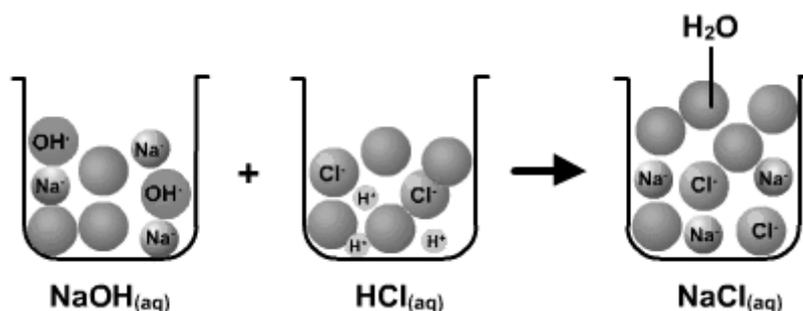


WASHINGTON LATIN PUBLIC CHARTER SCHOOL

HONORS CHEMISTRY 2019-20

UNIT 5A

CHEMICAL REACTIONS I – ACIDS, BASES AND SALTS



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Key words: Arrhenius, Bronsted-Lowry, acid, base, salt, neutralisation, hydrogen ion, hydroxide ion, proton donor, proton acceptor, alkali, pH, acidic, alkaline, neutral, strong acid, weak acid, strong base, weak base, indicator, phenolphthalein, methyl orange, universal indicator, quantitative analysis, titration, equivalence point, initial volume, final volume, titre volume, concordancy

# 1) Acids, Bases, Salts and Neutralisation

## (i) Acids

- The **Arrhenius** definition of an acid is a substance which releases hydrogen ( $H^+$ ) ions when dissolved in water; common examples are:
  - hydrogen chloride ( $HCl$ ) dissolves in water to produce hydrogen ( $H^+$ ) and chloride ( $Cl^-$ ) ions:  $HCl(g) \rightarrow H^+(aq) + Cl^-(aq)$ ; the resulting solution is known as **hydrochloric acid**
  - hydrogen nitrate ( $HNO_3$ ) dissolves in water to produce hydrogen ( $H^+$ ) and nitrate ( $NO_3^-$ ) ions:  $HNO_3(l) \rightarrow H^+(aq) + NO_3^-(aq)$ ; the resulting solution is known as **nitric acid**
  - hydrogen sulphate ( $H_2SO_4$ ) dissolves in water to produce hydrogen ( $H^+$ ) and nitrate ( $SO_4^{2-}$ ) ions:  $H_2SO_4(l) \rightarrow 2H^+(aq) + SO_4^{2-}(aq)$ ; the resulting solution is known as **sulphuric acid**
- Arrhenius acids therefore only exist in aqueous solution (aq)
- The **Bronsted-Lowry** definition of an acid is a proton ( $H^+$ ) donor; the  $H^+$  from the acid reacts with other ions and molecules in solution;  $HCl$ ,  $HNO_3$  and  $H_2SO_4$  are all Bronsted-Lowry acids because they lose their  $H^+$  ions (protons) when in solution
- The Bronsted-Lowry definition of an acid includes all substances which are classified as Arrhenius acids – eg  $HCl$ ,  $HNO_3$  and  $H_2SO_4$  are all Bronsted-Lowry acids because they lose their  $H^+$  ions when in solution; this definition also includes some substances which are not classified as Arrhenius acids (usually because they do not dissolve in water)

## (ii) Bases and Alkalis

- The **Bronsted-Lowry** definition of a base is a proton ( $H^+$ ) acceptor; the base accepts  $H^+$  ions (protons) from an acid
- The four most common types of base are:
  - $OH^-$  (hydroxide) ions, which accept  $H^+$  ions to form water:  $H^+ + OH^- \rightarrow H_2O$
  - $CO_3^{2-}$  (carbonate) ions, which accept  $H^+$  ions to form carbon dioxide and water:  $CO_3^{2-} + 2H^+ \rightarrow CO_2 + H_2O$
  - $O^{2-}$  (oxide) ions, which accept  $H^+$  ions to form water:  $H^+ + O^{2-} \rightarrow H_2O$
  - $NH_3$ , which accepts  $H^+$  ions to become  $NH_4^+$  (ammonium):  $NH_3 + H^+ \rightarrow NH_4^+$
- Some Bronsted-Lowry bases dissolve in water but most do not; when bases dissolve in water they generally release hydroxide ( $OH^-$ ) ions; a substance which dissolves in water to produce a solution containing  $OH^-$  ions is called an **alkali**; an alkali is therefore a soluble base
- The **Arrhenius** definition of a base is a substance which releases hydroxide ( $OH^-$ ) ions when dissolved in water; an Arrhenius base is therefore the same as an alkali
- All alkalis are bases, but some bases are not alkalis because they do not dissolve in water; a summary of which bases are soluble in water (and are therefore alkalis) can be shown in the table below:

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Alkalis (soluble bases)	Bases which are not alkalis because they are insoluble in water:
Sodium hydroxide, NaOH Potassium hydroxide, KOH Barium hydroxide, Ba(OH) <sub>2</sub> Sodium carbonate, Na <sub>2</sub> CO <sub>3</sub> Potassium carbonate, K <sub>2</sub> CO <sub>3</sub> Ammonia, NH <sub>3</sub>	Almost all other metal hydroxides, oxides and carbonates, eg magnesium hydroxide, Mg(OH) <sub>2</sub> calcium carbonate, CaCO <sub>3</sub> copper carbonate, CuCO <sub>3</sub> copper oxide, CuO magnesium oxide, MgO

### (iii) Salts and Neutralisation

- When an acid reacts with a base, it loses its H<sup>+</sup> ion; the H<sup>+</sup> 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<sup>+</sup> 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
- There are four main types of base, so there are four main types of neutralisation reaction in which a salt can be made from an acid and a base:
  - from metal hydroxides: acid + metal hydroxide → metal salt + water  
Eg  $\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$  (the salt produced is sodium chloride)
  - from metal oxides: acid + metal oxide → metal salt + water  
Eg  $2\text{HNO}_3 + \text{CaO} \rightarrow \text{Ca}(\text{NO}_3)_2 + \text{H}_2\text{O}$  (the salt produced is calcium nitrate)
  - from metal carbonates: acid + metal carbonate → metal salt + carbon dioxide + water  
Eg  $\text{H}_2\text{SO}_4 + \text{K}_2\text{CO}_3 \rightarrow \text{K}_2\text{SO}_4 + \text{CO}_2 + \text{H}_2\text{O}$  (the salt produced is potassium sulphate)
  - from ammonia: acid + ammonia → ammonium salt  
Eg  $\text{HCl} + \text{NH}_3 \rightarrow \text{NH}_4\text{Cl}$  (the salt produced is ammonium chloride)
- Useful applications of neutralization reactions include:
  - using CaO to neutralize soil which is too acidic
  - using Mg(OH)<sub>2</sub> (milk of magnesia) to neutralize indigestion caused by excess stomach acid (HCl)
  - using baking powder (NaHCO<sub>3</sub>) to neutralize bee stings (which are acidic)
- Most useful neutralization reactions involve insoluble bases – this is because insoluble bases are able to neutralize excess acid without themselves dissolving in water to produce harmful OH<sup>-</sup> ions

### (iv) Preparation of Soluble Salts by Neutralization

- The most suitable method of preparing a salt depends on the solubility of the salt and the solubility of the acid and the 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:
- Soluble salts are best prepared 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
  - copper sulphate is an example of a salt which can be prepared by this method:  
 $\text{CuO}(\text{s}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{CuSO}_4 + \text{H}_2\text{O}(\text{l})$  (CuO = insoluble base)

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- Group I and ammonium salts cannot be prepared by neutralization of insoluble bases as ammonia and all Group I bases are soluble; these salts can only be prepared by neutralizing acids with 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 salt can be extracted from the salt solution by gradually evaporating off the water
  - ammonium sulphate is an example of a salt which must be prepared by this method

## 2) Acidity, Alkalinity and the pH Scale

### (i) The pH Scale

- Water dissociates very slightly to produce  $H^+$  and  $OH^-$  ions:  $H_2O \rightleftharpoons H^+ + OH^-$ ; as a result, all aqueous solutions contain both  $H^+$  and  $OH^-$  ions; in pure water, the concentration of  $H^+$  and  $OH^-$  is around  $1 \times 10^{-7}$  mol/L
- Solutions which contain acids contain many more  $H^+$  ions than pure water; as a result, these solutions also contain fewer  $OH^-$  ions than pure water, because the extra  $H^+$  ions push the above equilibrium to the left; the higher the concentration of  $H^+$  ions, the lower the concentration of  $OH^-$  ions; any solution in which the concentration of  $H^+$  ions is greater than the concentration of  $OH^-$  ions is said to be an **acidic** solution
- Solutions which contain alkalis contain many more  $OH^-$  ions than pure water; as a result, these solutions also contain fewer  $H^+$  ions than pure water, because the extra  $OH^-$  ions push the above equilibrium to the left; the higher the concentration of  $OH^-$  ions, the lower the concentration of  $H^+$  ions; any solution in which the concentration of  $OH^-$  ions is greater than the concentration of  $H^+$  ions is said to be an **alkaline** solution
- A solution in which the concentration of  $H^+$  ions is equal to the concentration of  $OH^-$  ions, as is the case in pure water, is said to be a **neutral** solution

Concentration of $H^+$ ions (mol/L)	Concentration of $OH^-$ ions (mol/L)	Type of solution
0.1 ( $1 \times 10^{-1}$ )	$1 \times 10^{-13}$	acidic
0.01 ( $1 \times 10^{-2}$ )	$1 \times 10^{-12}$	acidic
0.001 ( $1 \times 10^{-3}$ )	$1 \times 10^{-11}$	acidic
$1 \times 10^{-5}$	$1 \times 10^{-9}$	acidic
$1 \times 10^{-7}$	$1 \times 10^{-7}$	neutral
$1 \times 10^{-9}$	$1 \times 10^{-5}$	alkaline
$1 \times 10^{-11}$	0.001 ( $1 \times 10^{-3}$ )	alkaline
$1 \times 10^{-13}$	0.1 ( $1 \times 10^{-1}$ )	alkaline

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- The level of acidity or alkalinity of a solution (ie the relative concentrations of  $H^+$  and  $OH^-$  ions) is measured on a scale called the **pH scale**; the pH of a solution is the negative logarithm (to base 10) of the hydrogen ion concentration:
  - If the  $H^+$  concentration is 0.1 (ie  $1 \times 10^{-1}$ ) mol/L, the pH of the solution is 1
  - If the  $H^+$  concentration is 0.001 (ie  $1 \times 10^{-3}$ ) mol/L, the pH of the solution is 3
  - If the  $H^+$  concentration is  $1 \times 10^{-7}$  mol/L, the pH of the solution is 7
  - If the  $H^+$  concentration is  $1 \times 10^{-11}$  mol/L, the pH of the solution is 11
  - If the  $H^+$  concentration is  $1 \times 10^{-13}$  mol/L, the pH of the solution is 13
  - the lower the pH, the more acidic the solution (the higher the concentration of  $H^+$  ions and the lower the concentration of  $OH^-$  ions); the most acidic solutions possible have a pH of around -1
  - the higher the pH, the more alkaline the solution (the higher the concentration of  $OH^-$  ions and the lower the concentration of  $H^+$  ions); the most alkaline solutions possible have a pH of around 15
  - neutral solutions have a pH of 7
- The relationship between pH, acidity and alkalinity is summarised in the table below:

pH	-1	1	3	5	7	9	11	13	15
Acidity	strongly acidic		slightly acidic		Neutral	slightly alkaline		Strongly alkaline	
$[H^+]$	very high		quite high		Normal	quite low		very low	
$[OH^-]$	very low		quite low		Normal	quite high		very high	

- Examples of the pH of common solutions are:
  - 1 mol/L HCl has a pH of 0
  - lemon juice and vinegar have a pH of approx. 2
  - orange juice has a pH of approx. 3
  - rainwater has a pH of approx. 6
  - pure water has a pH of 7
  - household bleach has a pH of approx. 12
  - 1 mol/L NaOH has a pH of 14

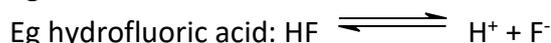
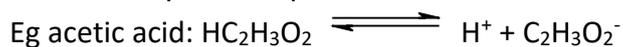
### 3) Strong and Weak Acids and Bases

#### (i) strong and weak acids

- Acids react with water to form solutions containing  $H^+$  ions:  
Eg  $HCl \rightarrow H^+ + Cl^-$       Eg  $HNO_3 \rightarrow H^+ + NO_3^-$ 
  - 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

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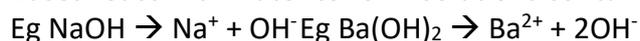
- Most acids, however, only dissociate slightly in water; the  $H^+$  ions, anions and undissociated molecules form a dynamic equilibrium:



- 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**

### (ii) 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 undissolved or undissociated  $NaOH$  or  $Ba(OH)_2$ ; 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 dissolved; these are also classified as weak bases because they do not produce solutions containing a lot of  $OH^-$  ions due to their low solubility

### (iii) Neutralizing strong and weak acids

- When weak acids are neutralised by strong bases:
  - the  $OH^-$  ions react with the  $H^+$  ions which have already dissociated; these 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
  - acetic acid, for example, only slightly dissociates in water but reacts completely with sodium hydroxide to form sodium acetate and water:  $HC_2H_3O_2 + NaOH \rightarrow NaC_2H_3O_2 + H_2O$
  - so weak acids will require exactly the same quantity of base as strong acids for complete neutralization:
    - 50 mL of 0.1 mol/L  $HCl$  will require 50 mL of 0.1 mol/L  $NaOH$  for complete neutralization
    - 50 mL of 0.1 mol/L  $HC_2H_3O_2$  will also require 50 mL of 0.1 mol/L  $NaOH$  for complete neutralization, even though it is a much weaker acid than  $HCl$

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- When a solution of a strong acid reacts with a solution of a strong base, the  $\text{H}^+$  ions react with the  $\text{OH}^-$  ions to make  $\text{H}_2\text{O}$ :  $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$ ; 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 neutralization reactions between strong acids and strong bases in solution have the same molar enthalpy change ( $-58 \text{ kJmol}^{-1}$ )
  - eg  $\text{HNO}_3 + \text{NaOH} \rightarrow \text{NaNO}_3 + \text{H}_2\text{O}$  ( $\Delta H = -58 \text{ kJmol}^{-1}$ );  $\text{HCl} + \text{KOH} \rightarrow \text{KCl} + \text{H}_2\text{O}$  ( $\Delta H = -58 \text{ kJmol}^{-1}$ )
- Solutions of weak acids only contain small quantities of  $\text{H}^+$  ions initially; the remaining molecules of weak acid dissociate gradually as the neutralization progresses; so during the neutralisation of weak acids, two reactions are taking place:
  - Step 1: the acid dissociates to form more  $\text{H}^+$ : eg  $\text{HC}_2\text{H}_3\text{O}_2 \rightarrow \text{H}^+ + \text{C}_2\text{H}_3\text{O}_2^-$   
this step is endothermic by around  $3 \text{ kJ/mol}$  (it depends on the acid)
  - Step 2: the resulting  $\text{H}^+$  reacts with the  $\text{OH}^-$ :  $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$ 
    - this step has a molar enthalpy change of  $-58 \text{ kJmol}^{-1}$  for all acids
    - the overall molar enthalpy change of neutralisation is therefore  $-55 \text{ kJ/mol}$
    - weak acids have a less exothermic molar enthalpy of neutralisation than strong acids, because the  $\text{H}^+$  ion has to be removed from the acid first, which absorbs some of the energy

### (iv) Other differences between strong and weak acids (and bases)

#### pH

- The pH of a solution depends on the concentration of  $\text{H}^+$  ions
  - strong acids, which dissociate fully into  $\text{H}^+$  ions, therefore produce more  $\text{H}^+$  ions than weak acids of the same molarity and have a lower pH
  - $0.1 \text{ mol/L HCl}$  has a pH of 1.0;  $0.1 \text{ mol/L HC}_2\text{H}_3\text{O}_2$  has a pH of 2.9
  - Strong bases have a higher pH than weak bases of the same molarity;  $0.1 \text{ mol/L NaOH}$  has a pH of 13.0;  $0.1 \text{ mol/L NH}_3$  has a pH of 11.1

#### Conductivity

- Strong acids and bases are completely dissociated into their ions:
  - a  $1 \text{ mol/L}$  solution of  $\text{HCl}$  contains  $1 \text{ mol/L H}^+$  and  $1 \text{ mol/L Cl}^-$ ; it does not contain any undissociated  $\text{HCl}$
  - a  $1 \text{ mol dm}^{-3}$  solution of  $\text{NaOH}$  contains  $1 \text{ mol dm}^{-3} \text{ OH}^-$  and  $1 \text{ mol dm}^{-3} \text{ Na}^+$  it does not contain any undissociated  $\text{NaOH}$
  - 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:
  - $1 \text{ mol/L}$  solution of  $\text{HC}_2\text{H}_3\text{O}_2$  contains mostly undissociated  $\text{HC}_2\text{H}_3\text{O}_2$ ; it contains a very small quantity of  $\text{C}_2\text{H}_3\text{O}_2^-$  and  $\text{H}^+$  ions (around  $0.01 \text{ mol/L}$ )
  - a  $1 \text{ mol/L}$  solution of  $\text{NH}_3$  contains mostly undissociated  $\text{NH}_3$ . It contains a very small quantity of  $\text{NH}_4^+$  and  $\text{OH}^-$  ions (around  $0.01 \text{ mol/L}$ )
  - solutions of weak acids and bases contain far fewer ions than solutions of strong acids and alkalis of the same molarity; they therefore have a much lower electrical conductivity than solutions of strong bases and alkalis of the same molarity

## Reactivity

- Weak acids react with bases much less quickly than strong acids, as the concentration of free H<sup>+</sup> ions in solution is much lower; similarly, weak bases react with acids much less quickly than strong bases
- The effect is most significant when acids are reacting with an insoluble base:
 

$\text{MgO(s)} + 2\text{HCl(aq)} \rightarrow \text{MgCl}_2\text{(aq)} + \text{H}_2\text{O(l)}$	(faster – HCl is a strong acid)
$\text{MgO(s)} + 2\text{HC}_2\text{H}_3\text{O}_2\text{(aq)} \rightarrow \text{Mg(C}_2\text{H}_3\text{O}_2)_2\text{(aq)} + \text{H}_2\text{O(l)}$	(slower - HC <sub>2</sub> H <sub>3</sub> O <sub>2</sub> is a weak acid)

## 4) Indicators

### (i) Pure Indicators

- An indicator is a substance which appears as one colour in acidic solutions and as a different colour in alkaline solutions; indicators can therefore be used to determine whether a solution is acidic or alkaline; most indicators are soluble organic acids or bases
- Most indicators are weak acids in which the acid and its anion have different colours:
 

HA	$\rightleftharpoons$	H <sup>+</sup>	+	A <sup>-</sup>
Colour 1				Colour 2

  - at low pH, the H<sup>+</sup> ions shift the position of equilibrium to the left and the indicator will appear as colour 1
  - at high pH, the [H<sup>+</sup>] is low so the equilibrium shifts to the right to produce more H<sup>+</sup> and the indicator will appear as colour 2
  - **hence indicators therefore appear as one colour in low pH, and a different colour in high pH**
- The pH range over which an indicator changes colour varies from indicator to indicator:

Indicator	Colour 1 (low pH)	Colour 2 (high pH)	pH of colour change
Methyl orange	Pink	Yellow	3 – 5
Phenolphthalein	Colourless	Pink	8 – 10

- The effect of acids and bases on indicators depends on the exact pH of the solution, and the pH at which the indicator changes colour:
  - Strongly acidic solutions with a pH of 3 or below will cause all three indicators to appear as Colour 1
  - Weakly acidic and neutral solutions with a pH of 5 – 7 will cause methyl orange to appear as Colour 2 but will cause phenolphthalein to appear as colour 1
  - Alkaline solutions with a pH of 10 or more will cause all three indicators to appear as Colour 2
- Single indicators, therefore, cannot give you the exact pH of a solution; they will simply tell you whether the pH is above or below a certain number

(ii) **Universal Indicator**

- Universal indicator is a mixture of several different indicators, each of which changes colour at a different pH; as a result, universal indicator appears at a different colour for every pH unit between 3 and 11:
  - at pH values of 3 and below, universal indicator appears red
  - between the pH values of 3 and 5, universal indicator appears orange
  - between the pH values of 5 and 6, universal indicator appears yellow
  - between the pH values of 6 and 8, universal indicator appears green
  - between the pH values of 8 and 9, universal indicator appears blue
  - between the pH values of 9 and 11, universal indicator appears indigo
  - at pH values of 11 and above, universal indicator appears violet
- Universal indicator is most useful for finding out the approximate pH of aqueous solutions; it can be used in paper form or in solution form; in most cases, a colour chart is also provided and the colour of the indicator can be compared to the chart to deduce the pH of the solution; it is not an accurate method of determining pH, but is very simple and easy to use

## 5) Quantitative Analysis - Acid-Base Titrations

- **Quantitative analysis** is the experimental determination of how much of a substance is present in a sample; it is possible to carry out the quantitative analysis of acids and bases using a technique called **titration**
- A **titration** is an experiment used to find out what exact volume of one solution is required to react completely with a fixed volume of another solution in a flask; it is the most common example of **volumetric analysis** (a type of quantitative analysis which measures volumes); from the titration result, you can deduce the concentration of the alkali if the concentration of the acid is known, or vice versa
- In most titrations, a fixed volume (usually 15, 20 or 25 mL) of alkali is prepared in a conical flask using a pipette and an indicator added to the alkali; the acid is then added gradually from the burette into the conical flask until the indicator changes colour (it is better to place the acid in the burette as alkalis can leave residue in the burette)
- Burettes are calibrated to measure the volume of solution delivered – the readings on the burette do not tell you the volume of solution actually in the burette; you must first record the burette reading before any solution has been delivered, and then record it again after the reaction is complete; the difference between these two volumes is the volume of solution delivered; this is also known as the titre volume:
 

**titre volume = final burette reading – initial burette reading**
- It is important to choose the indicator correctly – methyl orange cannot be used for titrations involving weak acids; phenolphthalein cannot be used for titrations involving weak alkalis
- It is common to repeat titrations several times in order to achieve consistent results. Titre volumes within 0.2 mL of each other are said to be **concordant**

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- Here is a simple worked example of how to use a titre volume to deduce the molarity of the unknown solution:

Question: In a titration 28.3 mL of a 0.10 mol/L solution of NaOH was required to change the color of the indicator in 25 mL of a solution of H<sub>2</sub>SO<sub>4</sub>. What was the molarity of the H<sub>2</sub>SO<sub>4</sub> solution?

Worked answer:

- Equation:  $\text{H}_2\text{SO}_4 + 2\text{NaOH} \rightarrow \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O}$
  - Moles of NaOH used in titration =  $28.3/1000 \times 0.1 = 2.8 \times 10^{-3}$
  - 2:1 ratio so moles of H<sub>2</sub>SO<sub>4</sub> present in conical flask =  $2.8 \times 10^{-3}/2 = 1.4 \times 10^{-3}$
  - so molarity of H<sub>2</sub>SO<sub>4</sub> =  $1.4 \times 10^{-3}/25 \times 1000 = 0.056 \text{ mol/L}$
- In many cases, the solution in the pipette needs to be prepared by dissolving a solid sample or by diluting a more concentrated solution; this requires a standard solution of the substance to be prepared first