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 a defined as a proton (H^+) donor

Bases and Alkalis

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

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)

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

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

$HCl + NaOH \rightarrow NaCl + H_2O$

$2HNO_3 + Mg(OH)_2 \rightarrow Mg(NO_3)_2 + 2H_2O$

$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$

$2HCl + CaO \rightarrow CaCl_2 + H_2O$

$HCl + NH_3 \rightarrow NH_4Cl$

Acid + Carbonate → Salt + Water + Carbon Dioxide

$H_2SO_4 + K_2CO_3 \rightarrow K_2SO_4 + CO_2 + H_2O$

$2HCl + Na_2CO_3 \rightarrow 2NaCl + CO_2 + H_2O$

$2HCl + CaCO_3 \rightarrow CaCl_2 + CO_2 + H_2O$

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

Titrations

The method for carrying out the titration

- **rinse equipment** (burette with acid, pipette with alkali, conical flask with distilled water)
- **pipette 25 cm^3 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

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

Working out average titre results
Only make an average of the concordant titre results

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

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.

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^3)**

Common Titration Equations

$CH_3CO_2H + NaOH \rightarrow CH_3CO_2^-Na^+ + H_2O$

$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$

$HCl + NaOH \rightarrow NaCl + H_2O$

$NaHCO_3 + HCl \rightarrow NaCl + CO_2 + H_2O$

$Na_2CO_3 + 2HCl \rightarrow 2NaCl + CO_2 + H_2O$

Titrating 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 ($NaHCO_3$) and calcium carbonate ($CaCO_3$) 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_2 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.

More complicated titration calculations- taking samples

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

$CH_3CO_2H + NaOH \rightarrow CH_3CO_2^-Na^+ + H_2O$

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_3CO_2H
1 moles NaOH : 1 moles CH_3CO_2H
So 0.00250 NaOH : 0.00250 moles CH_3CO_2H

Step 3 work out concentration of diluted CH_3CO_2H in 23.1 (and 250 cm^3)in moldm^{-3}

conc= amount/Volume

= 0.00250 / 0.0231

= 0.108 mol dm^{-3}

Step 4 work out concentration of original concentrated CH_3CO_2H in 25cm^3 in moldm^{-3}

conc = 0.108 x 10 = 1.08 mol dm^{-3}

Step 5 work out concentration of CH_3CO_2H in original concentrated 25 cm^3 in gdm^{-3}

conc in gdm^{-3} = conc in mol dm^{-3} x *M*r

= 1.08 x 60 = 64.8 g dm^{-3}

Example 2. An unknown metal carbonate reacts with hydrochloric acid according to the following equation.
$M_2CO_3(aq) + 2HCl(aq) \rightarrow 2MCl(aq) + CO_2(g) + H_2O(l)$
A 3.96 g sample of M_2CO_3 was dissolved in distilled water to make 250 cm^3 of solution. A 25.0 cm^3 portion of this solution required 32.8 cm^3 of 0.175 mol dm^{-3} hydrochloric acid for complete reaction. Calculate the *M*r of M_2CO_3 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_2CO_3 in 25.0 cm^3 put in conical flask
use balanced equation to give moles of M_2CO_3
2 mol HCl : 1 mol M_2CO_3
So 0. 00574 NaOH : 0.00287 moles M_2CO_3

3. Calculate the number of moles M_2CO_3 acid in original 250 cm^3 of solution

Moles in 250cm^3 = 0.00287 x10
=0.0287

4. work out the *M*r of M_2CO_3

*M*r= mass / amount

= 3.96/ 0.0287= 138.0

5. Work out *A*r of M = $\frac{\underline{(138-12- 16x3)}}{2}$

Ar of M = 39 M= potassium

Common Titration Equations

$CH_3CO_2H + NaOH \rightarrow CH_3CO_2^-Na^+ + H_2O$

$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$

$HCl + NaOH \rightarrow NaCl + H_2O$

$NaHCO_3 + HCl \rightarrow NaCl + CO_2 + H_2O$

$Na_2CO_3 + 2HCl \rightarrow 2NaCl + CO_2 + H_2O$

Example 3

950 mg of impure calcium carbonate tablet was crushed. 50.0 cm^3 of 1.00 mol dm^{-3} 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^3 with distilled water. 10.0 cm^3 of this solution was titrated with 11.1cm^3 of 0.300 mol dm^{-3} sodium hydroxide solution.
What is the percentage of $CaCO_3$ 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^3
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^3 of solution

Moles in 100cm^3 = 0.00333 x10
=0.0333

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

In original HCl 50.0 cm^3 of 1.00 mol dm^{-3} there is 0.05moles

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

5 Use balanced equation to give moles of $CaCO_3$
$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$
2 mol HCl : 1 mol $CaCO_3$
So 0.0167 HCl : 0.00835 moles $CaCO_3$

6. work out the mass of $CaCO_3$ in original tablet

mass= amount x *M*r

= 0.00835 x 100 = 0.835 g

percentage of $CaCO_3$ by mass in the tablet = 0.835/0.950 x100

= 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 ±0.5 of the smallest scale reading.
The uncertainty of a measurement (two judgements) is at least ±1 of the smallest scale reading.
- If the apparatus has a digital scale, the uncertainty is ± the resolution of the apparatus in each measurement

Calculating Apparatus Uncertainties

Each type of apparatus has a sensitivity uncertainty

- balance ± 0.001 g (if a 3 d.p. balance)
- volumetric flask ± 0.1 cm^3
- 25 cm^3 pipette ± 0.1 cm^3
- burette (start & end readings and end point) ± 0.10 cm^3

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
% uncertainty = 0.10/average titre result x100

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 +/– 0.05 cm^3 then during a titration two readings would be taken so the uncertainty on the titre volume would be +/– 0.10 cm^3.

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.

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 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.

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*r of 203 and the real value is 214, then the calculation is as follows:
Calculate difference 214-203 = 11
% = 11/214 x100
=5.41%

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.