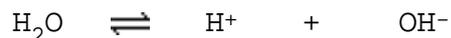


# ACIDS AND ALKALIS

Water ionises as follows:



Due to the reverse reaction, very few ions are actually present. The mixture contains mostly  $\text{H}_2\text{O}$  molecules. Since the rate of ionisation is equal to the rate at which the two ions recombine, the concentrations of  $\text{H}^+$ ,  $\text{OH}^-$ , and  $\text{H}_2\text{O}$  remain constant with time. We say that the concentrations have reached **equilibrium**.

Water is **NEUTRAL**: the concentrations of  $\text{H}^+$  and  $\text{OH}^-$  are equal.

In an **ACIDIC** solution the concentration of  $\text{H}^+$  is greater than the concentration of  $\text{OH}^-$ . A common household acid is vinegar.

In an **ALKALINE** solution the concentration of  $\text{OH}^-$  is greater than than the concentration of  $\text{H}^+$ . A common household alkali is Ammonia solution.

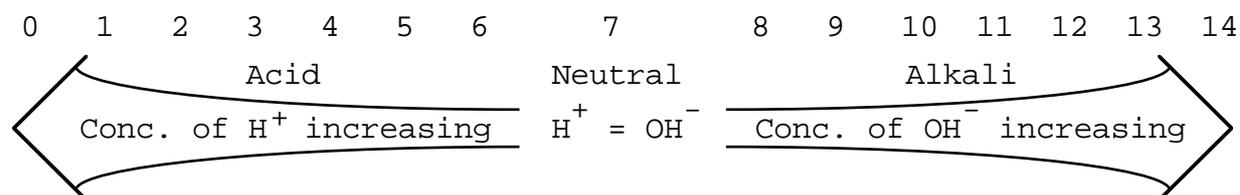
## THE pH SCALE

To measure whether a solution is acidic, alkaline or neutral we use the pH scale which ranges from below 0 to above 14.

Solution	pH
Acidic	less than 7
Alkali	greater than 7
Neutral	7

The higher the conc. of  $\text{H}^+$  in an acid the lower the pH.

The higher the conc. of  $\text{OH}^-$  in an alkali the higher the pH.



Diluting an acidic solution therefore lowers the conc. of  $\text{H}^+$ , reduces the acidity and raises the pH.

Diluting an alkaline solution lowers the conc. of  $\text{OH}^-$ , reduces the alkalinity and lowers the pH.

## FORMATION OF ACIDS

Most non-metal oxides, which dissolve in Water, form acids.

### Example 1 : Sulphur dioxide

This gas, produced on burning fossil fuels, reacts with rain water:



The Sulphurous acid is further oxidised by the Oxygen in the air to Sulphuric acid  $\text{H}_2\text{SO}_4$ .

### Example 2 : Nitrogen dioxide

This gas is produced in the car engine.  $\text{N}_2$  and  $\text{O}_2$  in the air, heated by the spark, react to form  $\text{NO}_2$ . The  $\text{NO}_2$  then reacts with rain water thus :



Both Sulphuric and Nitric acids are therefore present in rain water making it acidic - acid rain.

Acid rain damages buildings, metal structures, soils, plant and animal life.

## FORMATION OF ALKALIS

Most metal oxides, which dissolve in Water, form alkalis.

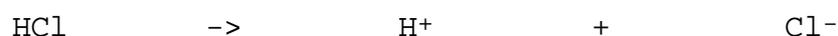
### Example : Calcium oxide



## STRONG AND WEAK ACIDS

Acids are proton ( $\text{H}^+$ ) donors.

A strong acid dissociates completely in Water  
e.g. Hydrogen chloride (Hydrochloric acid)



If we dissolve 1000 molecules of Hydrogen chloride in water we get 1000  $\text{H}^+$  ions !

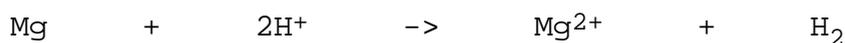
Sulphuric acid  $\text{H}_2\text{SO}_4$  and Nitric acid  $\text{HNO}_3$  are also strong acids.

A weak acid only partially dissociates e.g. Ethanoic acid



Dissociation is difficult. The two ions have a tendency to recombine. If we dissolve 1000 molecules of Ethanoic acid in water we only get 4 H<sup>+</sup> ions !!

Consequently, a solution of Ethanoic acid has a higher pH, lower electrical conductivity and a slower reaction with Magnesium :



than a solution of Hydrochloric acid of equal concentration.

## STRONG AND WEAK BASES

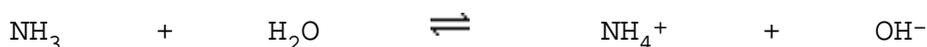
Bases are proton acceptors.

Strong bases (e.g. the oxide ion O<sup>2-</sup>) accept protons readily e.g. oxides react with Water forming alkalis:



There is little reversal ; solutions are strongly alkaline (pH 13-14) and good conductors of electricity since there are so many ions.

Weak bases (e.g. Ammonia NH<sub>3</sub>) accept protons less readily e.g. Ammonia reacts with Water forming an alkali:



There is a tendency for the two ions to recombine. Most of the OH<sup>-</sup> recombine so the solution is only weakly alkaline (pH 9-12) and a poor conductor of electricity since there are so few ions.

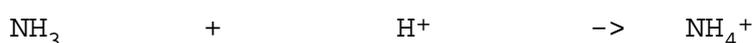
There are two other important bases :

Hydroxide ion            OH<sup>-</sup>

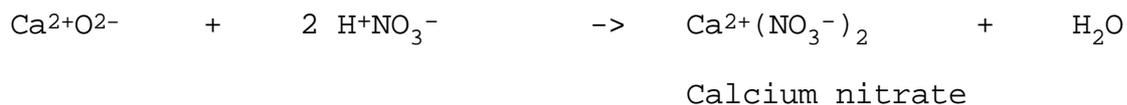
Carbonate ion            CO<sub>3</sub><sup>2-</sup>

## Reaction of bases with Acids

In general the reactions are :



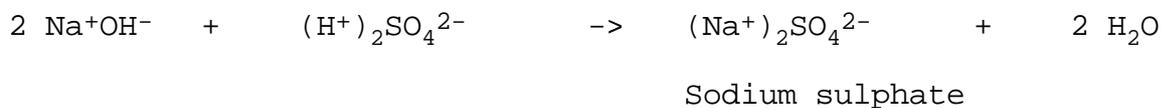
Ammonia reacts more slowly than the other three because it is a weaker base.

**Example 1**

Calcium nitrate is a useful fertiliser.

$\text{Ca}^{2+}$  and  $\text{NO}_3^-$  ions take no part in the reaction. They are called '**spectator ions**'.

The equation, written without spectator ions, is just:

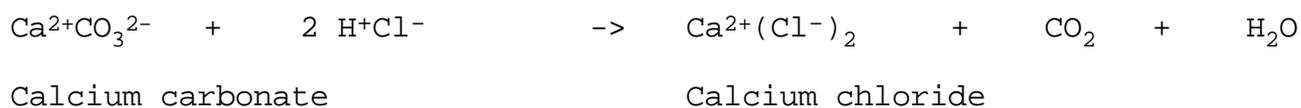
**Example 2**

The equation, written without spectator ions, is just:

**Example 3**

Ammonium nitrate is a useful fertiliser.

The equation, written without spectator ions, is just:

**Example 4**

Acid rain attacks rocks containing limestone and buildings made from marble, both impure forms of Calcium carbonate.

The equation, written without spectator ions, is just:



**EXPERIMENT**

Prepare Calcium chloride by the reaction in example 4.



When the reaction is over (no more bubbles of Carbon dioxide) the excess Calcium carbonate can be filtered off since it is **not** soluble in Water. The Calcium chloride is obtained by evaporation of the filtrate.

**NEUTRALISATION**

In the above four reactions, the  $H^+$  ion in the acid is removed and the pH moves up towards 7. We call this type of reaction '**neutralisation**'. Everyday examples are the treatment of acid indigestion with Magnesium hydroxide tablets and the use of Calcium hydroxide to reduce acidity in soil and lochs.

A '**salt**' is the compound formed by replacing the  $H^+$  ion in an acid by another positive ion. The salts in the above four examples are named: Calcium nitrate, Sodium sulphate, Ammonium nitrate and Calcium chloride.

**MOLARITY**

A solution's molarity, or concentration, is the number of moles of dissolved substance per 1000  $cm^3$  of solution e.g.

A 2M solution of Sodium fluoride contains 2 moles of Sodium fluoride in every 1000  $cm^3$  of solution.

1000  $cm^3$  of this solution would be prepared by first dissolving 2 moles of Sodium fluoride (84 g) in a little Water and then making up the volume to 1000  $cm^3$  in a volumetric flask.

**PROBLEM**

Calculate the number of moles of Sodium fluoride in 25  $cm^3$  of a 5 mol  $l^{-1}$  solution.

Answer:

$$\begin{aligned}
 \text{No. of moles of Sodium fluoride in } 1000 \text{ cm}^3 &= 5 \text{ mol} \\
 \Rightarrow \dots \dots \dots \dots \dots \dots \dots \dots \dots 25 \text{ cm}^3 &= \frac{5 \times 25}{1000} \\
 &= \underline{0.125 \text{ mol}}
 \end{aligned}$$

In general

$$\text{No. of moles} = \frac{MV}{1000}$$

where M = concentration (mol l<sup>-1</sup>)  
V = volume (cm<sup>3</sup>)

### PROBLEM

Calculate the weight of Glucose present in 200 cm<sup>3</sup> 3 mol l<sup>-1</sup> solution.

Answer:

$$\text{No. of moles of Glucose} = \frac{MV}{1000} = \frac{3 \times 200}{1000} = 0.6 \text{ mol}$$

$$\text{Wt. of 1 mole of Glucose } C_6H_{12}O_6 = 180 \text{ g}$$

$$\Rightarrow \dots \dots 0.6 \dots \dots = 180 \times 0.6 = \underline{108 \text{ g}}$$

### PROBLEM

20 cm<sup>3</sup> 0.1 mol l<sup>-1</sup> Hydrochloric acid were required to neutralise 25 cm<sup>3</sup> Sodium hydroxide.

Calculate the molarity of the Sodium hydroxide solution.

Answer:



$$\text{No. of moles } H^+Cl^- = \frac{MV}{1000}$$

$$= \frac{0.1 \times 20}{1000}$$

$$= 0.002 \text{ mol}$$

$$\Rightarrow \text{No. of moles } Na^+OH^- = 0.002 \text{ mol}$$

$$\Rightarrow \frac{MV}{1000} = 0.002$$

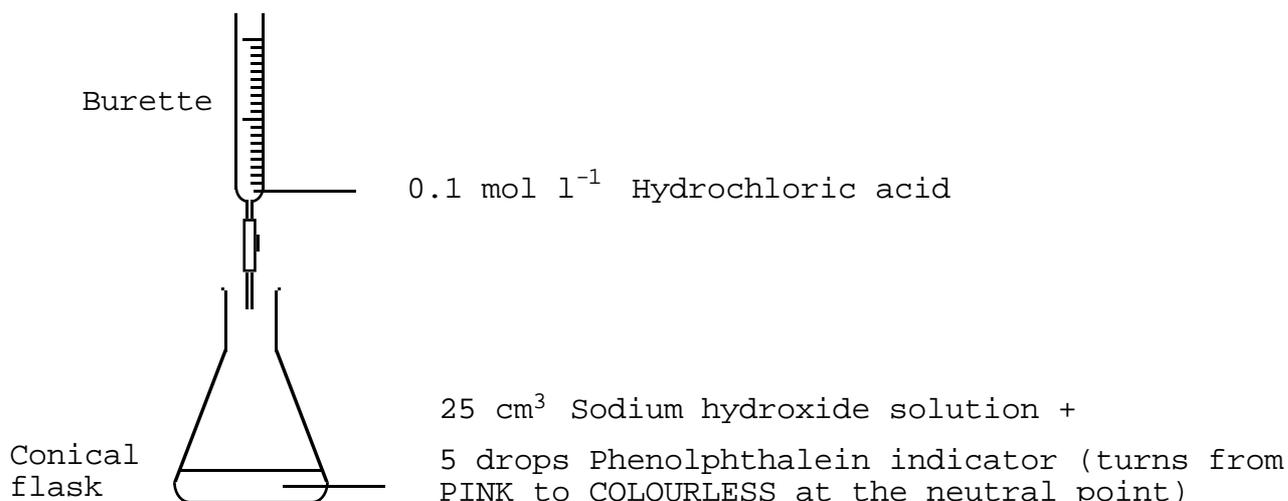
$$\Rightarrow M = \frac{0.002 \times 1000}{V}$$

$$= \frac{0.002 \times 1000}{25}$$

$$= \underline{0.08 \text{ mol l}^{-1}}$$

## TITRATION

The precise volumes required for this last problem would have been obtained by '**titration**'.



## PRECIPITATION REACTIONS

If we try making Lead(II) sulphate by the neutralisation methods described above .... it does not work! Lead(II) sulphate is insoluble in water. Insoluble salts **cannot** be prepared by the methods of neutralisation described above : the salt would coat reactant surfaces and prevent attack by the acid.

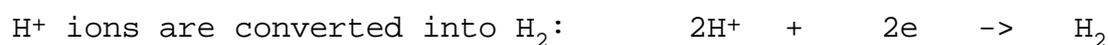
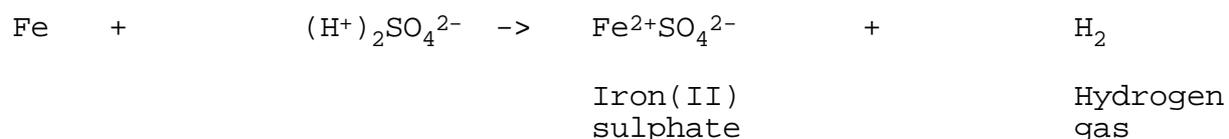
Insoluble salts are prepared by **PRECIPITATION**. Two solutions are mixed to give an insoluble product called a **precipitate**.

Lead(II) sulphate can be prepared by mixing solutions of Lead(II) nitrate and Sodium sulphate:



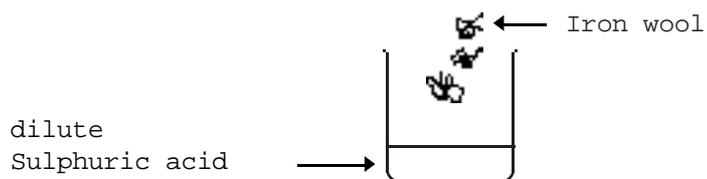
## REACTION OF METALS WITH ACIDS

Metals neutralise acids forming a salt and Hydrogen e.g.



Structures made from Iron are rapidly attacked by acid rain.

**Experiment** : Prepare Iron(II) sulphate.



The Hydrogen burns with a 'POP'. This is the test for Hydrogen. When the reaction is over (no more bubbles of Hydrogen) the excess metal can be filtered off and the Iron(II) sulphate obtained by evaporation of the filtrate.