FACTFILE:
GCE CHEMISTRY

ACID BASE TITRATIONS

Rewarding Learning

Acid base titrations

Students should be able to:

1.9.1 demonstrate understanding of the concept of weak and strong acids and bases in terms of dissociation of hydrogen ions and hydroxide ions;

1.9.2 demonstrate understanding of the techniques and procedures used when experimentally carrying out acid-base titrations involving strong acid/strong base, strong acid/weak base and weak acid/strong base, for example determination of the degree of hydration in a sample of sodium carbonate and analysis of vinegar;

1.9.3 select the correct indicator for each type of titration and recall the colour changes of phenolphthalein and methyl orange at the end point;

1.9.4 identify uncertainties in measurements and calculate the uncertainty when two burette readings are used to calculate a titre value;

1.9.5 select appropriate titration data, ignoring outliers, in order to calculate mean titres;

1.9.6 calculate concentrations and volumes for titration calculations;

1.9.7 demonstrate knowledge of the term molarity, M, and the units of concentration, for example mol dm^{-3}, g dm^{-3};

1.9.8 describe the techniques and procedures used to prepare a standard solution of required concentration.

Solution Concentration

A solution consists of a substance that is dissolved – the **solute** – and the substance that is doing the dissolving – the **solvent**. Solutions may be coloured or colourless, but they are always clear and never cloudy.

For solid substances the chemical amount of the substance (in moles) can be calculated from the mass of the substance and its molar mass. However, most chemical reactions take place in solution. If we want to know the amount of a substance in solution, then we must know the concentration of the solution and its volume.

The concentration of a solution is the number of moles or mass present in a stated volume. Concentration in chemistry can be expressed as the molar concentration (molarity, M) of a solution. Molarity is the concentration in mol dm^{-3} expressed using M.

$$\text{moles (n)} = \frac{\text{volume (cm}^3\text{)} \times \text{concentration (mol dm}^{-3}\text{)}}{1000}$$

$$\text{concentration (mol dm}^{-3}\text{)} = \frac{\text{moles (n)} \times 1000}{\text{volume (cm}^3\text{)}}$$

$$\text{volume (cm}^3\text{)} = \frac{\text{moles (n)} \times 1000}{\text{concentration (mol dm}^{-3}\text{)}}$$

Remember that: 1000 cm^3 = 1 dm^3 = 10^{-3} m^3.

Example

What is the molar concentration of a solution of 80 g of sodium hydroxide in 4 dm^3 of solution?

M_r (NaOH) = 40

Moles of NaOH = $\frac{80}{40}$ = 2 mol

Concentration = $\frac{\text{Amount of solute (moles)}}{\text{Volume of solution } (dm^3)} = \frac{2}{4} = 0.5 \text{ mol } dm^{-3}$

Volumetric Analysis

A titration is a laboratory procedure where a measured volume of one solution of known concentration is added to a known volume of another reagent of unknown concentration until the reaction is complete. The operation is an example of volumetric (titrimetric) analysis. The **end point** is usually shown by the colour change of an indicator. The main pieces of apparatus used in a titration are a burette, a pipette with safety filler, a volumetric flask and several conical flasks.

A standard solution is a solution for which the concentration is known. To prepare a standard solution, a definite amount of solute must be dissolved in a definite volume of solvent to give a definite concentration of solution. The definite amount of material is measured by weighing, and the definite volume of solution prepared in a volumetric flask. A ***volumetric flask*** contains a definite volume when correctly filled to the calibration mark. The procedure used to prepare a standard solution is summarised as:

- Weigh out an accurate mass of a solid in a beaker and dissolve in a small volume of deionised water stirring with a glass rod.
- Transfer the solution with rinsing to the volumetric flask.
- Rinse the beaker and glass rod with deionised water and add to the volumetric flask.
- Make up to the mark by adding deionised water to the volumetric flask until the bottom of the meniscus is on the mark.
- Stopper the flask and invert to mix thoroughly.

The titration procedure is summarised as follows:

- Rinse the burette with the solution you are going to fill it with. Fill the burette with this solution, which is usually the solution of known concentration (the standard solution).
- Using a pipette filler, rinse a pipette with the solution you are going to pipette into the conical flask. Pipette 25.0 cm^3 of this solution into three different conical flasks.
- Add 2-3 drops of a suitable indicator to each conical flask.
- Add the solution from the burette until the indicator just changes colour.
- Repeat the titration to achieve 2/3 concordant results, adding the solution dropwise near the end point. Calculate the average titre from two accurate titre values.

Safety during a titration

- The use of a pipette filler.
- The use of gloves when appropriate.

Accuracy during a titration

- Rinse the apparatus with the appropriate solution.
- Add the solution from the burette slowly/ dropwise at the endpoint/colour change of the indicator.
- Swirl the flask/wash down the sides of the flask with distilled/deionised water.
- Read the burette at the bottom of the meniscus.

Reliability during a titration

- Repeat the titration 2/3 times.
- Concordant readings obtained (within 0.2 cm^3 of each other).

Note that an outlier is a value which is not concordant. The rough titration is often an outlier and should not be used in calculation of average titres.

In one volumetric analysis, three titrations should be carried out. The first titration should be rough and be an overshoot. This will allow you to determine approximately where the end-point lies. The second and third titrations should be accurate within 0.1cm^3 of each other, with dropwise addition as the end point is reached. A typical table used to record results is shown below.

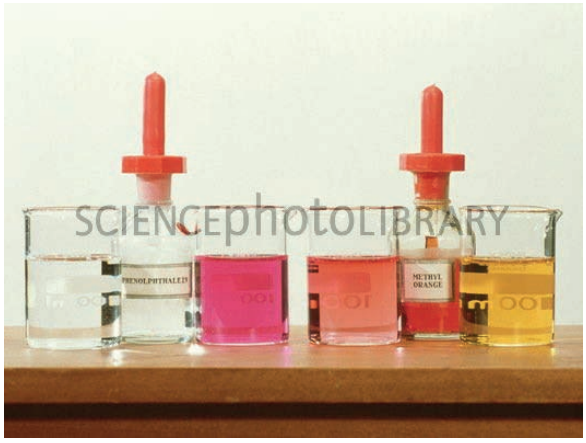
	Initial Burette reading (cm^3)	Final burette reading (cm^3)	Titre (cm^3)
Rough			
1st Accurate			
2nd Accurate			

Indicators are substances which change colour depending on the conditions. When an indicator changes colour during a titration, we say that the ***end point*** of the titration has been reached. An indicator is chosen so that the point at which it changes colour corresponds to the point, in the case of an acid-base titration, at which the solution in the flask has been completely neutralised. Some common indicators and their colour changes are shown in the table below:

Indicator	Colour in acid	Colour in alkali	Titrations suitable for
Methyl Orange	Red	Yellow	Strong acid-strong base Strong acid-weak base
Phenolphthalein	Colourless	Pink	Strong acid-strong base Weak acid-strong base

You must decide what type of titration is being carried out eg strong acid, strong base etc. and choose the appropriate indicator.

A weak acid/base is one which partially dissociates in solution. A strong acid/base is one which fully dissociates in solution.



Strong acids: Hydrochloric acid (HCl), Sulfuric acid (H_2SO_4), Nitric Acid (HNO_3)

Weak Acids: Ethanoic acid (CH_3COOH)

Strong Bases: Sodium hydroxide (NaOH), Potassium hydroxide (KOH)

Weak Bases: Ammonia (NH_3), Sodium carbonate (Na_2CO_3), Sodium hydrogencarbonate ($NaHCO_3$)

Example calculation

25.0 cm^3 of a solution of barium hydroxide, $Ba(OH)_2$, of unknown concentration is placed in a conical flask (using a pipette) and titrated with a solution of hydrochloric acid, HCl, which has a concentration of 0.0600 mol dm^{-3}. The volume of acid required is 20.40 cm^3. Calculate the concentration of the barium hydroxide solution.

$Ba(OH)_2(aq) + 2HCl(aq) \rightarrow BaCl_2(aq) + 2H_2O(l)$

Step 1: Calculate the amount of HCl delivered (the solution of known concentration) by using the expression

$$\text{moles (n)} = \frac{\text{volume x concentration}}{1000}$$

$$n = \frac{v \times c}{1000}$$

where c = 0.0600 moldm^{-3} and v =20.40 cm^3

$$n = \frac{20.40 \times 0.0600}{1000} = 1.22 \times 10^{-3} \text{ mol.}$$

Step 2: Calculate the amount of $Ba(OH)_2$ (the solution of unknown concentration) which reacts with this amount of HCl by substituting in the expression derived from the equation.

Ratio of HCl: $Ba(OH)_2$ = 2:1

Amount of $Ba(OH)_2$ = ½ × amount of HCl
= ½ × 1.22 x 10^{-3} mol = 6.10×10^{-4} mol

Step 3: Calculate the concentration of $Ba(OH)_2$ by substituting into the expression

$$c = \frac{n \times 1000}{v}$$

where n = 6.10×10^{-4} mol and v = 25.0 cm^3

$$c = \frac{n \times 1000}{v} = \frac{6.10 \times 10^{-4} \times 1000}{25.0}$$

$= 2.44 \times 10^{-2}$ mol dm^{-3}

There are two ways of doing a calculation. Either all the numbers are kept in the calculator or the individual steps are presented at each stage. There is also the problem of significant figures. But, whatever is done must be written correctly.

0.0600 is the concentration of the acid it has three significant figures; 20.40 is the titration value it has four significant figures because the last 0 could be a 5. That is why this zero is kept in the calculation but it could be omitted. The answer will need to be presented in terms of three significant figures because this is the justifiable degree of accuracy.

$0.0600 \times 20.40 = 1.224 \times 10^{-3} = 1.22 \times 10^{-3}$
(to three significant figures)

$= ½ \times 1.22 \times 10^{-3} = 6.10 \times 10^{-4}$
(accurately divided by 2)

If the numbers 0.06 and 20.4 had been left in the calculator and divided by 2 then
$0.06 \times 0.0204 \times ½ = 6.12 \times 10^{-4}$
=

Back to the calculation:

$$c = \frac{n \times 1000}{v} = \frac{6.10 \times 10^{-4} \times 1000}{25.0} = 2.44 \times 10^{-2} \text{ mol dm}^{-3}$$

The answer 6.12×10^{-4} would have given 2.45×10^{-2}.

Calculation of uncertainty

Uncertainty is an estimate attached to a measurement which characterises the range of values within which the true value is thought to lie. This is normally expressed as a range of values such as 44.0 ± 0.4.

When glassware is manufactured there will always be a maximum uncertainty/error. This is usually marked on the glassware. Error is the difference between an individual measurement and the true value of the quantity being measured.

The significance of error in a measurement depends upon how large a quantity is being measured. It is useful to quantify this uncertainty as a percentage uncertainty

$$\text{Percentage uncertainty} = \frac{\text{uncertainty}}{\text{measurement}} \times 100$$

When a burette is used to calculate an accurate titration there are two uncertainties.

Example

A burette has an error ± 0.05 cm^3. In a titration the initial burette reading was 0.05 cm^3 and the final burette reading was 24.55 cm^3. What is the percentage uncertainty in the titre value?

Answer

Initial burette reading = 0.05 cm^3
Final burette reading = 24.55 cm^3

The overall uncertainty in any volume measured in a burette always comes from the two measurements, so the overall uncertainty = 2 × 0.05 cm^3 = 0.01 cm^3

Titre value = 24.50 cm^3

$$\frac{\text{Percentage}}{\text{uncertainty}} = \frac{2 \times 0.05}{24.50} \times 100 = 0.4\%$$

Revision Questions

1 Washing soda is hydrated sodium carbonate and can be represented by the formula $Na_2CO_3xH_2O$. The value of x can be found by titrating a solution of washing soda against standard hydrochloric acid solution.

a) (i) What is meant by the term **standard** solution?

..... [1]

(ii) Write the equation for the reaction between sodium carbonate, Na_2CO_3, and excess hydrochloric acid.

..... [1]

b) In one experiment a 2.80 g sample of washing soda was made up to 250 cm^3 of solution in a volumetric flask. 25 cm^3 of this solution required 22.4 cm^3 of 0.1 mol dm^{-3} hydrochloric acid for neutralisation. Find the value of *x* using the following headings.

Moles of hydrochloric acid used

.....

Moles of sodium carbonate in 25 cm^3

.....

Moles of sodium carbonate in the sample

.....

Mass of water in the sample

.....

Moles of water in the sample

.....

Value of *x*

..... [5]

2 Analysis of a vinegar solution was carried out using the following procedure:

Transfer 25.0 cm^3 of undiluted vinegar into a 250 cm^3 volumetric flask and make the solution up to the mark using de-ionised water. Transfer 25.0 cm^3 portions of the diluted vinegar into three separate conical flasks and add a few drops of indicator of each flask. Titrate each solution with 0.1 mol dm^{-3} sodium hydroxide until and end point is reached.

A student obtained the following results:

	Initial burette reading (cm^3)	Final burette reading (cm^3)	Titre (cm^3)
Rough	0.0	21.7	21.7
1st accurate	21.7	43.1	
2nd accurate	0.0	21.3	

a) (i) Name a suitable indicator for this titration.

..... [1]

(ii) State the colour change which would be obtained at the end point.

from ..... to ..... [2]

b) (i) Write the equation for the reaction between vinegar (ethanoic acid) and sodium hydroxide.

..... [2]

(ii) Complete the results table and calculate the average titre.

..... [2]

(iii) Use the average titre to calculate the number of moles of sodium hydroxide used in the titration..

..... [1]

(iv) Calculate the concentration of ethanoic acid in the diluted vinegar.

..... mol dm^{-3} [1]

(v) Calculate the concentration of ethanoic acid in the undiluted vinegar.

..... mol dm^{-3} [1]

c) Describe, giving practical details, how you would prepare the solution of diluted vinegar and then transfer 25.0 cm^3 to a conical flask.

.....

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.....

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..... [4]

3 2.65 g of anhydrous sodium carbonate. Na_2CO_3, were dissolved in water and the solution made up to 250 cm^3 in a volumetric flask. The concentration of the solution was

A 0.025 mol dm^{-3}
B 0.050 mol dm^{-3}
C 0.100 mol dm^{-3}
D 0.200 mol dm^{-3}

4 For which one of the following titrations would phenolphthalein be a suitable indicator?

A ethanoic acid and sodium carbonate
B ethanoic acid and sodium hydroxide
C hydrochloric acid and aqueous ammonia
D hydrochloric acid and sodium carbonate