable way. Some of the characteristic reactions of acids are described in this section.

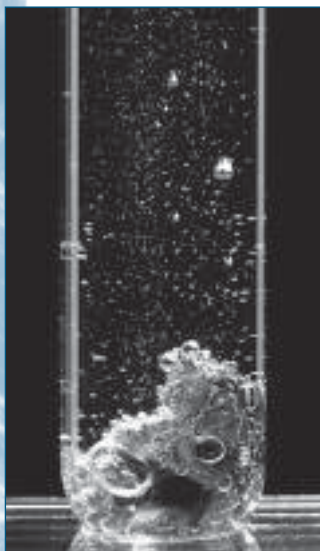
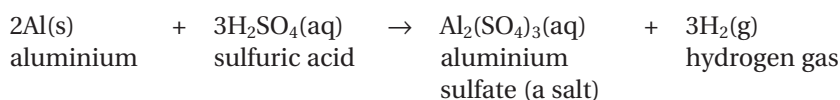
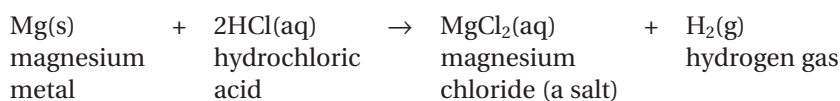
### Acids and metals

The position of hydrogen in the electrochemical series (table 7.2) indicates that acids will react with the highly reactive alkali metals (such as lithium and sodium), and even some less reactive metals (such as tin and lead) will react with acids.

When a metal reacts with an acid it forms a salt and hydrogen gas.

metal + acid  $\rightarrow$  salt + hydrogen gas

For example:



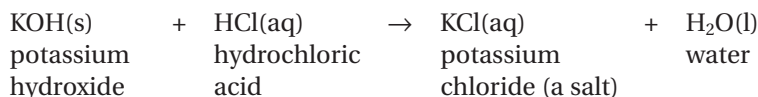
**figure 8.1** Hydrogen gas is produced from the reaction of some metals with dilute acid.

### Metal hydroxides and acids

Metal hydroxides react with acids. A salt and water are also the products in this case.

metal hydroxide + acid  $\rightarrow$  salt + water

For example:

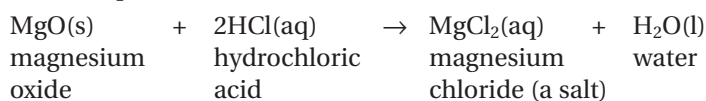


### Metal oxides and acids

When an oxidised form of some metals is placed in an acid it will react to produce a salt and water. The oxide behaves as a base.

metal oxide + acid  $\rightarrow$  salt + water

For example:

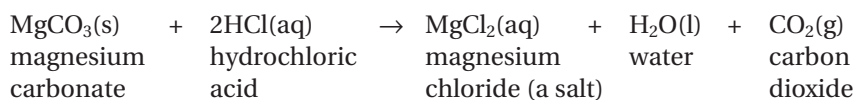


### Metal carbonates and acids

Carbonated forms of metals also react to form a salt and water. Carbon dioxide gas is an additional product of this particular reaction.

metal carbonate + acid  $\rightarrow$  salt + water + carbon dioxide gas

For example:



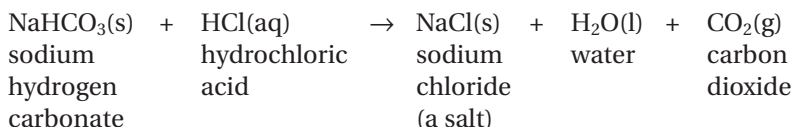
**figure 8.2** Effervescence in this antacid preparation is due to hydrogen ions and hydrogen carbonate ions reacting together to produce bubbles of carbon dioxide.

## Metal hydrogen carbonates and acids

Metal hydrogen carbonates react with acids to form a salt, water and carbon dioxide.

metal hydrogen carbonate + acid  $\rightarrow$  salt + water + carbon dioxide

For example:

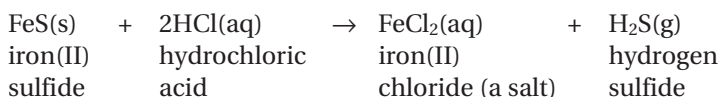


## Metal sulfides and acids

Another characteristic reaction of acids is their response to metal sulfides. Most metal sulfides will react with acids to form a salt and hydrogen sulfide.

metal sulfide + acid  $\rightarrow$  salt + hydrogen sulfide

For example:



Metal sulfides that have very low solubilities will not react in this way.

## Experiment 8.1 Reactions of hydrochloric acid

### Purpose

To investigate some common reactions of hydrochloric acid.

### Method

Note: A pH meter can be used as an alternative to universal indicator for this experiment. In this case, it may be necessary to replace test tubes with small beakers.

- 1 Add one drop of universal indicator to a test tube containing 20 drops of 0.1 M HCl. Add 0.1 M NaOH drop by drop and record any colour changes. Explain your observations and write an equation for the reaction that has occurred.
- 2 Carry out each of the following procedures and record your observations.
  - a Add about 1 mL of 1 M HCl to a test tube containing a small amount of solid  $\text{CaCO}_3$ . Test the gas evolved with lime water.
  - b Add about 5 mL of 1 M HCl to a small amount of solid CuO in a test tube. Shake the test tube and allow it to stand for 5 minutes.
  - c Add about 10 mL of 0.1 M HCl to a test tube containing a 1 cm strip of magnesium. Use an inverted test tube to collect the gas evolved and test it with a lighted taper.
  - d Add 1 mL of 0.1 M HCl to a small amount of solid  $\text{NaHCO}_3$  in a test tube. Test the gas evolved with lime water.
- 3 Summarise the reactions of hydrochloric acid observed in this experiment.

### Discussion

- 1 Write a full balanced equation for each reaction.
- 2 Write a balanced ionic equation for each reaction. Indicate whether the reaction is an acid–base reaction.

### materials

- beaker
- 0.1 M HCl
- 1 M HCl
- 0.1 M NaOH
- lime water
- universal indicator and pH chart or pH meter
- solid samples of  $\text{CaCO}_3$ , CuO,  $\text{NaHCO}_3$ , Mg
- test tubes/small beakers
- spatula
- teat pipette

### safety

- Wear gloves, safety glasses and a laboratory coat for this experiment.
- HCl is very corrosive and HCl vapour irritates the respiratory system.



## Questions

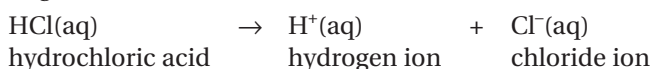
- 6 Predict the products of the reactions involving acids below and balance the equations.
- $\text{CuO} + \text{H}_2\text{SO}_4$
  - $\text{CaS} + \text{HNO}_3$
  - $\text{Sn} + \text{HCl}$
  - $\text{CaCO}_3 + \text{H}_3\text{PO}_4$
  - $\text{Mg}(\text{HCO}_3)_2 + \text{HCl}$
- 7 For each of the following reactions write an overall equation.
- lead(II) hydrogen carbonate and sulfuric acid
  - iron(III) hydroxide and hydrochloric acid
  - calcium carbonate and nitric acid
- 8 One particular brand of indigestion tablet contains the following active ingredients:
- calcium carbonate, 750 mg
  - magnesium carbonate, 120 mg
  - aluminium hydroxide, 120 mg.
- It is claimed that this preparation relieves the discomfort caused by excess stomach acid.
- Which acid is present in the human stomach?
  - Write a balanced equation for the reaction of each of the above active ingredients with stomach acid.
  - A short time after taking the tablets you burp. Why?

## 8.3 Defining acids and bases

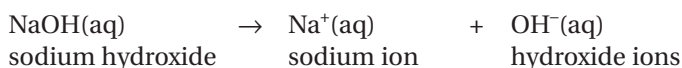
### Arrhenius acids and bases

Over time chemists have defined acids and bases in different ways. In 1887 the Swedish chemist Svante Arrhenius released his ideas on acids and bases. Arrhenius defined an acid as a substance that contains hydrogen and produces  $\text{H}^+$  ions in aqueous solution and a base as a substance that contains one or more hydroxide groups and produces  $\text{OH}^-$  ions in aqueous solution.

For example, an acid such as hydrochloric acid will ionise in water to release hydrogen ions.

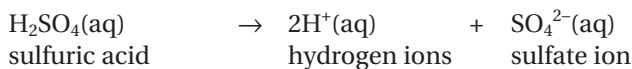


A base such as sodium hydroxide will ionise in water to release hydroxide ions.



**Monoprotic** acids only donate one proton per molecule. Hydrochloric acid is an example of a monoprotic acid. Some acids ionise to produce more than one hydrogen ion. These are known as **polyprotic** acids.

Sulfuric acid is referred to as a **diprotic** acid because it ionises to give two protons.



Phosphoric acid ( $\text{H}_3\text{PO}_4$ ) ionises to produce three protons and is known as a **triprotic** acid.

The Arrhenius theory did not allow for some compounds that were regarded as acids or bases but could not be described by the Arrhenius definitions. Ammonia ( $\text{NH}_3$ ) has all the properties of a base yet it does not contain hydroxide ions. The definition of acids and bases needed to be further developed.

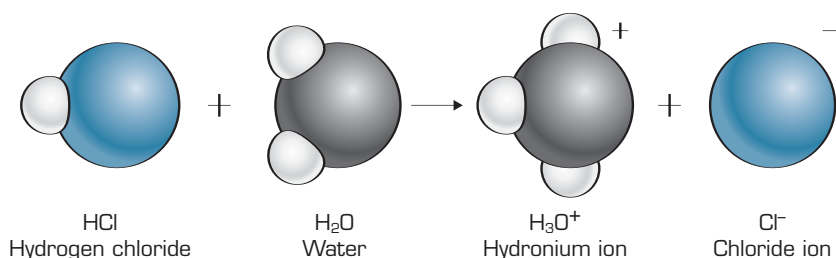
A proton donor is a substance that releases a proton. A proton acceptor receives the proton.

A hydrogen ion ( $\text{H}^+$ ) consists of a nucleus that contains a single proton.

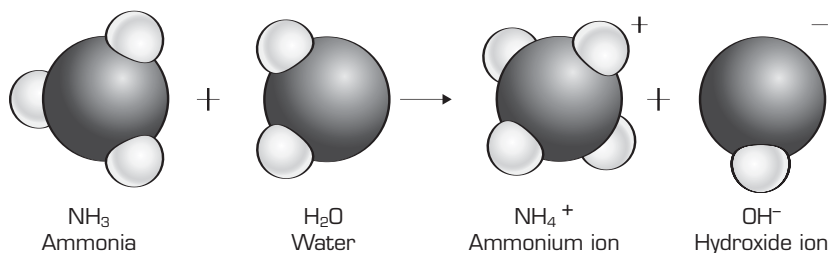
A **monoprotic acid** can donate only one hydrogen ion when reacting with a base.

**Polyprotic acids** can donate more than one hydrogen ion when reacting with a base.





**figure 8.3** In this reaction hydrochloric acid acts as the proton donor and water as the proton acceptor.



**figure 8.4** In this reaction the base ammonia acts as a proton acceptor and water as the proton donor.

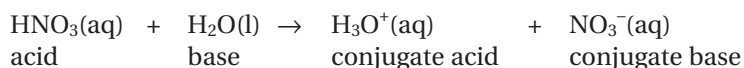
## Brønsted–Lowry acids and bases

In 1923 Johannes Brønsted and Thomas Lowry developed a more inclusive definition of acids and bases. The Brønsted–Lowry theory defines acids and bases by the way they react with each other. When an acid and a base react, the acid acts as a **proton donor** and the base as a **proton acceptor**. An acid donates a proton in the form of a  $\text{H}^+$  ion, and a base accepts the  $\text{H}^+$  ion.

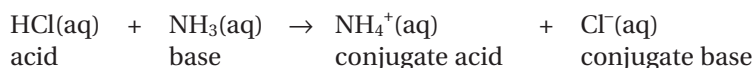
$\text{H}^+$  ions react with water molecules to form the hydronium ions ( $\text{H}_3\text{O}^+$ ) in aqueous solution.

According to the Brønsted–Lowry theory the base forms a **conjugate acid** when it receives a proton. Conversely an acid forms a **conjugate base** after donating a proton.

For example, in the reaction between nitric acid and water, water acts as a base and accepts a proton from nitric acid. It then forms a conjugate acid. The acid donates a proton and forms a conjugate base.



Similarly, in the reactions between hydrochloric acid and ammonia, the acid donates a proton to form a conjugate base. Ammonia (a base) accepts a proton and forms a conjugate acid.



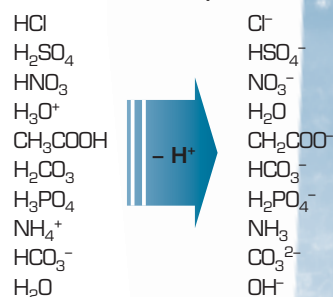
Conjugate acids are paired with the base and conjugate bases are paired with the acid in a neutralisation reaction. In the reactions between nitric acid and water, nitric acid and the nitrate ion are a **conjugate acid–base pair**. Water and the hydronium ion are also a conjugate acid–base pair.

In the reaction between hydrochloric acid and ammonia, the hydrochloric acid and chloride ion are a conjugate acid–base pair, and ammonia and the ammonium ion are also a conjugate acid–base pair.

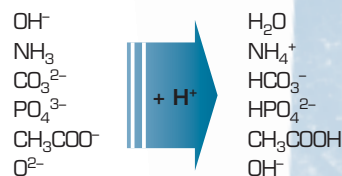


A **conjugate acid** is the species formed when a base accepts a proton. A **conjugate base** is the species formed when an acid donates a proton.

**Acids ... donate a proton to form:**



**Bases ... accept a proton to form:**



**figure 8.5** Some common acids and bases.





Substances that can act as either an acid or a base are said to be **amphiprotic**.

Amphiprotic substances are also called amphoteric substances.

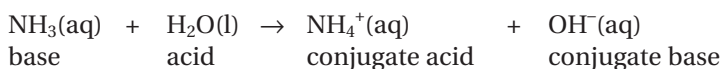
## Worksheet 1

Acids and bases

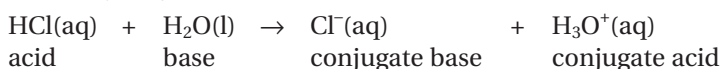
## Amphiprotic substances

Amphiprotic substances can act as either an acid or a base according to the Brønsted–Lowry theory.

For example water can act as a hydrogen ion donor (acid) in the reaction:



or as a hydrogen ion acceptor (base) in the reaction:



Water is an example of an amphiprotic substance. Some others are listed below.

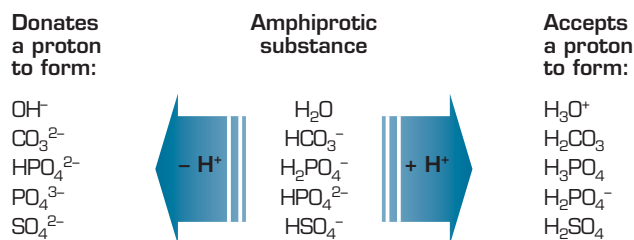


figure 8.6 Some amphiprotic substances.

## Experiment 8.2

### Amphiprotic substances

#### materials

- 0.1 M NaHCO<sub>3</sub>
- 0.1 M NaHSO<sub>4</sub>
- 0.1 M Na<sub>2</sub>HPO<sub>4</sub>
- 0.1 M NaH<sub>2</sub>PO<sub>4</sub>
- universal indicator and pH chart or pH meter
- semi-micro test tubes or small beakers if pH meter is to be used

#### safety

- Wear safety glasses and a laboratory coat for this experiment.

#### Purpose

To determine the relative acid and base strengths of some amphiprotic substances.

#### Method

- 1 Place 20 drops of each solution into separate semi-micro test tubes. Add one drop of universal indicator to each solution. Record the colour of each solution. Alternatively, a pH meter can be used.
- 2 Use the indicator chart to determine the pH of each solution.

#### Theory

Amphiprotic substances can act as both acids and bases. In aqueous solution, an amphiprotic substance will undergo two hydrolysis reactions, one as an acid and one as a base. In most cases, one of these reactions will predominate. If the acid reaction proceeds to a greater extent than the base reaction, the concentration of hydronium ions (H<sub>3</sub>O<sup>+</sup>) is greater than the concentration of hydroxide ions (OH<sup>-</sup>) and the pH of the solution will be less than 7. If the base reaction proceeds further than the acid reaction, then the pH of the solution will be greater than 7.

In this experiment, the sodium salt of each amphiprotic anion is used. The Na<sup>+</sup>(aq) ion does not hydrolyse to any appreciable extent in water. Any change of pH can therefore be attributed to the acid–base reactions of the amphiprotic anion in each solution.

#### Discussion

- 1 Write equations for the acid and base hydrolysis undergone by each of the amphiprotic substances tested. In each case, deduce whether the substance is stronger as an acid or as a base.

## Questions

- 9 Define the terms acid and conjugate acid using the Brønsted–Lowry theory.
- 10 Define the terms conjugate base and base using the Brønsted–Lowry theory.
- 11 Identify the acid/conjugate base and base/conjugate acid pairs in each of the following neutralisation reactions.
- a  $\text{H}_2\text{SO}_4(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{HSO}_4^-(\text{aq}) + \text{H}_3\text{O}^+(\text{aq})$
- b  $\text{CH}_3\text{COOH}(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{CH}_3\text{COO}^-(\text{aq}) + \text{H}_2\text{O}(\text{l})$
- c  $\text{CO}_3^{2-}(\text{aq}) + \text{CH}_3\text{COOH}(\text{aq}) \rightarrow \text{HCO}_3^-(\text{aq}) + \text{CH}_3\text{COO}^-(\text{aq})$
- d  $\text{HCO}_3^-(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{CO}_3^{2-}(\text{aq}) + \text{H}_2\text{O}(\text{l})$

## 8.4 Concentration and strength

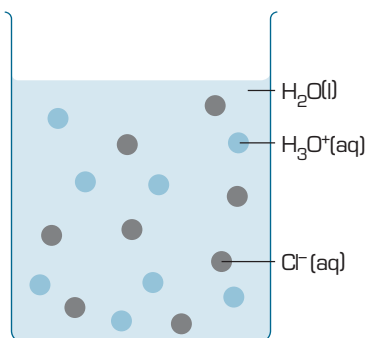
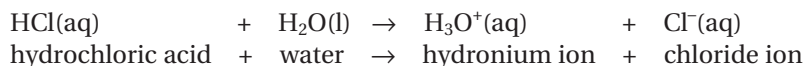
### Dilute and concentrated

The terms dilute and concentrated have no relationship to the strength of an acid or a base. These terms only describe the amount of acid or base dissolved in a specified volume of solution.

For example, concentrated hydrochloric acid would have a large amount of hydrogen chloride dissolved in a fixed volume of water. A dilute solution of hydrochloric acid would only have a small amount of hydrogen chloride dissolved in a fixed volume of water. Both solutions would contain a strong acid, but the more concentrated solution will contain more hydrogen ions per litre of solution.

### Strong and weak

The strength of acids is determined by their ability to ionise in solution. When a strong acid such as hydrochloric acid is dissolved in water it forms hydronium and chloride ions.

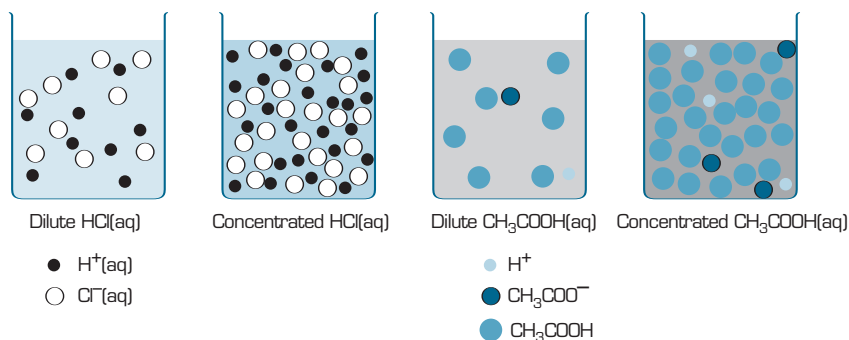


**figure 8.7** Strong acids like HCl will ionise completely in solution.

Because it ionises completely, hydrochloric acid is referred to as a strong acid. Weaker acids such as vinegar (acetic acid) only partially ionise in solution and are described as weak acids.



An **electrolyte** is a substance that dissolves in water to form ions that are free to move and hence conduct electricity.

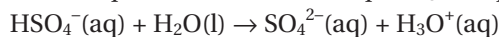
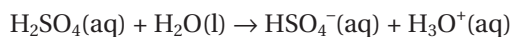


**figure 8.8** A concentrated strong acid (HCl) and a concentrated weak acid ( $\text{CH}_3\text{COOH}$ ). You can see the number of hydrogen ions is higher in the strong acid solution.

Strong acids react faster and are more corrosive than weak acids of the same concentration. A strong acid is a better electrolyte than a weak acid, i.e. a strong acid will conduct electricity much better than weak acid solution of the same concentration. This is because there would be more dissolved ions in a solution of strong acid which are able to conduct a current.

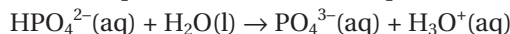
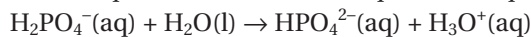
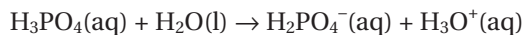
Most acids are **monoprotic** and only donate one proton per molecule. Monoprotic acids can be weak or strong also. Hydrochloric acid is a strong monoprotic acid and acetic acid (vinegar) is a weak monoprotic acid.

Sulfuric acid is referred to as a **diprotic acid** because it donates two protons.



Sulfuric acid is a strong acid because it ionises to a large extent in aqueous solution, not because it is diprotic.

Phosphoric acid is a **triprotic acid** and can donate three protons. It is still only a weak acid because only a small proportion of phosphoric acid will ionise when reacting with a base.



Whether an acid is monoprotic, diprotic or triprotic does not determine its strength. The **strength of an acid** is determined by the extent to which it ionises in aqueous solution.

## Experiment 8.3

### Strong and weak acids

#### materials

- 3 × 100 mL beakers
- 4 × test tubes
- test-tube rack
- 10 mL measuring cylinder
- 100 mL measuring cylinder
- 10 mL of 1 M  $\text{CH}_3\text{COOH}$
- 10 mL of 1 M HCl
- methyl violet indicator

*Methyl violet indicator is added to test tubes containing 1 M  $\text{CH}_3\text{COOH}$ , 1 M HCl, 0.1 M HCl and 0.01 M HCl. The colour of the indicator in the 1 M  $\text{CH}_3\text{COOH}$  most closely matches the colour in the 0.01 M HCl.*

#### Purpose

To compare the extent of ionisation of 1 M  $\text{CH}_3\text{COOH}$  and 1 M HCl.

#### Method

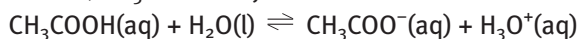
- 1 Prepare a solution of approximately 0.1 M HCl by adding 10 mL of 1 M HCl to 90 mL of water.
- 2 Prepare a 0.01 M HCl solution by diluting the 0.1 M HCl solution in the same way.



- Place solutions of 1 M, 0.1 M and 0.01 M HCl in test tubes to a depth of about 5 cm. To each, add three drops of methyl violet indicator.
- Add three drops of methyl violet to a test tube containing a similar quantity of 1 M CH<sub>3</sub>COOH. Match the indicator colour in 1 M CH<sub>3</sub>COOH to one of the HCl solutions.

### Theory

Unlike HCl, CH<sub>3</sub>COOH only ionises to a small extent in water:



Because just 1% of the CH<sub>3</sub>COOH molecules are present as CH<sub>3</sub>COO<sup>-</sup>, the concentration of H<sub>3</sub>O<sup>+</sup> ions in 1 M CH<sub>3</sub>COOH is about 0.01 M (pH 2). Methyl violet changes colour at a pH value of about 1.

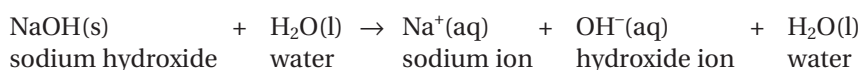
### Discussion

- What is the concentration of H<sub>3</sub>O<sup>+</sup> ions in each of the HCl solutions?
- Using your answer to Question 1, estimate the concentration of H<sub>3</sub>O<sup>+</sup> ions in 1 M CH<sub>3</sub>COOH.
- Approximately what percentage of the CH<sub>3</sub>COOH molecules have been ionised?

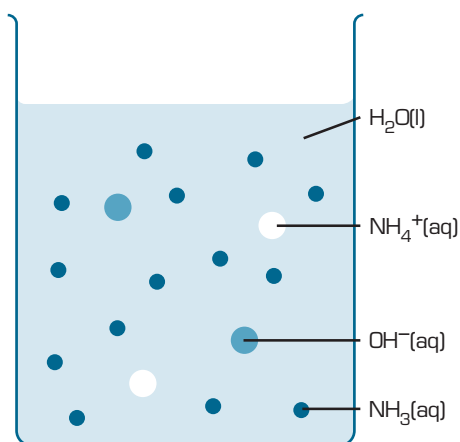
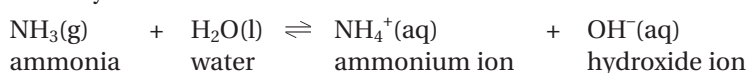
### Alternative activity

If two equal amounts of marble chips are simultaneously added to each of the acids in the test tubes, the rates of evolution of carbon dioxide gas may also be used as a measure of relative acidity. Alternatively, magnesium may be added and the rates of evolution of hydrogen gas compared.

The same principles apply for the strength of bases as well. A strong base such as sodium hydroxide will ionise completely to form sodium and hydroxide ions in solution.



A weaker base such as ammonia will not ionise as readily as sodium hydroxide. Only a fraction of ammonia will react with water to form ammonium ions and hydroxide ions.



**figure 8.9** Weak bases like ammonia do not ionise completely in solution.

### safety

- Wear safety glasses and a laboratory coat for this experiment.

### Worksheet 2

Acid-base reactions



The double arrow  $\rightleftharpoons$  indicates that a process is reversible. Not only do reactants form products but the products can also react together to reform the original reactants.



## Questions

- 12 a What type of acid (monoprotic, diprotic, etc.) is  $\text{H}_3\text{AsO}_4$ ?  
b Write an equation(s) to show how it ionises in water.  
c Identify the types of polyprotic acids in table 8.1.
- 13 What is the difference between a strong acid and a concentrated acid?
- 14 Why do strong acids conduct electricity better than weak acids?
- 15 What volume of 10 M hydrochloric acid would be required to prepare 250 mL of 0.30 M acid solution?

## 8.5 The pH scale

### pH

A Danish scientist called Soren Sorenson proposed the pH scale in 1909. The concentration of hydrogen ions in a solution is used to compare the strength of an acid or a base. An acid solution contains a higher concentration of hydrogen ions and will have a lower pH. A basic solution contains a lower concentration of hydrogen ions and will have a higher pH.

The pH scale usually ranges from 1 to 14, but can be outside this range in cases of very strong, concentrated acids and bases. A pH of 7 indicates a solution is neutral and is neither an acid nor a base. It is possible to have a pH of less than 1 or greater than 14. The numbers merely provide a means of making a comparison between solutions. A pH of 0–3 can be described as very acidic, 3–6 as mildly acidic, 8–11 as mildly basic and 11–14 as very basic.

The pH of a solution is calculated using the negative logarithm of the concentration of hydrogen ions and is expressed as

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

### ✓ Worked example 8.5A

If a solution of hydrochloric acid has a concentration of hydrogen ions equal to 1.0 M then the negative logarithm of this concentration defines this solution as having a pH of 0.

$$\begin{aligned}\text{pH} &= -\log[\text{H}^+] \\ \text{pH} &= -\log[1.0] = 0.0\end{aligned}$$

This would indicate a strongly acidic solution.

### ✓ Worked example 8.5B

A solution of sodium hydroxide had a hydrogen ion concentration of  $1.0 \times 10^{-14}$  M. The negative logarithm of this solution defines the solution as having a pH of 14.

$$\begin{aligned}\text{pH} &= -\log[\text{H}^+] \\ \text{pH} &= -\log[1 \times 10^{-14}] = 14\end{aligned}$$

This would indicate a very alkaline (or basic) solution.

Square brackets [ ] are used to represent concentration in  $\text{mol L}^{-1}$ .

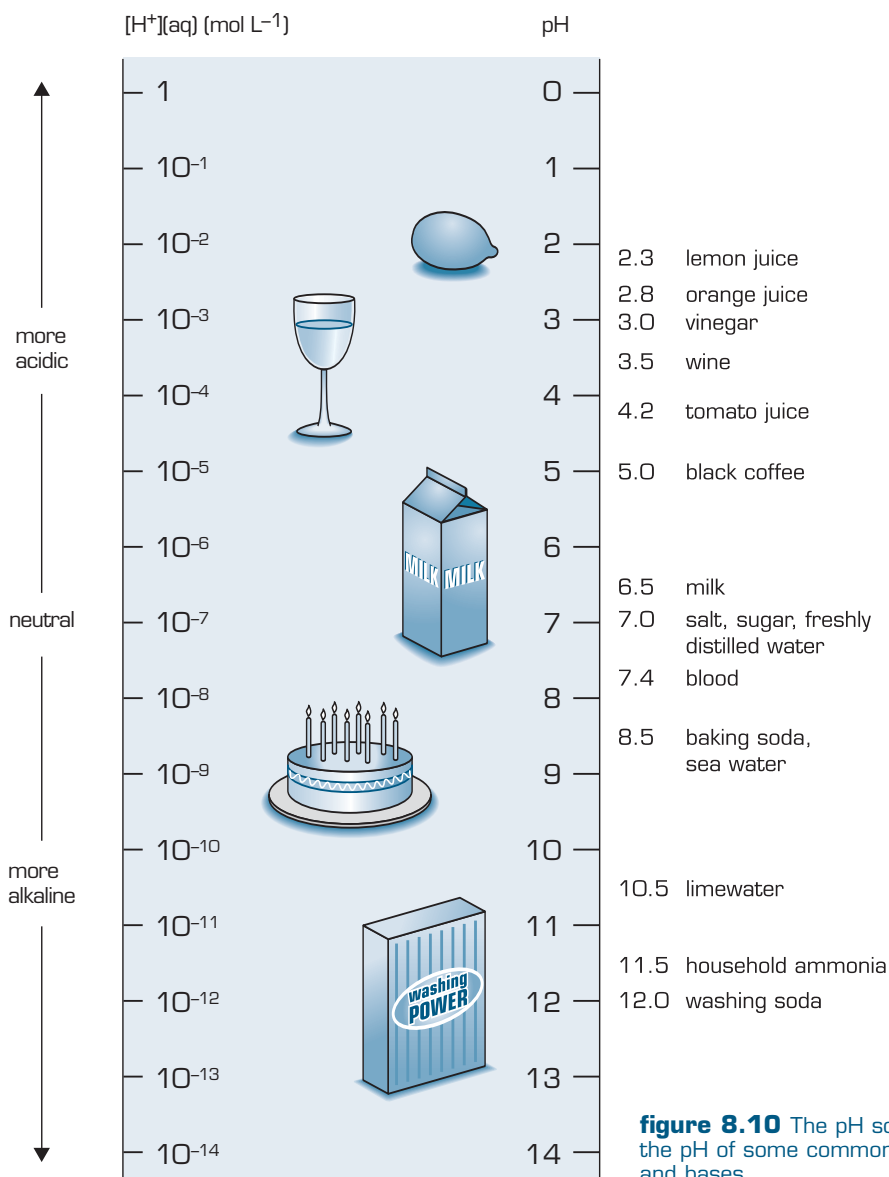




## Worked example 8.5C

The pH of a particular solution can also be used to determine the concentration of hydrogen ions present. For example, a solution of acetic acid with a pH of 5.0 would have a hydrogen ion concentration of  $1 \times 10^{-5}$  M.

$$\begin{aligned} \text{pH} &= -\log[\text{H}^+] \\ 5.0 &= -\log[\text{H}^+] \\ [\text{H}^+] &= 10^{-5.0} \\ [\text{H}^+] &= 1 \times 10^{-5} \text{ M} \\ (\text{i.e. } [\text{H}^+] &= 10^{-\text{pH}}) \end{aligned}$$



The pH scale is used widely throughout industry and in the home. Pool testing kits, used to maintain water quality in swimming pools, use the pH scale. Soap makers monitor the pH of their product to ensure it won't be too caustic for your skin and eyes.



The pH of blood should remain at around 7.35–7.45 and stomach acid should be between 1–2 for an individual to remain healthy. If the pH of a lake drops from 6–7 to 3–4 then all the fish and plant life will die. The pH scale makes communicating the acidity level of a substance very easy.

## Experiment 8.4

### Measuring pH

#### materials

- several different 50 mL water samples collected from local streams, drains, dams or ponds and from the local water supply
- pH meter or pH paper and pH colour code reference chart
- beaker
- test tubes

#### safety

- Wear safety glasses and a laboratory coat for this experiment.

#### Purpose

To measure the pH of samples of water from several different sources.

#### Method

##### A. Using a pH meter

- 1 If necessary calibrate the pH meter using the instructions provided with the instrument.
- 2 Insert the probe of the pH meter into a beaker containing one of the 50 mL water samples. Allow to stand for 30 seconds. Record the pH reading.
- 3 Repeat this procedure for each water sample.

##### B. Using pH paper

Add a 2 cm length of pH paper to a test tube containing about 20 mL of one of the water samples. Estimate the pH of the sample by matching the colour of the pH paper with the colour code reference chart.

#### Discussion

- 1 What might lead to a sample of reservoir water having a pH value less than 7?
- 2 What is the pH of a solution in which the concentration of hydrogen ions is:
  - a  $10^{-2}$  M?
  - b 0.000 01 M?

### Questions

- 16 Determine the pH of the solutions described below.
- a  $[H^+] = 0.01$  M
  - b  $[H^+] = 1 \times 10^{-11}$  M
  - c  $[H^+] = 10^{-5}$  M
  - d  $[H^+] = 1 \times 10^{-7}$  M
  - e  $[H^+] = 2.5 \times 10^{-7}$  M
  - f  $[H^+] = 2.3 \times 10^{-1}$  M
  - g  $[H^+] = 0.5$  M
  - h  $[H^+] = 5 \times 10^{-8}$  M
- 17 Determine the  $[H^+]$  of the following solutions.
- a pH = 6.0
  - b pH = 7.0
  - c pH = 13.0
  - d pH = 3.5
  - e pH = 12.7
  - f pH = 5.4
  - g pH = 1.6
  - h pH = 8.1
- 18 Calculate the pH of a solution that contains 50 g of  $HNO_3$  dissolved in enough water to create 1 L of solution.
- 19 Coca-Cola has a pH of 3 and black coffee has a pH of 5. How many more times acidic is Coca-Cola than black coffee?

## 8.6 Salts

When aqueous solutions of acids and bases are mixed, a reaction occurs in which the hydrogen ion from the acid and the hydroxide ion from the base combine in a neutralisation process that produces a **salt** and **water**. After a neutralisation reaction, the salt solutions consist of a mixture of positive and negative ions (and their names are in the salt name). For example, sodium chloride (NaCl) is a mixture of  $\text{Na}^+$  and  $\text{Cl}^-$  ions, calcium chloride ( $\text{CaCl}_2$ ) is a mix of  $\text{Ca}^{2+}$  and  $\text{Cl}^-$  ions etc.

Although all acid–base reactions are referred to generally as neutralisation reactions, strictly speaking **neutralisation** involves mixing an acid ( $\text{pH} < 7$ ) with a base or alkali ( $\text{pH} > 7$ ), which react to form a **neutral** salt solution of  $\text{pH} 7$ . This is not the case in the majority of acid–base reactions, as the salts can undergo hydrolysis (react with water) either accepting or donating a proton to form conjugate acid–base pairs and elevating or lowering the  $\text{pH}$ .

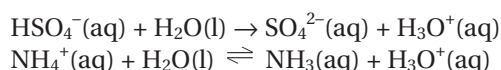
### Acidity of salt solutions

Salts are a class of compounds that consist of a cation other than  $\text{H}^+$  and an anion other than  $\text{OH}^-$  or  $\text{O}^{2-}$ . The acidity ( $\text{pH}$ ) of salt solutions depends on whether or not their anions and cations hydrolyse. If a cation or anion donates a proton to water it forms hydronium ions ( $\text{H}_3\text{O}^+$ ) that causes the salt solution to be acidic. If an anion accepts protons from water, hydroxide ( $\text{OH}^-$ ) ions are formed, causing the salt solution to be basic.

### Acidic salts

Examples of acidic salts include  $\text{NaHSO}_4$  and  $\text{NH}_4\text{Cl}$ .

When combined with water, the hydrogen sulfate ion ( $\text{HSO}_4^-$ ) and the ammonium ion ( $\text{NH}_4^+$ ) will donate a proton to the water to form a hydronium ion.

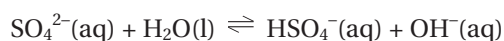


### Basic salts

Some anions of salts are able to accept protons from water molecules. In these cases the salt will be alkaline. Cations cannot accept protons and therefore cannot make a salt alkaline.

Examples of basic salts include  $\text{Na}_2\text{SO}_4$  and  $\text{KCH}_3\text{COO}$ .

When combined with water the sulfate ion ( $\text{SO}_4^{2-}$ ) accepts a proton from water.



The ethanoate ion ( $\text{CH}_3\text{COO}^-$ ) accepts a proton from water to form ethanoic acid molecules.

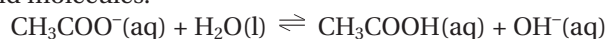


Table 8.4 classifies common ions according to whether they cause aqueous solutions to be acidic, basic or neutral.

Small highly charged metal ions such as iron(III), zinc and aluminium are weakly acidic in aqueous solution. For example, the hydrated iron(III) ion hydrolyses according to the following reaction:



Two examples of how salts are useful in chemistry are as **buffers**, to regulate the  $\text{pH}$  of a solution, and as **indicators**, in monitoring the  $\text{pH}$  of a substance.



A soluble base is also known as an alkali.



Hydrolysis is a reaction in which an aqueous ion accepts or donates a proton to water.



**table 8.4** Acid–base properties of some common ions in aqueous solution.

Type of ion	Acidic	Basic	Neutral
cation	NH <sub>4</sub> <sup>+</sup> Fe <sup>3+</sup> Al <sup>3+</sup> Zn <sup>2+</sup>	none	metal ions from Group 1 and Group 2
anion	HSO <sub>4</sub> <sup>-</sup> H <sub>2</sub> PO <sub>4</sub> <sup>-</sup>	F <sup>-</sup> , S <sup>2-</sup> , SO <sub>4</sub> <sup>2-</sup> , ClO <sup>-</sup> , CH <sub>3</sub> COO <sup>-</sup> , CO <sub>3</sub> <sup>2-</sup> , HCO <sub>3</sub> <sup>-</sup> , PO <sub>4</sub> <sup>3-</sup> , HPO <sub>4</sub> <sup>2-</sup> , C <sub>2</sub> O <sub>4</sub> <sup>2-</sup>	Cl <sup>-</sup> , NO <sub>3</sub> <sup>-</sup> , Br <sup>-</sup> , I <sup>-</sup>

A **buffer** is a solution that resists changes in pH by neutralising both acids and bases.

## Buffers

**Buffers** are substances that resist changes in pH despite the addition of small amounts of acids or bases to a solution.

They usually consist of a weak acid or base, and work by removing any hydrogen and hydroxide ions from a solution. By removing both of these ions, the pH will remain constant.

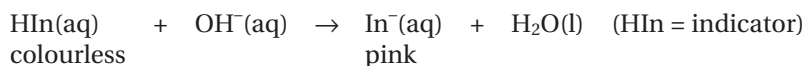
Buffers are used in industries where fluctuations in pH are harmful to the manufacturing process. An example of this is the dyeing of fabrics. The dyes will not bind to the fabric as effectively if the pH is not maintained at the correct level. Our bodies have natural buffers in our blood and gastric juice to maintain the pH at healthy levels. Even the ocean maintains a healthy pH by removing excess acid using a buffer. Excessive addition of an acid or base can overpower a buffering system.

## Indicators

The most common way to determine the pH of an acid or base is with a substance called an **indicator**. Indicators change colour depending on the presence or the strength of an acid or base.

Indicators can also give an indication of the pH of a particular substance and are used in a chemical technique called titration to determine the exact concentration of acid or base in a particular solution.

Indicators are actually conjugate acids and bases that form different colours. The indicator phenolphthalein, for example, remains colourless in acidic solutions but turns pink in alkaline solutions.



This colour change occurs around a pH of 8–10.

Other indicators are different colours and change colour at a different pH range. Methyl red changes from red to yellow at pH range of 4–6. Bromocresol green changes from red to blue at a pH range of 4–6. Others such as litmus change from red to blue at pH range of 5–8

When an indicator gains or releases a hydrogen ion it changes how that substance absorbs and reflects different colours of light, causing a colour change.

There are some limitations that govern the use of indicators. First, changes in temperature will marginally affect the pH at which they will change colour. Second, indicators will also only give a rough guide to the pH because they usually change colour over a range of two pH points.

**Indicators** are chemical substances that change colour according to the pH.

table 8.5 Common indicators.

Indicator	Colour of acid form	Colour of base form
phenolphthalein	colourless	pink
methyl orange	pink	yellow
bromothymol blue	yellow	blue

## Experiment 8.5

### Unknown solutions

#### Purpose

To use the indicator methyl orange to identify each one of four different solutions as either water, 0.1 M HCl, 0.01 M HCl or 0.1 M NaOH.

#### Method

Devise a method to identify the four solutions.

#### Theory

Methyl orange is pink at pH less than 3 and yellow at pH greater than 4.

#### Discussion

- 1 Describe the method you used to identify the four solutions.
- 2 Estimate the pH of each solution.
- 3 Phenolphthalein is an indicator which changes from colourless at pH less than 8 to pink at pH greater than 10. What observations would you expect to make if you used phenolphthalein instead of methyl orange to identify the four solutions?
- 4 Sodium hydroxide is a strong alkali. Strong alkalis may be neutralised with strong acids. When sodium hydroxide is spilled onto clothes or skin the correct safety procedure is to wash immediately with water. Why should a strong acid not be used to neutralise the alkali in this situation?

#### materials

- 10 mL of each unknown solution
- methyl orange indicator
- test tubes
- test-tube rack

#### safety

- Wear safety glasses and a laboratory coat for this experiment.
- Any spillages should be mopped up immediately, and if the solution is splashed onto clothes or skin it should be washed off using plenty of water.

## Experiment 8.6

### Universal indicator colour display

*Two red solutions are mixed. The colour of the mixture changes from red to yellow to green to blue to purple.*

#### Purpose

To show the change in colour of universal indicator as the pH changes in a chemical reaction.

#### Method

##### Prior to the demonstration

- 1 Prepare a 500 mL solution containing 0.5 g  $\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$ , 18.5 g KI and 3 mL of 0.1 M  $\text{CH}_3\text{COOH}$  and enough universal indicator to produce a strong red colour. The solution is stable for about 6 hours.
- 2 Prepare a 500 mL solution containing 3 mL of 0.1 M  $\text{CH}_3\text{COOH}$ , 5 mL of 20 vol  $\text{H}_2\text{O}_2$  and enough universal indicator to produce a strong red colour.

#### materials

- 2 × 600 mL beakers
- 250 mL measuring cylinder (or 600 mL tall beaker)
- 2 × 10 mL measuring cylinders
- glass stirring rod
- white cardboard
- 0.5 g  $\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$
- 18.5 g KI
- 6 mL of 0.1 M  $\text{CH}_3\text{COOH}$
- universal indicator
- 5 mL of 20 vol  $\text{H}_2\text{O}_2$





## safety

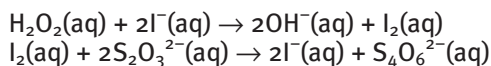
- Wear gloves, safety glasses and a laboratory coat for this experiment.
- Avoid contact of  $\text{H}_2\text{O}_2$  with the eyes and skin.

## For the demonstration

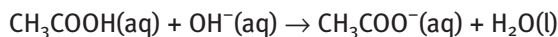
Pour equal volumes of the two solutions into a measuring cylinder and observe the colour changes. A white card behind the cylinder is helpful.

## Theory

A series of reactions is involved:



As the  $\text{OH}^-$  is formed, it reacts with the ethanoic acid:



causing the pH to increase and the indicator colour to change.

## Discussion

- 1 What does the colour of the universal indicator in the original solutions tell you about their pH?
- 2 What colour do you expect to observe when the two solutions are mixed?
- 3 Suggest reasons for the colour changes observed when the solutions are mixed.
- 4 Use the equations of the reactions involved in this process to explain the increase in pH observed.
- 5 Many indicators have a single acid colour and a single base colour. Suggest why universal indicator shows a number of colours.

## Alternative activities

- This demonstration is particularly effective when carried out in a Petri dish on an overhead projector.
- The rate of the reaction is determined by the concentration of  $\text{H}_2\text{O}_2$ . Different concentrations of  $\text{H}_2\text{O}_2$  may be used to show the effect of concentration of reactants on reaction rate.

## Questions

- 20 Write equations for the hydrolysis of the following ions.
- a fluoride
  - b ammonium
  - c ethanoate
  - d dihydrogenphosphate
  - e hypochlorite
  - f aluminium
- 21 Predict whether aqueous solutions of the following salts will be acidic, basic or neutral.
- a aluminium iodide
  - b calcium hydrogensulfate
  - c potassium sulfate
  - d potassium hydrogenphosphate
  - e sodium carbonate
  - f sodium sulfide

## Summary ▶▶▶▶

- ▶ **Acids** and **bases** are very common substances.
- ▶ Bases that dissolve in water are called **alkalis**.
- ▶ Acids and bases will react to form a salt and water in a **neutralisation reaction**.
- ▶ Acids undergo characteristic reactions.
- ▶ Acids and bases can be defined by the **Arrhenius** and **Brønsted–Lowry** theories.
- ▶ The Brønsted–Lowry theory defines acids as **proton donors** that form **conjugate bases**, and bases as **proton acceptors** that form **conjugate acids**.
- ▶ **Amphiprotic** substances can act as acids or bases.
- ▶ Acids and bases can be strong or weak. Strong acids make excellent **electrolytes**.
- ▶ Most acids are **monoprotic**, but some can be **polyprotic—diprotic** or **triprotic**.
- ▶ The strength of acids and bases can be compared on the **pH scale**. The **pH** of a solution is the negative logarithm of the concentration of hydrogen ions.
- ▶ **Buffers** are solutions that resist changes in pH.
- ▶ **Indicators** are substances that change colour depending on the pH.
- ▶ The **salt** of an acid–base reaction may be acidic, basic or neutral. If the **anion** or **cation** can **donate a proton** to water in solution, the salt is **acidic**. If the **anion** can **accept a proton** from water then the salt is **basic**.

## Key terms

acid  
acidic salt  
alkali  
amphiprotic  
Arrhenius theory  
base

basic salt  
Brønsted–Lowry theory  
buffers  
conjugate acid  
conjugate base  
diprotic

electrolyte  
indicators  
monoprotic  
neutral salt  
neutralisation reaction  
pH

pH scale  
polyprotic  
proton acceptors  
proton donor  
salt  
triprotic



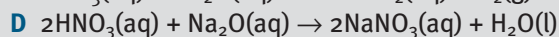
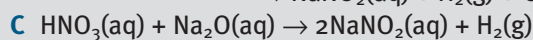
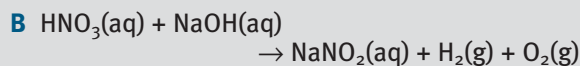
## Chapter Review

### 8.1 Introducing acids and bases

- 22 Which of the following is the formula of a common base?
- A HCl
  - B HNO<sub>3</sub>
  - C NH<sub>3</sub>
  - D CH<sub>3</sub>COOH
- 23 What amount (in mol) of sulfuric acid is needed to exactly neutralise 0.01 mol of sodium hydroxide?
- A 0.2
  - B 0.02
  - C 0.01
  - D 0.005
- 24
- a Give the name and formula of a common base.
  - b Give the name and formula of a common acid.
  - c Write an equation showing a reaction between the acid and base you have chosen.
- 25 Acids and bases have different characteristics.
- a List three characteristics of an acid.
  - b List three characteristics of a base.
  - c You are given a clear solution in a beaker. Briefly outline how you would safely determine if it was an acid or a base.

### 8.2 Reactions involving acids

- 26 Which of the following gases is produced when dilute acid solution reacts with a metal?
- A H<sub>2</sub>O
  - B H<sub>2</sub>
  - C CO<sub>2</sub>
  - D O<sub>2</sub>
- 27 Which of the following reactions represents a neutralisation reaction?
- A NaOH(aq) + HCl(aq) → NaCl(aq) + H<sub>2</sub>O(l)
  - B C(g) + O<sub>2</sub>(g) → CO<sub>2</sub>(g)
  - C 2NaOH(aq) + CuSO<sub>4</sub>(aq) → Cu(OH)<sub>2</sub>(s) + Na<sub>2</sub>SO<sub>4</sub>(aq)
  - D 2H<sub>2</sub>O<sub>2</sub>(l) → O<sub>2</sub>(g) + 2H<sub>2</sub>O(l)
- 28 Which of the following is not a likely product when an acid reacts with a metal carbonate?
- A H<sub>2</sub>O
  - B CO<sub>2</sub>
  - C H<sub>2</sub>CO<sub>3</sub>
  - D CuCO<sub>3</sub>
- 29 Which of the following equations correctly represents an acid reacting with a metal oxide?
- A HNO<sub>3</sub>(aq) + NaOH(aq) → NaNO<sub>3</sub>(aq) + H<sub>2</sub>O(l)



- 30 Which of the following is most likely to be produced when an acid reacts with a metal sulfide?
- A H<sub>2</sub>
  - B H<sub>2</sub>S
  - C SO<sub>2</sub>
  - D SO<sub>3</sub>
- 31 Carbon dioxide can be produced in the laboratory by reacting dilute sulfuric acid with sodium carbonate.
- a Write a full balanced chemical equation for the reaction that occurs.
  - b If 50.0 g of sodium carbonate is reacted with excess dilute sulfuric acid, calculate the mass of carbon dioxide produced.
  - c If 3.00 kg of carbon dioxide is produced, calculate the mass of sodium carbonate reacted.
  - d Suggest two safety precautions that should be followed during the procedure.
- 32 Magnesium metal reacts with hydrochloric acid.
- a Write a full balanced chemical equation for the reaction.
  - b Calculate the mass of hydrogen gas produced when 5.0 g of magnesium reacts.
  - c 0.50 g of magnesium is added to 150 mL of 0.10 M hydrochloric acid.
    - i Which reactant will be in excess and by how much?
    - ii Calculate the mass of hydrogen gas produced.
- ### 8.3 Defining acids and bases
- 33 Which of the following substances is both amphiprotic and polyprotic?
- A H<sub>2</sub>SO<sub>4</sub>
  - B HCO<sub>3</sub><sup>-</sup>
  - C H<sub>2</sub>PO<sub>4</sub><sup>-</sup>
  - D H<sub>3</sub>PO<sub>4</sub>
- 34
- a Write concise definitions for the following terms.
    - i Brønsted–Lowry base
    - ii strong acid
    - iii molarity
    - iv conjugate acid
  - b Explain, with the aid of equations, why HCO<sub>3</sub><sup>-</sup> is classified as amphiprotic.

- 35 The following equation represents a Brønsted–Lowry acid–base reaction:  
 $\text{HNO}_3(\text{aq}) + \text{NH}_3(\text{aq}) \rightarrow \text{NH}_4\text{NO}_3(\text{aq})$
- Which reactant is acting as an acid?
  - Give a reason for your answer to part a.
  - Estimate the pH of a weak  $\text{NH}_3$  solution. Explain your reasoning.
- 36 Write balanced equations for the three ionisation stages of phosphoric acid.

## 8.4 Concentration and strength

- 37 If 250.0 mL of 2.0 M HCl is diluted to 1 L, what amount of HCl (in mol) is present in 80.0 mL of diluted solution?
- 38 Calculate the concentration of each of the following diluted solutions.
- 20 mL of water is added to 6 mL of 0.5 M HCl.
  - 1.0 L of water is added to 8.0 L of 0.5 M HCl.
  - 3 L of water is added to 4 L of 0.5 M  $\text{H}_2\text{SO}_4$ .
  - 50 mL of 2 M HCl is added to 180 mL of water.
- 39 Calculate the amount (in mol) of solute in:
- 1.6 L of 0.75 M hydrochloric acid solution (HCl)
  - 3.0 L of 0.3 M nitric acid solution ( $\text{HNO}_3$ )
  - a mixture of 70 mL of 1.0 M HCl and 150 mL of 0.3 M HCl.

## 8.5 The pH scale

- 40 Human gastric juice has a pH of 1. What is the concentration of hydrogen ions in gastric juice?

- 41 2.7% of a 2.0 M monoprotic acid is ionised in a solution. Calculate the pH of the solution.
- 42 Human blood has a pH of 7.4. Is blood acidic, basic or neutral? What assumption have you made?
- 43 A solution of hydrochloric acid has a pH of 2.
- What is the molar concentration of hydrogen ions in the solution?
  - What amount of hydrogen ions (in mol) would be present in 500 mL of this solution?
- 44 Calculate the pH of:
- a solution made by dissolving 0.100 mol of  $\text{HNO}_3$  in water to make 100 mL of solution
  - a solution made by adding 10 mL of 1.5 M HCl to 90 mL water
  - a solution made by adding 50.0 mL water to a 100 mL solution of HCl with a pH of 2.00.

## 8.6 Salts

- 45 Classify the following salts as acidic, basic or neutral.
- $\text{BaCl}_2$
  - $\text{K}_2\text{SO}_4$
  - $\text{NaH}_2\text{PO}_4$
  - $\text{Na}_2\text{CO}_3$
  - $\text{NaHSO}_4$
  - $\text{NH}_4\text{Cl}$
  - $\text{K}_3\text{PO}_4$
  - $\text{CaI}_2$
  - $\text{KHCO}_3$

