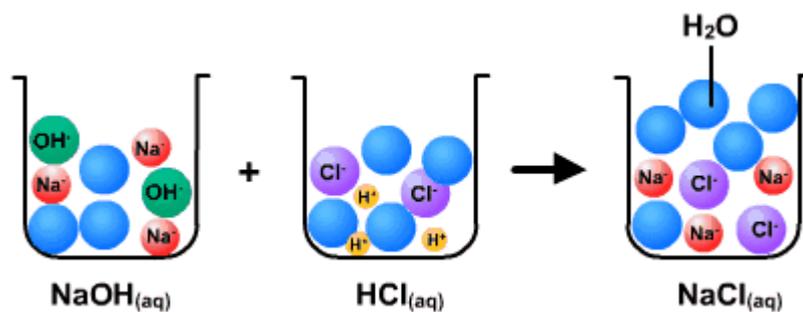


UNIT 5

ACIDS, BASES AND SALTS

PART 1 – PROPERTIES OF ACIDS, BASES AND SALTS



Contents

1. Acids, Bases, Salts and Neutralisation
2. Physical Properties of Acids, Bases and Salts
3. Strong and Weak Acids and Bases
4. Salt Hydrolysis

Key words: Arrhenius, Bronsted-Lowry, Acid, Base, Salt, Neutralisation, Hydrogen ion, Hydroxide ion, Proton Donor, Proton Acceptor, Alkali, Amphoteric, Hygroscopic, Deliquescent, Efflorescent, Water of Crystallisation, Strong Acid, Weak Acid, Strong Base, Weak Base, Salt Hydrolysis

Units which must be completed before this unit can be attempted:

Unit 1 – Atomic Structure and the Periodic Table

Unit 2 – Particles, Structure and Bonding

Unit 3 – Amount of Substance

Unit 4 – Introduction to Physical Chemistry

1) Acids, Bases, Salts and Neutralisation

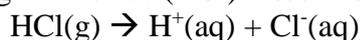
(a) Acids

(i) Arrhenius definition

An acid is a substance which releases hydrogen (H⁺) ions when dissolved in water. This is the Arrhenius definition of an acid.

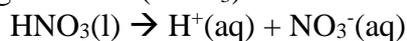
Common examples:

Hydrogen chloride (HCl) dissolves in water to produce hydrogen (H⁺) and chloride (Cl⁻) ions:



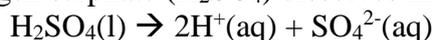
The resulting solution is known as **hydrochloric acid**.

Hydrogen nitrate (HNO₃) dissolves in water to produce hydrogen (H⁺) and nitrate (NO₃⁻) ions:



The resulting solution is known as **nitric acid**.

Hydrogen sulphate (H₂SO₄) dissolves in water to produce hydrogen (H⁺) and nitrate (SO₄²⁻) ions:



The resulting solution is known as **sulphuric acid**.

Acids are usually found and used in aqueous solution (aq).

(ii) Bronsted-Lowry definition

H⁺ ions can react with other ions and molecules in solution. The acid loses its H⁺ ion during these reactions.

H⁺ ions are protons. **An acid can therefore be described as a proton donor**. This is the Bronsted-Lowry definition of an acid.

(b) Bases**(i) Arrhenius definition**

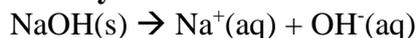
A base is a substance which releases hydroxide (OH⁻) ions when dissolved in water. This is the Arrhenius definition of a base.

There are several common types of base:

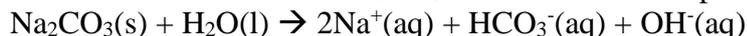
Some **metal oxides** such as CaO react with water to produce hydroxide (OH⁻) ions:



Some **metal hydroxides** such as NaOH dissolve in water to produce hydroxide (OH⁻) ions:



Some **metal carbonates** such as Na₂CO₃ react with water to produce hydroxide (OH⁻) ions:



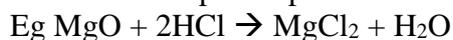
Ammonia (NH₃) reacts with water to produce hydroxide (OH⁻) ions:

**(ii) Bronsted-Lowry definition**

The hydroxide ions (OH⁻) produced by bases in water react readily with the protons (H⁺) in acids to form water. An alternative definition of a base is therefore a species which can accept H⁺ ions, or a **proton acceptor**. This is the Bronsted-Lowry definition of a base and includes various substances which do not dissolve in water and are therefore not classified as Arrhenius bases.

All metal oxides contain the oxide ion (O²⁻).

An oxide ion can accept two protons to form a water molecule: O²⁻ + 2H⁺ → H₂O



All metal oxides are Bronsted-Lowry bases, even if they do not dissolve in water.

All metal hydroxides contain the hydroxide ion (OH⁻). A hydroxide ion can accept a proton to form a water molecule: OH⁻ + H⁺ → H₂O



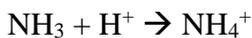
All metal hydroxides are Bronsted-Lowry bases, even if they do not dissolve in water.

All metal carbonates contain the carbonate ion (CO₃²⁻). A carbonate ion can accept two protons to form carbon dioxide and water: CO₃²⁻ + 2H⁺ → CO₂ + H₂O



All metal carbonates are Bronsted-Lowry bases, even if they do not dissolve in water.

Ammonia is also a Bronsted-Lowry base as it can accept a proton to form an ammonium ion:



According to Bronsted-Lowry theory, and reaction which involves the transfer of protons (ie from an acid to a base) is an acid-base reaction.

Although all Bronsted-Lowry bases react with acids, not all dissolve in water. A Bronsted-Lowry base which dissolves in water to produce a solution containing OH⁻ ions is called an **alkali**.

An alkali is a soluble base. An alkali is the same, therefore, as an Arrhenius base.

(c) Salts and Neutralisation Reactions

When an acid reacts with a base, it loses its H^+ ion. The H ion is replaced either with the metal ion from the base or with an ammonium ion. The resulting compound is called a salt.

A salt is the product formed when the H^+ ion from an acid is replaced by a metal ion or an ammonium ion.

A reaction between an acid and a base to make a salt is called a **neutralisation** reaction.

As there are four different types of base, there are four different types of neutralisation reaction in which a salt can be made from an acid and a base:

- (i) Acid + metal hydroxide \rightarrow metal salt + water
Eg $HCl + NaOH \rightarrow NaCl + H_2O$ (the salt produced is sodium chloride)
- (ii) Acid + metal oxide \rightarrow metal salt + water
Eg $HNO_3 + CaO \rightarrow Ca(NO_3)_2 + H_2O$ (the salt produced is calcium nitrate)
- (iii) Acid + metal carbonate \rightarrow metal salt + carbon dioxide + water
Eg $H_2SO_4 + K_2CO_3 \rightarrow K_2SO_4 + CO_2 + H_2O$ (the salt produced is potassium sulphate)
- (iv) Acid + ammonia \rightarrow ammonium salt
Eg $HCl + NH_3 \rightarrow NH_4Cl$ (the salt produced is ammonium chloride)

Test Your Progress: Topic 5 Exercise 1

Deduce the formulae of the following substances and indicate whether they are acids, bases or salts:

- | | |
|------------------------|-----------------------|
| a) Sodium oxide | j) Calcium nitrate |
| b) Calcium hydroxide | k) Sulphuric acid |
| c) Ammonium nitrate | l) Ammonium chloride |
| d) Potassium carbonate | m) Nitric acid |
| e) Strontium sulphate | n) Potassium sulphate |
| f) Ammonium sulphate | o) Magnesium oxide |
| g) Hydrochloric acid | p) Caesium bromide |
| h) Rubidium hydroxide | q) Barium sulphate |
| i) Magnesium carbonate | r) Strontium nitrate |

Test Your Progress: Topic 5 Exercise 2

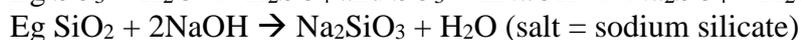
Write balanced symbol equations for the following neutralisation reactions:

- | | |
|---|--|
| a) Nitric acid with potassium hydroxide | g) Nitric acid with calcium carbonate |
| b) Sulphuric acid with sodium hydroxide | h) Hydrochloric acid with barium carbonate |
| c) Hydrochloric acid with calcium hydroxide | i) Sulphuric acid with sodium carbonate |
| d) Nitric acid with calcium oxide | j) Nitric acid with ammonia |
| e) Hydrochloric acid with barium oxide | k) Sulphuric acid with ammonia |
| f) Sulphuric acid with magnesium oxide | l) Hydrochloric acid with ammonia |

(d) Amphoteric substances

Proton donors (acids) will react with proton acceptors (bases), usually to form salts.

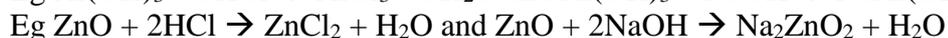
As well as the common acids (HCl, HNO₃ and H₂SO₄), which contain H⁺ ions, all oxides and hydroxides of non-metals (ie all covalently bonded oxides and hydroxides) can behave as acids. Some dissolve in water to form solutions containing H⁺ ions. Some do not dissolve in water but react with alkalis to form salts:



Oxides and hydroxides of metals, however (ie all ionically bonded oxides and hydroxides) can behave as bases. Some dissolve in water to form solutions containing OH⁻ ions. Some do not dissolve in water but react with acids to form salts:

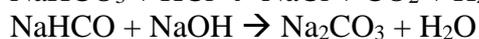


There are a small number of oxides and hydroxides of metals which can form salts by reacting with both acids and alkalis:



These substances can react with both acids and bases. They are proton donors and proton acceptors and are said to be **AMPHOTERIC**. The bonding in these oxides and hydroxides tends to be intermediate between ionic and covalent.

Other substances can show amphoteric properties by behaving as both proton donors and proton acceptors. Hydrogencarbonate ions (HCO₃⁻) are a good example:



(e) Solubility of acids, bases and salts

(i) acids

Most acids are soluble in water. Hydrochloric acid, sulphuric acid and nitric acid are all highly soluble in water.

(ii) bases

The solubility of bases in water varies:

Metal hydroxides are usually insoluble in water. They are therefore bases but not alkalis.

The exceptions are the Group I hydroxides (LiOH, NaOH, KOH etc) and barium hydroxide (Ba(OH)₂), which are highly soluble in water. Strontium hydroxide (Sr(OH)₂) and calcium hydroxide (Ca(OH)₂) are slightly soluble in water. These hydroxides are alkalis as well as bases.

Metal oxides are usually insoluble in water. They are therefore bases but not alkalis.

The exceptions are the Group I oxides (Li₂O, Na₂O, K₂O etc) and barium oxide (BaO), which are highly soluble in water. Strontium oxide (SrO) and calcium oxide (CaO) are slightly soluble in water.

When these oxides dissolve in water, they react with water to form their respective hydroxides.



Metal carbonates are usually insoluble in water. They are therefore bases but not alkalis.

The exceptions are the Group I carbonates (Li₂CO₃, Na₂CO₃, K₂CO₃ etc) which are highly soluble in water. These hydroxides are alkalis as well as bases.

Ammonia is highly soluble in water. It is therefore an alkali as well as a base.

(iii) Salts

Most common salts are soluble, with some exceptions:

- All nitrate salts are soluble
- Most common chloride salts are soluble (although there are some less common exceptions)
- Most common sulphate salts are soluble, except barium sulphate and strontium sulphate (which are completely insoluble) and calcium sulphate (which is only slightly soluble), and there are some less common exceptions.
- All Group I salts and all ammonium salts are soluble

(f) Preparing simple salts by neutralisation

The most suitable method of preparing a salt by a neutralisation reaction depends on the solubility of the salt and the solubility of the acid and base used to make it. Insoluble salts are generally not prepared by neutralisation reactions, but soluble salts frequently are:

There are two common types of preparation:

(i) Preparing a soluble salt from an acid and an insoluble base

If the salt being prepared is soluble, the easiest way to make it is from an acid and an insoluble base.

An excess of the insoluble base must be added to ensure that the acid reacts completely.

The salt solution can be separated from the excess base by simple filtration.

The salt can be extracted from the salt solution by gradually evaporating off the water.

Example: preparing the soluble salt copper sulphate from sulphuric acid (the acid) and copper oxide (the base):

Practical: prepare a sample of copper sulphate from copper oxide and sulphuric acid

- 1) Measure out 20 cm³ of 0.5 mol dm⁻³ sulphuric acid into a 100 cm³ beaker.
- 2) Warm the beaker gently on a tripod until the temperature reaches 50 °C.
- 3) Add 1 g of copper oxide gradually to the beaker over a period of 2 minutes, stirring gently.
- 4) Heat gently for a few minutes until it is almost boiling
- 5) Allow the mixture to cool.
- 6) Place a folded piece of filter paper inside a filter funnel, and then place the funnel into the neck of a 250 cm³ conical flask.
- 7) Pour the warm mixture into the filter funnel and allow the solution to filter through. A clear blue solution should collect in the conical flask.
- 8) Rinse the beaker and then pour the clear blue solution back into it. Label the beaker with your name and leave it for a week.
- 9) Use a spatula to remove any crystals in the beaker and dry them with a piece of filter paper.

How many moles of sulphuric acid (H₂SO₄) are used in the reaction?

How many moles of copper oxide (CuO) are used in the reaction?

Why is it important that the CuO is in excess?

https://www.youtube.com/watch?v=qIOMlwBoe_4

(ii) Preparing a soluble salt from an acid and an alkali

Group I and ammonium salts cannot be prepared from insoluble bases as ammonia and all Group I bases are soluble. These salts can only be prepared from acids and alkalis.

The acid and alkali must be added in the exact amounts required, as there is no way to separate the salt produced from any remaining acid or alkali. The final mixture should therefore only contain the salt and water.

The salt can be extracted from the salt solution by gradually evaporating off the water.

Example: preparing the soluble salt sodium chloride from hydrochloric acid (the acid) and sodium hydroxide (the alkali).

Practical: prepare a sample of ammonium sulphate from sulphuric acid and ammonia

- 1) Measure out 25 cm^3 of 2 mol dm^{-3} ammonia solution into a 100 cm^3 beaker.
- 2) Add 25 cm^3 of 1 mol dm^{-3} sulphuric acid solution to the same beaker.
- 3) Pour the resulting mixture into an evaporating dish and heat until most of the water has evaporated.
- 4) Leave the remaining solution for a week.
- 5) Use a spatula to remove any crystals in the beaker and dry them with a piece of filter paper.

Write an equation for the reaction between sulphuric acid and ammonia.

How many moles of sulphuric acid are used in the reaction?

How many moles of ammonia are used in the reaction?

Why is it important to use these exact quantities of the reactants?

<https://www.youtube.com/watch?v=Yh7z2ThG32o>

2) Physical Properties of acids, bases and salts

(a) Electrical conductivity in solution

Every ionic compound, either in the solid state or in solution, can be classified as an acid, base or salt, based on the cations and anions it contains:

If the cation is H^+ , the ionic compound is an **acid**.

If the anion is O^{2-} , OH^- , CO_3^{2-} or HCO_3^- , the ionic compound is a **base**.

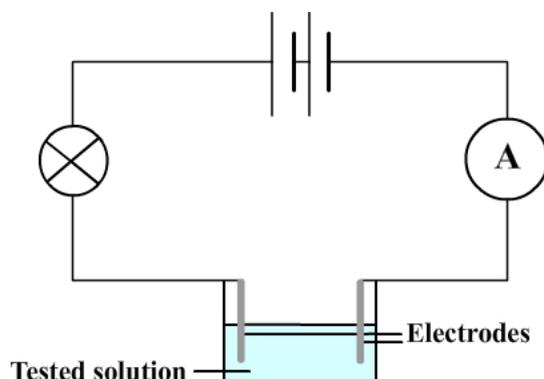
Otherwise the ionic compound is a **salt**.

Most ionic compounds dissolve in water to produce free ions. As a result, aqueous solutions of ionic compounds (acids, bases or salts) behave as electrical conductors.

Molten ionic compounds also contain ions which can move. As a result, molten acids, bases and salts also behave as electrical conductors.

Any substance which can conduct electricity in the molten or aqueous states is called an **electrolyte**.

The electrical conductivity of a solution can easily be measured using a simple circuit. A cell, bulb and ammeter are all connected in series. The circuit is completed by two electrodes dipped into the solution but otherwise not in direct contact. If the bulb lights, the solution contains electrolytes. The greater the ammeter reading, the greater the concentration of electrolytes in the solution.



Molten and aqueous solutions of acids, bases and salts are **electrolytes**. Any solution which shows electrical conductivity must therefore contain an acid, a base or a salt (or a mixture of them).

(b) Taste

Acids have a sour taste, caused by the hydrogen ions.

Alkalis have a soapy taste, caused by the hydroxide ions.

Many salts have a salty taste, mainly caused by sodium ions but also caused by some other cations.

(c) Deliquescence, Efflorescence and Hygroscopy

Many acids, bases and salts in the solid and aqueous states are hygroscopic, deliquescent, or efflorescent.

A **hygroscopic** substance is one which can absorb water from the atmosphere.

This happens because the ions in the solid are strongly attracted to water. Most concentrated solutions of acids and alkalis are hygroscopic and so are most solid salts if they are soluble in water.

Concentrated sulphuric acid, solid sodium chloride and solid sodium hydroxide are all hygroscopic.

Hygroscopic substances need to be stored in dry conditions to stop them from absorbing water.

Solids which are very hygroscopic are also deliquescent. A **deliquescent** solid is one which can absorb so much water from the atmosphere that it will eventually dissolve in the water it has absorbed.

Sodium hydroxide and calcium chloride are both deliquescent solids.

Hygroscopic and deliquescent solids are often used as **drying agents** because of their ability to absorb water from the atmosphere.

An efflorescent solid is one which contains water within its crystal structure but which will gradually release the water on exposure to the atmosphere. Efflorescence is the opposite of hygroscopy.

Many bases and salts, and some acids, in the solid state contain water molecules locked in between the ions in the crystal structure. These are known as hydrated solids. They often have a fixed chemical formula; the water is shown after the unit formula, separated by a dot:

Examples of hydrated solids	Unit formula
Hydrated copper sulphate	$\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$
Hydrated sodium carbonate	$\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$
Hydrated ethanedioic acid	$\text{H}_2\text{C}_2\text{O}_4 \cdot 2\text{H}_2\text{O}$

The water locked into the solid in these compounds is known as water of crystallisation.

Some of these compounds lose this water on exposure to air over a period of time. This is known as efflorescence. In many cases, there is an observable change as a result.

For example: hydrated copper sulphate is blue. When exposed to the atmosphere, it gradually loses its water and becomes anhydrous copper sulphate, which is white.

For example: wet plaster contains hydrated calcium sulphate ($\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$). When plaster is left to dry, it loses its water and hardens.

3) Strong and Weak Acids and Bases

(a) Strong and weak acids

Acids react with water to form solutions containing H^+ ions:



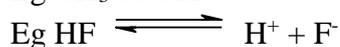
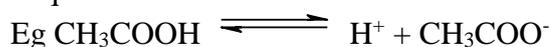
In hydrochloric acid and nitric acid, the molecule dissociates completely into the H^+ ion and the anion. There is no molecular HCl or HNO_3 remaining.

An acid which completely dissociates in water to form hydrogen ions is called a **strong** acid.

A strong acid is a substance which completely dissociates in water to form H^+ ions.

HCl , HNO_3 and H_2SO_4 are strong acids.

Most acids, however, only dissociate slightly in water. The H^+ ions, anions and undissociated molecules form a dynamic equilibrium:



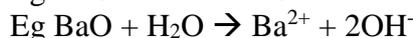
If the equilibrium position lies mainly on the left-hand side (ie only a small number of the molecules have dissociated), then the acid is called a **weak** acid.

A weak acid is a substance which only slightly dissociates in water to form H^+ ions.

All carboxylic acids, such as methanoic acid and ethanoic acid, are weak acids.

(b) Strong and weak bases

Bases react with water to form solutions containing OH⁻ ions:



In sodium hydroxide and barium hydroxide, the solid dissolves completely into the OH⁻ ion and the cation. There is no undissociated NaOH or BaO in solution.

A base which completely dissociates in water to form OH⁻ ions is called a **strong** base.

A strong base is a substance which completely dissociates in water to form OH⁻ ions.

NaOH, KOH and BaO are strong bases.

Some bases are highly soluble in water, but do not fully react with water to form hydroxide ions. The OH⁻ ions, cations and undissociated molecules form a dynamic equilibrium:



If the equilibrium position lies mainly on the left-hand side (ie only a small number of the molecules have reacted), then the bases is called a **weak** base.

A weak base is a substance which only slightly dissociates in water to form OH⁻ ions.

Ammonia and sodium carbonate are weak bases.

There are also some bases, such as Ca(OH)₂, which are only slightly soluble in water but which fully dissociate into OH⁻ ions once they are dissolves. Technically, these are strong bases but they are often classified as weak bases because they do not produce solutions containing a lot of OH⁻ ions due to their low solubility.

(c) Properties of strong acids and bases compared to weak acids and bases**(i) Electrical conductivity**

Strong acids and bases are completely dissociated into their ions:

A 1 mol dm⁻³ solution of HCl contains 1 mol dm⁻³ H⁺ and 1 mol dm⁻³ Cl⁻. It does not contain any undissociated HCl.

A 1 mol dm⁻³ solution of NaOH contains 1 mol dm⁻³ OH⁻ and 1 mol dm⁻³ Na⁺. It does not contain any undissociated NaOH.

These solutions both contain the same number of ions and so will both show the approximately the same electrical conductivity in aqueous solution.

Weak acids and bases are only slightly dissociated into their ions:

A 1 mol dm⁻³ solution of CH₃COOH contains mostly undissociated CH₃COOH. It contains a very small quantity of CH₃COO⁻ and H⁺ ions (around 0.01 mol dm⁻³).

A 1 mol dm⁻³ solution of NH₃ contains mostly undissociated NH₃. It contains a very small quantity of NH₄⁺ and OH⁻ ions (around 0.01 mol dm⁻³).

Solutions of weak acids and bases contain far fewer ions than solutions of strong acids and alkalis of the same concentration.

Solutions of weak acids and bases have a far lower electrical conductivity than solutions of strong acids and alkalis of the same concentration.

(ii) Enthalpy of neutralisation

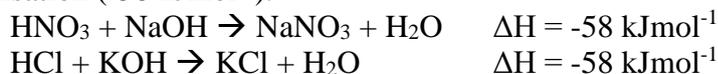
Solutions of strong acids contain fully dissociated H⁺ ions.

Solutions of strong bases contain fully dissociated OH⁻ ions.

When a solution of a strong acid reacts with a solution of a strong base, the H⁺ ions react with the OH⁻ ions to make H₂O: **H⁺ + OH⁻ → H₂O**. The remaining ions are not involved in the reaction.

This is true of a neutralisation reaction between any strong acid and any strong base – in all cases, an identical chemical reaction is taking place.

As a result, all reactions between strong acids and strong bases in solution have the same enthalpy of neutralisation (-58 kJ mol⁻¹).



Solutions of weak acids only contain small quantities of H^+ ions.

When weak acids are neutralised by strong bases, the OH^- ions react with the H^+ ions which quickly run out. As a result, the weak acid dissociates more, producing more H^+ ions (according to Le Chatelier's principle). This process continues until the weak acid has fully dissociated.

So during the neutralisation of weak acids, two reactions are taking place:

- the acid dissociates to form more H^+ : $\text{CH}_3\text{COOH} \rightarrow \text{CH}_3\text{COO}^- + \text{H}^+$
- the resulting H^+ reacts with the OH^- : $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$

The first reaction, the dissociation, is endothermic by around 2 kJmol^{-1} . The second reaction, the neutralisation, has a molar enthalpy change of -58 kJmol^{-1} . The overall molar enthalpy change of neutralisation is therefore -56 kJmol^{-1} .



Weak acids have a less exothermic enthalpy of neutralisation than strong acids, because the H^+ ion has to be removed from the acid first, which absorbs some of the energy.

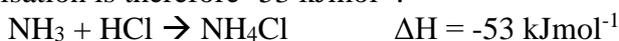
Solutions of weak bases only contain small quantities of OH^- ions.

When weak bases are neutralised by strong acids, the H^+ ions react with the OH^- ions which quickly run out. As a result, the weak base dissociates more, producing more OH^- ions (according to Le Chatelier's principle). This process continues until the weak base has fully dissociated.

So during the neutralisation of weak bases, two reactions are taking place:

- the base dissociates to form more OH^- : $\text{NH}_3 + \text{H}_2\text{O} \rightarrow \text{NH}_4^+ + \text{OH}^-$
- the resulting OH^- reacts with the H^+ : $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$

The first reaction, the dissociation, is endothermic by around 5 kJmol^{-1} . The second reaction, the neutralisation, has a molar enthalpy change of -58 kJmol^{-1} . The overall molar enthalpy change of neutralisation is therefore -53 kJmol^{-1} .



Weak bases have a less exothermic enthalpy of neutralisation than strong bases, because the OH^- ion has to be produced from the base first, which absorbs some of the energy.

Practical: compare the enthalpy of neutralisation of a strong base and a weak base by the same acid

- 1) Place a clean, dry polystyrene cup inside a glass beaker.
- 2) Pour 25 cm^3 of 2 mol dm^{-3} HCl into the cup using a measuring cylinder. Record the initial temperature of the solution.
- 3) Measure out 25 cm^3 of 2 mol dm^{-3} NaOH using a measuring cylinder and record the initial temperature of the solution. Hence deduce the average initial temperature of the solutions.
- 4) Add the NaOH solution to the HCl solution and stir continuously, until the temperature either remains constant for over a minute or starts to change in the opposite direction. Record the final temperature reached.
- 5) Deduce the maximum temperature change.
- 6) Repeat steps 1 – 5, but this time use 25 cm^3 of 2 mol dm^{-3} NH_3 instead of NaOH .
- 7) Which reaction gives you a greater temperature change? Can you explain why?

(d) Preparation of simple salts from acids and salts

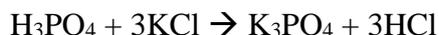
If a strong acid reacts with a salt of a weaker acid, the stronger acid will displace the weaker acid from its salt, by forcing the salt of the other acid to accept the hydrogen ion.

For example, sulphuric acid is a stronger acid than hydrochloric acid. As a result, if concentrated sulphuric acid is reacted with any salt of hydrochloric acid (eg potassium chloride), the sulphuric acid will give its hydrogen ion to the chloride ion, forming hydrochloric acid and a sulphate salt (eg potassium sulphate).



This method can be used to make salts of strong acids from any weaker acid.

The reaction can also be used to make phosphate salts from chloride salts by reacting them with concentrated phosphoric acid. This reaction works even though phosphoric acid is a weaker acid than hydrochloric acid:



This reaction works because of Le Chatelier's principle: the HCl formed, which is a gas, escapes, which prevents the reaction from going in the other direction. This can be a good way of making some phosphate salts.

4) Salt Hydrolysis

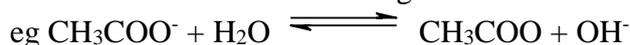
Salts can be formed from:

- strong acids and strong bases (eg NaNO_3 , K_2SO_4 and BaCl_2)
- weak acids and strong bases (eg CH_3COONa , KCN)
- strong acids and weak bases (eg NH_4Cl , NH_4NO_3)

Salts can also be formed from weak acids and weak bases but these tend not to be stable.

(i) Anion hydrolysis

In some cases, the anion in the salt (which is formed when the acid loses its H^+) can react partially with the water in the solution to reform the original acid:



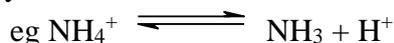
The resulting solution, therefore, contains OH^- ions as well as the salt.

This only happens when the anion comes from a weak acid. If an acid is reluctant to give up its H^+ ion in the first place, its anion is more likely to pick up an H^+ from surrounding water molecules to reform the original acid. Salts which do this are known as **basic salts**.

Salts made from weak acids, such as CH_3COONa and KCN , are basic salts.

(ii) Cation hydrolysis

In some cases, the cation in the salt (which is formed when the base loses its OH^- or accepts H^+) can react partially with the water in the solution to reform the original base:

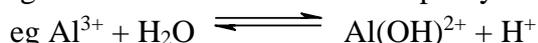


The resulting solution, therefore, contains H^+ ions as well as the salt.

This usually happens when the cation comes from a weak base. If a base is reluctant to accept an H^+ ion in the first place, its cation is more likely to release an H^+ and reform the original base. Salts which do this are known as **acidic salts**.

Salts made from weak bases, such as NH_4Cl , NH_4NO_3 , are acidic salts.

Some highly charged cations also react with water to form H^+ ions, in order to reduce their charge; they are accepting OH^- ions from the water and partly reforming the original hydroxide:



Salts which do this are also known as **acidic salts**.

Salts containing highly charged cations, such as AlCl_3 , are acidic salts.

In conclusion:

- salts made from weak acids can themselves react with water and behave like weak bases; they are known as **basic salts**.
- salts made from weak bases, or salts containing highly charged cations, can themselves react with water and behave like very bases; they are known as **acidic salts**.
- all other salts do not react with water and are known as **neutral salts**.

The tendency of a salt to itself react with water and show acidic or basic properties is known as **salt hydrolysis**.

Test Your Progress: Topic 5 Exercise 3

Predict whether the following salts will undergo hydrolysis. If they do, write an equation to show how one of the ions in the salt reacts with water to produce H^+ or OH^- ions.

- a) sodium nitrate
- b) ammonium nitrate
- c) sodium ethanoate
- d) potassium sulphate
- e) aluminium sulphate
- f) potassium cyanide

Test Your Progress: Topic 5 Exercise 4 (summary)

1. What are the Arrhenius and Bronsted-Lowry definitions of an acid?
2. What are the Arrhenius and Bronsted-Lowry definitions of a base?
3. What is an alkali?
4. What is a salt?
5. Describe the preparation of a salt from an acid and an insoluble base.
6. Describe the preparation of a salt from an acid and a soluble base.
7. Explain the meaning of the term amphoteric and give one example of an amphoteric substance.
8. Explain the meaning of the terms hygroscopic, deliquescent and efflorescent.
9. Give one important application of hygroscopic and deliquescent substances.
10. Explain the difference between a weak acid and a strong acid.
11. Describe two different experiments you could carry out which of two acids was stronger.
12. Describe the preparation of a salt from an acid and a salt