# PART 3 – QUANTITATIVE ANALYSIS

## INQUIRY QUESTION: HOW ARE SOLUTIONS OF ACIDS AND BASES ANALYSED?

**Students:**

- conduct practical investigations to analyse the concentration of an unknown acid or base by titration
- investigate titration curves and conductivity graphs to analyse data to indicate characteristic reaction profiles, for example:

  – strong acid/strong base
  – strong acid/weak base
  – weak acid/strong base (ACSCH080, ACSCH102)

- model neutralisation of strong and weak acids and bases using a variety of media

## VOLUMETRIC ANALYSIS

Volumetric analysis is a quantitative technique that involves reactions in solution. It requires the determination of the volume of a standardised solution that reacts with the substance (or 'unknown') being analysed. A technique called **titration** which is used in the analysis. In acid–base analysis the reaction involves a neutralisation reaction.

Volumetric analysis requires that one solution must have an accurately known concentration. This solution is called the **standard solution**. The standard solution may be either a **primary standard** or a secondary standard. The preparation of primary standards is discussed below.

## STANDARD SOLUTIONS

In preparing a standard solution of definite concentration the chemical to be used (the 'Primary standard') must satisfy the following criteria:

(a) It must be a water-soluble solid.
(b) It must have high purity.
(c) It must not lose or gain water or react with oxygen or carbon dioxide in the air.
(d) It must have a high molecular mass (reducing the effect of measurement errors).

The most common substances to be used as Primary standards included sodium carbonate ($Na_2CO_3$) and oxalic acid ($H_2C_2O_4.2H_2O$).

Sodium Hydroxide is **NEVER** used as a primary standard. This is because it undergoes two reactions when exposed to air, both of which affect the mass of the sodium hydroxide. Firstly, sodium hydroxide is **deliquescent**, which means it absorbs water from the air. Secondly, sodium hydroxide reacts with carbon dioxide in the air.

## PREPARING A STANDARD SOLUTION.

1. Rinse volumetric flask with distilled H2O.
2. Weigh PRIMARY STANDARD in a small beaker.
3. Dissolve PRIMARY STANDARD in small volume of distilled $H_2O$.
4. Pour solution into volumetric flask.
   i. Use filter funnel.
   ii. Rinse beaker and funnel several times with distilled $H_2O$ and add rinsings to flask.
   iii. Invert to mix the solution, until all solid is dissolved.
5. Add distilled $H_2O$ until bottom of meniscus reaches calibration line. Place lid on and invert to mix solution.

## EXPLAIN THE PROCESS OF TITRATING

In this process the concentration of a solution is accurately determined by measuring exactly how much of the standard solution is required to neutralise a known volume of it. This is because neutralisation occurs when the number of moles of acid is stoichiometrically equal to the number of moles of base. This is called the **EQUIVALENCE POINT**.

The problem is that acid solutions are colourless, base solutions are colourless and so is the product solution. How can we tell when the reaction has reached the equivalence point? The answer is to use an appropriate indicator. The indicator should be chosen so that it changes colour at the pH of the equivalence point. The point when the indicator changes colour is called the **END POINT**. The goal is the choose an indicator where the end point will be same as the equivalence point.

| Indicator | Colour Change at different pH levels |
|---|---|
| Methyl orange | Red below 3.1 – Yellow above 4.4 |
| Bromothymol Blue | Yellow below 6.2 – Green around 7 – Blue above 7.6 |
| Phenolphthalein | Colourless below 8.3 – Red above 10.0 |

**Note:**

pH meters improve accuracy, but appropriate indicators are a very useful alternative which do not need to be calibrated or checked.

## METHOD FOR TITRATION

1. Pour standard solution into a clean burette using a clean funnel. Ensure solution goes above 0 mL line.

2. Run excess solution out of the burette and discard (ensure new level is 0 mL).

3. Rinse a clean conical flask with distilled water.

4. Add the unknown solution to the flask by careful measurement using a pipette (previously rinsed with the unknown solution) and place it under the burette.

5. Add 2-5 drops of an appropriate indicator.

6. Slowly add the standard solution in the burette to the flask, swirling continuously until the indicator just changes colour (end point, and hopefully the equivalence point).

7. Read volume of burette.

8. Repeat three times. Calculate average. Calculate required concentration of unknown.

**QUESTION 41** (HSC 2006:9)
Which statement best describes the equivalence point in a titration between a strong acid and a strong base?

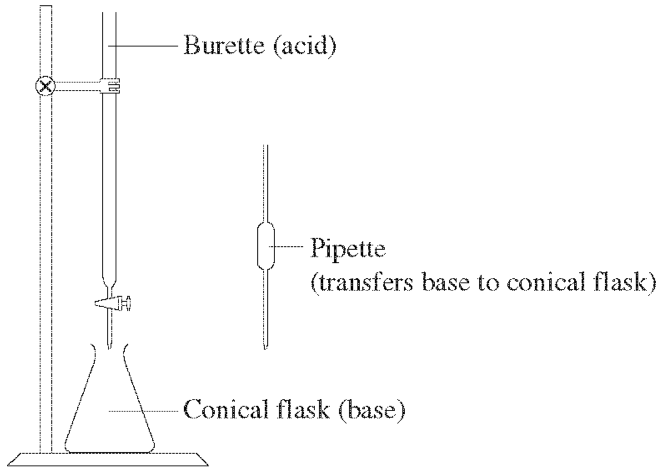
A The point at which the first sign of a colour change occurs.

B The point at which equal moles of acid and base have been added together.

C The point at which equal moles of $H^+$ ions and $OH^-$ ions have been added together.

D The point at which the rate of the forward reaction equals the rate of the reverse reaction.

**QUESTION 42** (HSC 2002:8)
In a titration, an acid of known concentration is placed in a burette and reacted with a base that has been pipetted into a conical flask.




What should each piece of glassware be rinsed with immediately before the titration?

| | Burette | Pipette | Conical Flask |
|---|---|---|---|
| A | Acid | Base | Water |
| B | Water | Water | Water |
| C | Acid | Base | Base |
| D | Water | Water | Base |

**QUESTION 43** (HSC 2003:14)
In a titration of a strong base with a strong acid, the following procedure was used:

1. A burette was rinsed with water and then filled with the standard acid.
2. A pipette was rinsed with some base solution.
3. A conical flask was rinsed with some base solution.
4. A pipette was used to transfer a measured volume of base solution into the conical flask.
5. Indicator was added to the base sample and it was titrated to the endpoint with the acid.

Which statement is correct?

A The calculated base concentration will be correct.

B The calculated base concentration will be too low.

C The calculated base concentration will be too high.

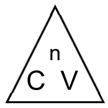
D No definite conclusion can be reached about the base concentration.

## 4 STEPS TO SUCCESS WITH TITRATION CALCULATIONS

1. Find the moles used of the known solution (moles = conc. x volume).
2. Find the ratio of reactants (usually using a balanced chemical equation).
3. Use this ratio to find the moles of the unknown.
4. Calculate the concentration of the unknown solution (conc. = mole / volume).

**The formula to use:**

(C = concentration in mol$^{-1}$, n = number of moles, V = volume in L)




**Remember:** The formula to use if dilutions are involved: $C_1 V_1 = C_2 V_2$

Sometimes the formula $C_A V_A / a = C_B V_B / b$ is taught. However, it is preferable to use the four steps outlined above, especially for back titrations.

**QUESTION 44** (HSC 2008:28)
A standard solution was prepared by dissolving 1.314 g of sodium carbonate in water.
The solution was made up to a final volume of 250.0 mL.

(a) Calculate the concentration of the sodium carbonate solution. (2 marks)