

Why do some substances produce acidic or alkaline solutions?

- Some compounds react with water to produce acidic or alkaline solutions.
- Water must be present for a substance to act as an acid or as a base (usually at gcse level!).
- Acids in aqueous solution produce hydrogen H^+ ions. The H^+ ion is a proton. In water this proton is hydrated (associated with water and more correctly expressed as H_3O^+ (aq)) but H^+ (aq) is adequate here.
 - eg hydrochloric acid: $HCl_{(g)} + aq \Rightarrow H^+_{(aq)} + Cl^-_{(aq)}$
 - or sulphuric acid: $H_2SO_{4(l)} + aq \Rightarrow 2H^+_{(aq)} + SO_4^{2-}_{(aq)}$
- Alkalis in aqueous solution produce OH^- (aq) hydroxide ions.
 - eg sodium hydroxide: $NaOH_{(s)} + aq \Rightarrow Na^+_{(aq)} + OH^-_{(aq)}$
 - or calcium hydroxide: $Ca(OH)_{2(s)} + aq \Rightarrow Ca^{2+}_{(aq)} + 2OH^-_{(aq)}$
- When alkalis and acids react, the 'general word' and e.g. 'molecular formula' neutralisation equation might be ...
 - ACID + ALKALI \Rightarrow SALT + WATER ... eg
 - hydrochloric acid + sodium hydroxide \Rightarrow sodium chloride + water
 - $HCl_{(aq)} + NaOH_{(aq)} \Rightarrow NaCl_{(aq)} + H_2O_{(l)}$
 - BUT the ionic equation for ANY neutralisation is
 - $H^+_{(aq)} + OH^-_{(aq)} \Rightarrow H_2O_{(l)}$
 - and the remaining ions eg $Na^+_{(aq)}$ and $Cl^-_{(aq)}$ become the salt crystals $NaCl_{(s)}$ on evaporating the water.
- Acids can be defined as proton donors. A base can be defined as a proton acceptor (Bronsted-Lowry theory).
 - eg here the hydroxide ion is the base and accepts a proton from an acid.
 - $H^+_{(aq)} + OH^-_{(aq)} \Rightarrow H_2O_{(l)}$
 - or here the hydrogen chloride is the acid and the ammonia is the base when ammonium chloride is formed when the two gases are mixed. The acid hydrogen chloride donates a proton to the base ammonia. (note: no water present!)
 - $HCl_{(g)} + NH_{3(g)} \Rightarrow NH_4^+Cl^-_{(s)}$
 - or copper(II) oxide (base) + sulphuric acid (acid) \Rightarrow copper(II) sulphate + water
 - $CuO_{(s)} + H_2SO_{4(aq)} \Rightarrow CuSO_{4(aq)} + H_2O_{(l)}$
 - ionically it is: $Cu^{2+}O^{2-}_{(s)} + 2H^+_{(aq)} \Rightarrow Cu^{2+}_{(aq)} + H_2O_{(l)}$
 - Acids are characterised by having at least one replaceable hydrogen atom in forming a salt
 - eg for acid \Rightarrow sodium salt or salts (from Na_2O , $NaOH$, $NaHCO_3$ or Na_2CO_3)
 - $HNO_3 \Rightarrow NaNO_3$
 - only one nitrate salt possible from one replaceable 'hydrogen'
 - $HCl \Rightarrow NaCl$
 - only one chloride salt possible from one replaceable 'hydrogen'
 - $H_2SO_4 \Rightarrow NaHSO_4 \Rightarrow Na_2SO_4$
 - two sulphate salts possible from two replaceable 'hydrogens'
 - $H_3PO_4 \Rightarrow KH_2PO_4 \Rightarrow K_2HPO_4 \Rightarrow K_3PO_4$
 - three phosphate salts possible from three replaceable 'hydrogens'
 - Several scientists have made contributions to ionic and acid-base theory eg

- **Arrhenius in 1887**, was one of the first scientists to suggest that substances could split into free positive and negative ions when dissolved in water, the so called '**electrolytic dissociation**' giving rise to electrically conducting solutions. His theory was considered a bit revolutionary, and he was given a low rating for his PhD at Paris at first! - however the 'professors' recanted when other scientists decided it was a good idea and in 1903 he was awarded the Nobel Prize for his ionic theory work!
- **Lowry and Bronsted (1923)** took further the work of Arrhenius and applied ionic theory to the concept of acids and bases - that is, that acids and bases are proton donors and acceptors (see above). It should be noted that the work of Arrhenius took much longer to be accepted than the work of Lowry and Bronsted because there was no pre-existing (and proven) theory of ion formation.
- **Acids and alkalis are further classified by the extent of their ionisation in water.**
 - They are described as strong or weak depending their degree of ionisation in water.
 - **Do not confuse** the terms **weak and strong** to do with how far the 'molecules' become ionised in water ...
 - with the terms **dilute and concentrated**. These refer to the concentration of the acid or alkali in terms of how much of the original material is dissolved in water as measured by eg molarity.
 - You need to read on and then return here to clarify the points.
- **A strong acid or alkali is one that is that is nearly or completely 100% ionised in water** (not an equilibrium situation)
 - **examples of strong acids** are hydrochloric, sulphuric and nitric acids.
 - eg the maximum (or nearly) hydrogen ion concentration results ...
 - nitric acid is: $\text{HNO}_{3(l)} + \text{aq} \Rightarrow \text{H}^+_{(aq)} + \text{NO}_3^-_{(aq)}$
 - and sulphuric acid is: $\text{H}_2\text{SO}_{4(l)} + \text{aq} \Rightarrow 2\text{H}^+_{(aq)} + \text{SO}_4^{2-}_{(aq)}$
 - **examples of strong alkalis** (soluble strong bases) are sodium hydroxide or potassium hydroxide etc. (usually Group 1 or 2 hydroxides).
 - eg the maximum (or nearly) hydroxide ion concentration results ...
 - potassium hydroxide is: $\text{KOH}_{(s)} + \text{aq} \Rightarrow \text{K}^+_{(aq)} + \text{OH}^-_{(aq)}$
 - or strontium hydroxide is: $\text{Sr}(\text{OH})_{2(s)} + \text{aq} \Rightarrow \text{Sr}^{2+}_{(aq)} + 2\text{OH}^-_{(aq)}$
- **A weak acid or alkali is only partially ionised in water.**
 - **examples of weak acids** are ethanoic, citric and carbonic acids.
 - eg for **ethanoic about 2% ionises**, the equilibrium lies mainly to the un-ionised form on the left and for the weaker carbonic acid even less is ionised. So only a relatively low concentration of free hydrogen ions form
 - $\text{CH}_3\text{COOH}_{(aq)} \rightleftharpoons \text{CH}_3\text{COO}^-_{(aq)} + \text{H}^+_{(aq)}$
 - $\text{H}_2\text{CO}_{3(aq)} \rightleftharpoons \text{HCO}_3^-_{(aq)} + \text{H}^+_{(aq)}$
 - **An example of a weak alkali** (weak soluble base) is ammonia solution, about 2% changes to the ionic forms on the right. So only a relatively low concentration of free hydroxide ions form.
 - $\text{NH}_3(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{NH}_4^+_{(aq)} + \text{OH}^-_{(aq)}$
 - or sodium carbonate: $\text{CO}_3^{2-} + \text{H}_2\text{O}(l) \rightleftharpoons \text{HCO}_3^-_{(aq)} + \text{OH}^-_{(aq)}$
 - both of which, when dissolved in water, produce hydroxide ions giving an alkaline solution, despite the fact that **OH** doesn't appear in their formulae!

- You can **distinguish between strong and weak acids** of the same concentration by using the **pH scale** (or crudely, by the rate of reaction with metals). More on pH scale and indicators on [Acids, Bases, Salts](#) page.
- The pH is **dependent on the relative concentrations of the $H^+_{(aq)}$ and the $OH^-_{(aq)}$ concentrations.**
 - a **high $H^+_{(aq)}$** concentration means a **low pH** and low $OH^-_{(aq)}$ concentration, usually **strong acid**
 - lower $H^+_{(aq)}$ concentration means higher pH and higher $OH^-_{(aq)}$ concentration, less acid
 - a **high $OH^-_{(aq)}$** concentration means a **high pH** and low $H^+_{(aq)}$ concentration, usually **strong base/alkali**
 - lower $OH^-_{(aq)}$ concentration means lower pH and higher $H^+_{(aq)}$ concentration, less alkaline
- In general: **pH 1-2 strong acids**, **pH 3-6 weak acids**, **pH 7 neutral**, **pH 8-11 weak base/alkali**, **pH 12-14 strong base/alkali**
- If you put magnesium ribbon into 1 molar solutions of hydrochloric acid (strong, high $H^+_{(aq)}$ concentration) and ethanoic acid (weak, much lower $H^+_{(aq)}$ concentration), you can see the difference in the fast and slow 'fizzing' rates!
- **Neutralisation ionically is: $H^+_{(aq)} + OH^-_{(aq)} \Rightarrow H_2O_{(l)}$ (exothermic)**
 - **The pH of a solution, or determining the neutralisation point, can be measured with**
 - an indicator colour comparison card or indicator added to a titration
 - a pH meter which is calibrated with 'buffer solutions' of exactly know pH.
 - **When mixing an acid and alkali the neutralisation end-point can also be determined by**
 - the point of maximum temperature rise
 - the point