Chapter 17: Titrations (Acid-base)

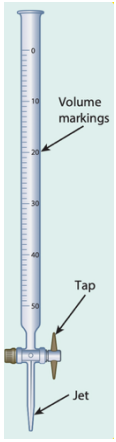
Volumetric analysis is a quantitative analytical method that involves measuring volumes (hence the name) **in order to determine the concentration of a solution**.

To do this, **experiments called titrations** are carried out.

Titration equipment

Burette

A burette is designed to deliver measurable quantities of a liquid with high accuracy.



It must first be rinsed:

- **1) with deionised water (to remove any impurities) & then…**
- **2) with the solution it will contain during the experiment (this will remove any drops of deionised water that would dilute the solution in the burette).**

Normal tap water contains different types of ions, which could interfere with the reaction and lead to inaccurate results. Therefore deionised water – water with all ions removed – is used during titrations.

Rinsing

- Clamp the burette vertically to a retort stand and place a beaker marked 'waste' underneath the tap. Ensure the tap is closed. Using a funnel, pour some deionised water into the burette. *(You could also just hold the burette in one hand & use a wash bottle in the other to add some deionised water into the burette. Ensure the tap is closed!)*
- Unclamp the burette, tilt it almost horizontally, and rotate to ensure the deionised water touches all inside surfaces.
- Re-clamp the burette and open the tap to allow the deionised water to run through the tap and jet (the part below the tap), ensuring they are also rinsed.
- Repeat this with the solution that the burette is about to contain.

Filling

- Using a retort stand and clamp, **clamp the burette vertically** (to ensure an accurate reading).
- Check that the **tap is closed**.
- Using a **funnel,** fill the burette with the required solution – e.g. hydrochloric acid.
- Add the solution until **it's above the 0cm^3 graduation mark**.
- **Remove the funnel** (to avoid drips into the burette, which could affect accuracy).
- Place an empty beaker marked 'waste' beneath the tap and **open the tap to allow the solution to flow through and fill the jet.**
- Ensure there are **no air bubbles**, especially in the jet (to remove air bubbles, quickly open and close the tap).
- Finally, adjust the level of the solution carefully until the **bottom of the meniscus is on the graduation mark at eye level.**

Link it: What type of error could occur if you didn't look at the meniscus at eye level?

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Pipette

- A pipette is a piece of equipment designed to measure and transfer an exact volume of liquid.
- A **pipette filler** is used to fill a pipette.
- The pipette is **rinsed 1) with deionised water 2) with the solution that it will contain.**
- When filling the pipette, always keep the tip of the pipette beneath the surface of the liquid you are filling it with, but ensure it doesn't touch the bottom of the beaker.
- **Fill the pipette until the liquid level is above the graduation mark, then slowly release liquid until the bottom of the meniscus is on the graduation mark at eye level**.
- To transfer the liquid, **touch the tip of the pipette against the inside of the conical flask and allow the liquid to drain out slowly**.
- There will be a drop of liquid left in the pipette at the end of this drainage time – don't worry, this drop was accounted for when the pipette was designed. **Do not blow or shake out this drop as that would give you an inaccurate volume of liquid.**

Volumetric flask

- The volumetric flask is used to make up solutions of a specific volume – e.g. 250cm^3, 500cm^3 or 1000cm^3.
- When filling the volumetric flask, a funnel is used.
- However, as the solution nears the graduation mark, the funnel is removed.
- **The last few drops of deionised water are added using a dropper until the bottom of the meniscus rests on the graduation mark at eye level.**
- The volumetric flask is **stoppered and repeatedly inverted** (approx. ten to twenty times) to **ensure the solution is homogeneous** (uniformly mixed).

Conical flask

- Ideally, a clean, dry conical flask should be used.
- If dirty, a conical flask can be cleaned by rinsing it with deionised water only.
- The conical flask **should <u>not</u> be rinsed with the solution** it will contain – e.g. the base, since that would mean we wouldn't know the exact volume of base present in the conical flask for the experiment.
- Designed so that it may be **swirled** (without splashing)

White tile

- Used to see the colour change of the indicator more clearly.
- Placed below the conical flask.

Wash bottle

- A wash bottle filled with deionised water is used to wash down the inner sides of the conical flask.
- This is done to ensure that every drop of liquid that leaves the burette (e.g. an acid) reacts with the liquid in the conical flask (e.g. a base).
- If drops of acid were left clinging to the side of the conical flask but did not react with the base, it would give an inaccurate result for the volume of acid used.

Acid-base titrations

- Acid-base titrations are carried out between an **acid and a base.**
- When an acid and base react together, they form a salt and water (this is a **neutralisation reaction**).
- The equivalence point for an acid-base reaction is when **$H^+ = OH^-$.**
- An indicator is added to **indicate the equivalence point with a colour change**.
- The **endpoint refers to the point at which the indicator changes colour**. If the correct indicator is used, the equivalence and endpoint should be the same.
- However, if the incorrect indicator is used, this will not be the case (for more on choosing the correct indicator, see Chapter 21).
- **Indicators are themselves weak acids or bases. Therefore only one to two drops should be used** so that they do not interfere with the reaction and give an inaccurate result.

Carrying out an acid-base titration

1. Prepare all equipment (rinsing appropriately, etc.).

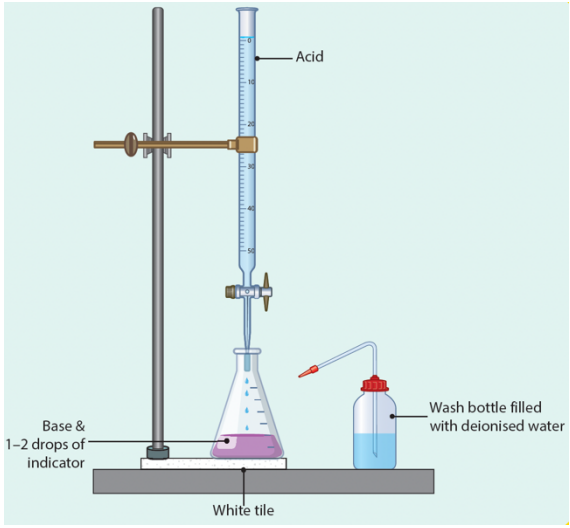
2. Fill the burette with the acid.

3. Use a pipette (and pipette filler) to transfer a known volume of the base to a conical flask.

4. Add one to two drops of indicator to the conical flask, and swirl to mix.

5. Place the conical flask on a white tile on the base of the retort stand.

6. Begin titrating by opening the tap of the burette and allowing the acid to flow into the conical flask. Keep one hand on the tap of the burette while using the other to swirl the conical flask continuously to ensure even mixing. From time to time, wash down the sides of the conical flask with a wash bottle of deionised water.



7. When the indicator changes colour, close the tap of the burette immediately to stop the flow of liquid. Note the volume of the acid required to cause this colour change. This is your first (rough) titration.

8. Now you know roughly how much of the acid is needed, you will carry out two further accurate titrations. This time, slow down the flow of the acid as you approach the end point (the volume required for the first titration). You should be adding the acid drop by drop as you approach the end point, and one drop of the acid should be enough to cause the colour change.

Ensuring accuracy during a titration

- A white tile is placed beneath the conical flask in order to see the colour change more clearly.
- A wash bottle filled with deionised water is used to rinse down the sides of the conical flask.
- All readings should be taken from the bottom of the meniscus at eye level.
- Continuously swirl the conical flask to keep the solution evenly mixed.
- Add the solution from the burette drop by drop when approaching the end point.

Now try Qs 1-11 in chapter 17 of the workbook (page 31).

Experimental Investigation 11: Carry out a strong acid-weak base titration

- Burette = hydrochloric acid (strong acid) – to be standardised, i.e. concentration to be determined
- Conical flask = anhydrous sodium carbonate solution (a standard solution of a weak base)
- Indicator = methyl orange
- Colour change = yellow/ orange → pink/red

See page 49 of the lab book for more.

Titration Calculations

The data obtained from carrying out titrations in the lab is used to calculate the concentration of the unknown solution. These are the four steps that you should always follow when carrying out a titration calculation.

Step 1: Begin by drawing a rough sketch of the titration at the top of your page. These questions can be quite wordy and it's crucial to be able to place the correct information with the correct solution.

Step 2: Start with the solution whose concentration and volume are given. Convert concentration to moles/L. Find out how many moles of this solution were used.

Step 3: Use the ratio (always given in the balanced equation) to find out how many moles of the other solution were used (the solution of unknown concentration).

Step 4: Use the answer from step 3, as well as the volume (cm^3) of the solution of unknown concentration used (always given in the question) to find out the molarity of the solution.

Example 1

Meadhbh and Nina carried out an experiment to determine the concentration of a solution of ammonia. 25cm^3 portions of the ammonia solution were titrated against a (previously standardised) 0.01M solution of hydrochloric acid in a burette. On average, 20.6cm^3 of hydrochloric acid was needed. The balanced equation for the titration was: $HCl + NH_3 \rightarrow NH_4Cl$. Calculate (a) the molarity of the ammonia solution; (b) the concentration in g/L of ammonia, NH_3, in the solution.

Step 1: Draw your rough sketch and fill in all known information.

Step 2: Start with the molarity of the solution that you know the most about. Find the number of moles used.

Step 3: Use the ratio to figure out how many moles of the other solution must have reacted.

Step 4: Use the number of moles used, plus the volume used, to calculate the molarity of the solution of unknown concentration.

Example 2

Louis and John carried out an experiment to determine the concentration of a solution of hydrochloric acid, HCl. Using a pipette, they transferred 25cm^3 portions of a 0.1M Na_2CO_3 solution to a conical flask and titrated it against a solution of HCl, according to the following balanced equation: $2HCl + Na_2CO_3 \rightarrow 2NaCl + H_2O + CO_2$. The three titres obtained were 19.8cm^3, 19.5cm^3 and 19.6cm^3. Calculate the molarity of the HCl solution.

Step 1: Rough sketch

Step 2: Find the number of moles used (of solution you know the most about)

Step 3: Ratio

Step 4: Calculate unknown concentration & answer questions asked

Now try the test yourself questions on page 251 of the textbook.

Experimental Investigation 12: Carry out a strong acid-strong base titration

- Burette = hydrochloric acid (strong acid) – a standardised solution
- Conical flask = sodium hydroxide solution (a strong base, but not a primary standard – since it absorbs water from the air - therefore this is not a standard solution, and must be standardised)
- Indicator = methyl orange
- Colour change = yellow/ orange → pink/red

See page 53 of the lab book for more.

Example 3

Alex and Leah carried out an experiment to determine the concentration of a solution of sodium hydroxide, NaOH. Using a pipette, they transferred 25.0cm^3 portions of this solution to a conical flask, and titrated it against a solution of (previously standardised) 0.05M HCl, according to the following balanced equation: HCl + NaOH → NaCl + H2O. The average of the two accurate titres was 22.5cm^3. Calculate (a) the molarity of the NaOH solution; (b) the concentration in g/L of NaOH.

Now try the test yourself questions on page 254 of the textbook.

Experimental Investigation 13: Investigate, using primary data, how to find the concentration of ethanoic acid in vinegar

- Burette = diluted vinegar solution – to be standardised, i.e. concentration to be determined
- Conical flask = sodium hydroxide solution – standardised
- Indicator = phenolphthalein
- Colour change = pink → colourless

See page 57 of the lab book for more.

Example 4

Sorcha and Naoise wanted to determine the concentration of ethanoic acid, CH_3COOH, in a sample of vinegar. First, they diluted the vinegar (from $25cm^3$ to $250cm^3$) and then they titrated this solution against $25cm^3$ portions of a previously standardised 0.1M solution of sodium hydroxide, NaOH. One rough and two accurate titrations were carried out. The average of the accurate titres was $18.75cm^3$.

The equation for the reaction is $CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$.

(a) Calculate the concentration of the diluted solution of ethanoic acid in (i) moles per litre; (ii) grams per litre.

(b) (i) State the concentration of ethanoic acid in the original vinegar sample in grams per litre. (ii) Express this concentration in terms of % w/v.

Example 5

To determine the concentration of ethanoic acid in a sample of vinegar, Theo and Yujing diluted 25cm^3 of vinegar to 250cm^3. They then titrated the diluted vinegar against a previously standardised solution that contained 1.20g of sodium hydroxide in 500cm^3 of solution. On average, 18.75cm^3 of the diluted vinegar was required to neutralise 25cm^3 of this sodium hydroxide solution. The equation for the titration reaction is:

$$CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$$

Calculate (a) the number of moles of sodium hydroxide in each 25cm^3 portion; (b) the number of moles of ethanoic acid per cm^3 of diluted vinegar; (c) the concentration of ethanoic acid in the original vinegar in moles per litre.

Now try the test yourself questions on page 260 of the textbook, and questions 14 – 25 in the workbook.

Test Yourself – Past papers

Now test yourself on the following past paper questions (from the old course, but applicable to yours!). Begin with some OL questions to warm up, then test yourself on the HL questions.

If you have forgotten your papers in school, the website https://betterbetterexaminations.ie/ has them all (and is easier to navigate than the SEC version!)

Past Paper Questions OL	✓
2025 Q2	
2024 Q2	
2023 Q2	
2022 Q2	
2021 Q2	
2020 Q2	

& all of the other Q2s from the ordinary level papers….

Past Paper Questions HL	✓
2024 Q1	
2016 Q1	
2012 Q1	
2008 Q1	