

## 2.1.4 Acids

An acid releases  $H^+$  ions in aqueous solution

The most common strong acids are :  
Hydrochloric ( HCl), sulfuric ( $H_2SO_4$ ) and nitric ( $HNO_3$ ) acid;  
Ethanoic acid  $CH_3COOH$  is a weak acid

There are several definitions of acids we use in chemistry. One we use later in the course is the Bronsted- Lowry acid which is defined as a proton ( $H^+$ ) donor

### Bases and Alkalis

Bases neutralise acids. Common bases are metal oxides, metal hydroxides and ammonia.

The Bronsted- Lowry base is defined as a proton ( $H^+$ ) acceptor

An Alkali is a soluble base that releases  $OH^-$  ions in aqueous solution;  
The most common alkalis are sodium hydroxide (NaOH), potassium hydroxide (KOH) and aqueous ammonia ( $NH_3$ )

A base readily accepts  $H^+$  ions from an acid:  
eg  $OH^-$  ions accepts an  $H^+$  ion forming  $H_2O$   
 $NH_3$  accepts an  $H^+$  ion forming  $NH_4^+$  ion

### Strong and weak acids

Strong acids **completely dissociate** when dissolved in water

Weak acids only **slightly dissociate** when dissolved in water, giving an equilibrium mixture

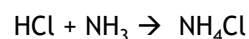
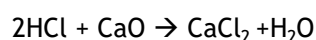
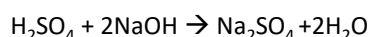
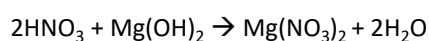
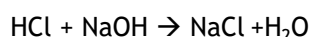
### Neutralisation Reactions

Neutralisation reaction form salts

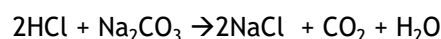
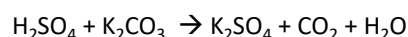
A **Salt** is formed when the  $H^+$  ion of an acid is replaced by a metal ion or an ammonium ion

### Common Neutralisation Reaction Equations

**ACID + BASE → SALT + WATER**



**Acid + Carbonate → Salt + Water + Carbon Dioxide**



Observations : In carbonate reactions there will be Effervescence due to the  $CO_2$  gas evolved and the solid carbonate will dissolve

# Titration

The method for carrying out the titration

- rinse equipment** (burette with acid, pipette with alkali, conical flask with distilled water)
- pipette 25 cm<sup>3</sup> of alkali into conical flask**
- touch surface of alkali with pipette** ( to ensure correct amount is added)
- adds acid solution from burette**
- make sure the jet space** in the burette **is filled** with acid
- add a few drops of indicator** and refer to colour change at end point
- phenolphthalein [pink (alkali) to colourless (acid): end point pink colour just disappears] [use if NaOH is used]
- methyl orange [yellow (alkali) to red (acid): end point orange] [use if HCl is used]
- use a white tile underneath the flask to help observe the colour change
- add acid to alkali whilst **swirling the mixture** and **add acid dropwise at end point**
- note burette reading** before and after addition of acid
- repeats titration** until **at least 2 concordant results** are obtained- two readings within 0.1 of each other

## Working out average titre results

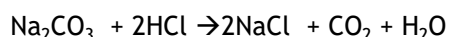
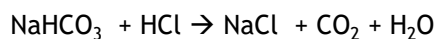
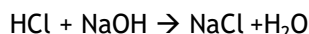
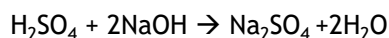
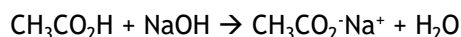
Only make an average of the concordant titre results

If **2 or 3 values are within 0.10cm<sup>3</sup>** and therefore **concordant** or close then we can say results are accurate and **repeatable** and **the titration technique is good/ consistent**

## Recording results

- Results should be clearly recorded in a table
- Result should be recorded in full (i.e. both initial and final readings)
- Record titre volumes to 2dp (0.05 cm<sup>3</sup>)**

## Common Titration Equations



## Titration mixtures

If titrating a mixture to work out the concentration of an active ingredient it is necessary to consider if the mixture contains other substances that have acid base properties.

If they don't have acid base properties we can titrate with confidence.

## Testing batches

In quality control it will be necessary to do titrations/testing on several samples as the amount/concentration of the chemical being tested may vary between samples.

## Safely dealing with excess acid

Sodium hydrogen carbonate (NaHCO<sub>3</sub>) and calcium carbonate (CaCO<sub>3</sub>) are good for neutralising excess acid in the stomach or acid spills because they are not corrosive and will not cause a hazard if used in excess. They also have no toxicity if used for indigestion remedies but the CO<sub>2</sub> produced can cause wind. Magnesium hydroxide is also suitable for dealing with excess stomach acid as it has low solubility in water and is only weakly alkaline so not corrosive or dangerous to drink (unlike the strong alkali sodium hydroxide). It will also not produce any carbon dioxide gas.

## Safety precautions

**Acids and alkalis are corrosive (at low concentrations acids are irritants)**

**Wear eye protection and gloves**

If spilled immediately wash affected parts after spillage

If substance is unknown treat it as potentially toxic and wear gloves.

If the jet space is not filled properly prior to commencing the titration it will lead to errors if it then fills during the titration, leading to a larger than expected titre reading.

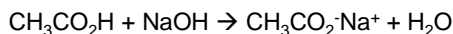
A conical flask is used in preference to a beaker because it is easier to swirl the mixture in a conical flask without spilling the contents.

Indicators are generally weak acids so we only add a few drops of them. If too much is added it will affect the titration result

Distilled water can be added to the conical flask during a titration to wash the sides of the flask so that all the acid on the side is washed into the reaction mixture to react with the alkali. It does not affect the titration reading as water does not react with the reagents or change the number of moles of acid added.

## More complicated titration calculations- taking samples

**Example 1:** A 25.0cm<sup>3</sup> sample of vinegar was diluted in a 250cm<sup>3</sup> volumetric flask. This was then put in a burette and 23.10cm<sup>3</sup> of the diluted vinegar neutralised 25.0 cm<sup>3</sup> of 0.100 M NaOH. What is the concentration of the vinegar in gdm<sup>-3</sup> ?



Step 1: work out amount, in mol, of sodium hydroxide  
 amount = conc x vol  
 = 0.10 x 0.025  
 = 0.00250 mol

Step 2: use balanced equation to give moles of CH<sub>3</sub>CO<sub>2</sub>H  
 1 moles NaOH : 1 moles CH<sub>3</sub>CO<sub>2</sub>H  
 So 0.00250 NaOH : 0.00250 moles CH<sub>3</sub>CO<sub>2</sub>H

Step 3 work out concentration of diluted CH<sub>3</sub>CO<sub>2</sub>H in 23.1 (and 250 cm<sup>3</sup>) in moldm<sup>-3</sup>

conc = amount/Volume  
 = 0.00250 / 0.0231  
 = 0.108 mol dm<sup>-3</sup>

Step 4 work out concentration of original concentrated CH<sub>3</sub>CO<sub>2</sub>H in 25cm<sup>3</sup> in moldm<sup>-3</sup>

conc = 0.108 x 10 = 1.08 mol dm<sup>-3</sup>

Step 5 work out concentration of CH<sub>3</sub>CO<sub>2</sub>H in original concentrated 25 cm<sup>3</sup> in gdm<sup>-3</sup>

conc in gdm<sup>-3</sup> = conc in mol dm<sup>-3</sup> x Mr  
 = 1.08 x 60 = 64.8 g dm<sup>-3</sup>

**Example 2.** An unknown metal carbonate reacts with hydrochloric acid according to the following equation.  
 $\text{M}_2\text{CO}_3(\text{aq}) + 2\text{HCl}(\text{aq}) \rightarrow 2\text{MCl}(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$   
 A 3.96 g sample of M<sub>2</sub>CO<sub>3</sub> was dissolved in distilled water to make 250 cm<sup>3</sup> of solution. A 25.0 cm<sup>3</sup> portion of this solution required 32.8 cm<sup>3</sup> of 0.175 mol dm<sup>-3</sup> hydrochloric acid for complete reaction. Calculate the Mr of M<sub>2</sub>CO<sub>3</sub> and identify the metal M

1. Calculate the number of moles of HCl used

amount = conc x vol  
 = 0.175 x 0.0328  
 = 0.00574 mol

2. Work out number of moles of M<sub>2</sub>CO<sub>3</sub> in 25.0 cm<sup>3</sup> put in conical flask

use balanced equation to give moles of M<sub>2</sub>CO<sub>3</sub>  
 2 mol HCl : 1 mol M<sub>2</sub>CO<sub>3</sub>  
 So 0.00574 NaOH : 0.00287 moles M<sub>2</sub>CO<sub>3</sub>

3. Calculate the number of moles M<sub>2</sub>CO<sub>3</sub> acid in original 250 cm<sup>3</sup> of solution

Moles in 250cm<sup>3</sup> = 0.00287 x 10  
 = 0.0287

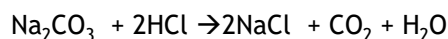
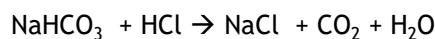
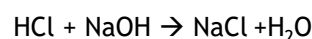
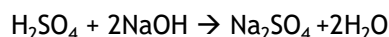
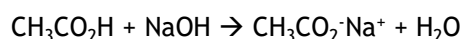
4. work out the Mr of M<sub>2</sub>CO<sub>3</sub>

Mr = mass / amount  
 = 3.96 / 0.0287 = 138.0

5. Work out Ar of M =  $\frac{138-12-16 \times 3}{2}$

Ar of M = 39 M = potassium

## Common Titration Equations



## Example 3

950 mg of impure calcium carbonate tablet was crushed. 50.0 cm<sup>3</sup> of 1.00 mol dm<sup>-3</sup> hydrochloric acid, an excess, was then added. After the tablet had reacted, the mixture was transferred to a volumetric flask. The volume was made up to exactly 100 cm<sup>3</sup> with distilled water. 10.0 cm<sup>3</sup> of this solution was titrated with 11.1cm<sup>3</sup> of 0.300 mol dm<sup>-3</sup> sodium hydroxide solution.

What is the percentage of CaCO<sub>3</sub> by mass in the tablet?

1. Calculate the number of moles of sodium hydroxide used

amount = conc x vol  
 = 0.30 x 0.0111  
 = 0.00333 mol

2. Work out number of moles of hydrochloric acid left in 10.0 cm<sup>3</sup>

use balanced equation to give moles of HCl  
 1 mol NaOH : 1 mol HCl  
 So 0.00333 NaOH : 0.00333 moles HCl

3. Calculate the number of moles of hydrochloric acid left in 100 cm<sup>3</sup> of solution

Moles in 100cm<sup>3</sup> = 0.00333 x 10  
 = 0.0333

4. Calculate the number of moles of HCl that reacted with the indigestion tablet.

In original HCl 50.0 cm<sup>3</sup> of 1.00 mol dm<sup>-3</sup> there is 0.05 moles

moles of HCl that reacted with the indigestion tablet. = 0.05 - 0.0333 = 0.0167

5 Use balanced equation to give moles of CaCO<sub>3</sub>

$\text{CaCO}_3(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{CaCl}_2(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$   
 2 mol HCl : 1 mol CaCO<sub>3</sub>  
 So 0.0167 HCl : 0.00835 moles CaCO<sub>3</sub>

6. work out the mass of CaCO<sub>3</sub> in original tablet

mass = amount x Mr  
 = 0.00835 x 100 = 0.835 g

percentage of CaCO<sub>3</sub> by mass in the tablet =  $\frac{0.835}{0.950} \times 100 = 87.9\%$

# Uncertainty

## Readings and Measurements

### Readings

the values found from a single judgement when using a piece of equipment

### Measurements

the values taken as the difference between the judgements of two values (e.g. using a burette in a titration)

In general, if uncertainty is not indicated on apparatus, the following assumptions are made:

For an analogue scale-

The uncertainty of a reading (one judgement) is at least  $\pm 0.5$  of the smallest scale reading.

The uncertainty of a measurement (two judgements) is at least  $\pm 1$  of the smallest scale reading.

- If the apparatus has a digital scale, the uncertainty is  $\pm$  the resolution of the apparatus in each measurement

### Calculating Apparatus Uncertainties

Each type of apparatus has a sensitivity uncertainty

- balance  $\pm 0.001$  g (if a 3 d.p. balance)
- volumetric flask  $\pm 0.1$  cm<sup>3</sup>
- 25 cm<sup>3</sup> pipette  $\pm 0.1$  cm<sup>3</sup>
- burette (start & end readings and end point)  $\pm 0.10$  cm<sup>3</sup>

Calculate the percentage error for each piece of equipment used by

$$\% \text{ uncertainty} = \pm \frac{\text{uncertainty}}{\text{Measurement made on apparatus}} \times 100$$

e.g. for burette

$$\% \text{ uncertainty} = 0.10 / \text{average titre result} \times 100$$

To calculate the maximum percentage apparatus uncertainty in the final result add all the individual equipment uncertainties together.

### Uncertainty of a measurement using a burette.

If the burette used in the titration had an uncertainty for each reading of  $\pm 0.05$  cm<sup>3</sup> then during a titration two readings would be taken so the uncertainty on the titre volume would be  $\pm 0.10$  cm<sup>3</sup>.

To decrease the apparatus uncertainties you can either decrease the sensitivity uncertainty by using apparatus with a greater resolution (finer scale divisions) or you can increase the size of the measurement made.

### Reducing uncertainties in a titration

Replacing measuring cylinders with pipettes or burettes which have lower apparatus uncertainty will lower the % uncertainty.

To reduce the % uncertainty in a burette reading it is necessary to make the titre a larger volume. This could be done by: increasing the volume and concentration of the substance in the conical flask or by decreasing the concentration of the substance in the burette.

If looking at a series of measurements in an investigation the experiments with the smallest readings will have the highest experimental uncertainties.

### Reducing uncertainties in measuring mass

Using a balance that measures to more decimal places or using a larger mass will reduce the % uncertainty in weighing a solid.

Weighing sample before and after addition and then calculating difference will ensure a more accurate measurement of the mass added.

### Calculating the percentage difference between the actual value and the calculated value

If we calculated an *M<sub>r</sub>* of 203 and the real value is 214, then the calculation is as follows:

Calculate difference 214-203 = 11

$$\% = 11/214 \times 100$$

$$= 5.14\%$$

If the %uncertainty due to the apparatus < percentage difference between the actual value and the calculated value then there is a discrepancy in the result due to other errors.

If the %uncertainty due to the apparatus > percentage difference between the actual value and the calculated value then there is no discrepancy and all errors in the results can be explained by the sensitivity of the equipment.