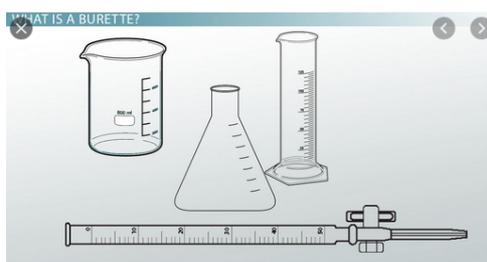


# Accuracy and Error in Titration

## Accuracy (temperature-dependent accuracy)

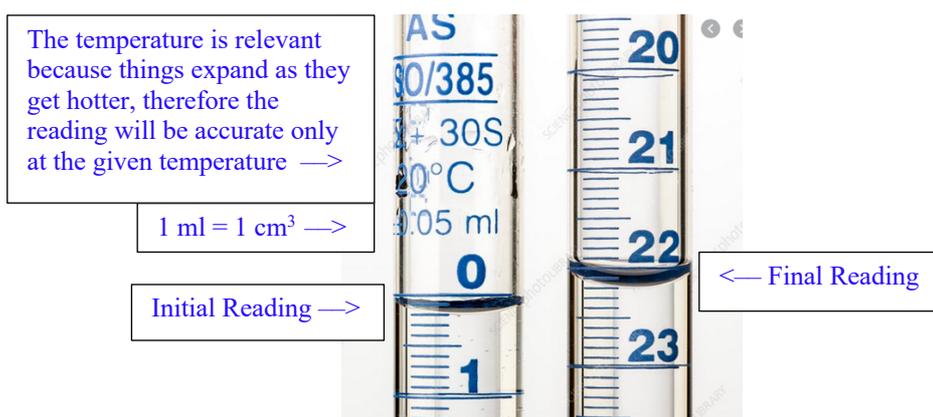
A manufacturer can make a product (e.g. a ruler or a burette) out of say plastic, but the length of the ruler (or the burette) will alter with the temperature of the room. The manufacturer will therefore specify that the measuring instrument will give a true reading at a given room temperature of say 20°C. However, at a higher temperature (of say 25°C) the ruler will have expanded and the reading it gives will no longer be *accurate*. It will no longer give a reading that reflects the true value of the reading.



Shown alongside are (from left to right)  
- a beaker  
- a conical flask  
- a measuring cylinder, and  
- a burette (lying on its side).

## Error (theoretical accuracy)

For a relatively cheap instrument, a manufacturer cannot guarantee that any given reading will be 100% accurate. The manufacturer will therefore specify the (usually) maximum error that will be involved in the measuring instrument. For example, in the burette below, the manufacturer has specified that at 20°C the error of any reading will be (at most)  $\pm 0.5 \text{ ml} = \pm 0.05 \text{ cm}^3$  (1 ml = 1 cm<sup>3</sup>). The initial reading at 20°C in this case is 0.2 cm<sup>3</sup>, but given the manufacturer's error involved, the true value of the reading could have been  $(0.2 \pm 0.05) \text{ cm}^3$  i.e. anywhere between **0.15 cm<sup>3</sup>** and **0.25 cm<sup>3</sup>**, and the true value for the Final Reading could have been anywhere between **22.15 cm<sup>3</sup>** and **22.25 cm<sup>3</sup>**. (In *each* reading at 20°C the maximum error will be 0.05 cm<sup>3</sup>.)<sup>1</sup> Please remember that **every time that a reading is taken, there will be a possible error – therefore in every titration there will be TWO POSSIBLE ERRORS (one when the Initial Reading is taken, and one when the Final Reading is taken).**



*I hope that that is clear.* Since every titration involves two readings (an Initial Reading and a Final Reading), then every titration will have **two** possible errors. In the case of the burette above, each possible error is 0.05 cm<sup>3</sup>, therefore the **total error** in each titration will be  $2 \times 0.05 \text{ cm}^3 = \mathbf{0.1 \text{ cm}^3}$ .

<sup>1</sup> In English, if a whole sentence has gone into “parenthesis” (i.e. it is enclosed in brackets/commas/dashes), then the final full-stop must go *inside* the parenthesis.

## Acceptable Error

Let us say that (for whatever reason) precisely 20 cm<sup>3</sup> of a liquid has to be measured by titration, and that a maximum possible percentage error of 0.5 % is not acceptable, and that the maximum percentage error has to be 0.25 %. The maximum possible error involved would then have to be halved. What could one do if a more accurate burette were not available?

Each titration will always involve two possible errors (one from the Initial Reading and one from the Final Reading), therefore if the volume of the liquid used from the burette were 40 cm<sup>3</sup> instead of 20 cm<sup>3</sup>, then the maximum possible percentage error involved would be halved. Let us look at the Maths involved. The maximum possible overall error would be 2 x 0.05 cm<sup>3</sup> = 0.1 cm<sup>3</sup> (0.05 cm<sup>3</sup> at each reading, therefore the maximum possible **percentage** error would be given by

$$\text{For } 20 \text{ cm}^3: \text{ Maximum possible percentage error} = (2 \times 0.05 \text{ cm}^3 = 0.1 \text{ cm}^3) \frac{[0.1 \text{ cm}^3 \times 100]}{20 \text{ cm}^3} \% = 0.5 \% .$$

but the maximum possible percentage error for 40 cm<sup>3</sup> would be given by

$$\text{For } 40 \text{ cm}^3: \text{ Maximum possible percentage error} = (2 \times 0.05 \text{ cm}^3 = 0.1 \text{ cm}^3) \frac{[0.1 \text{ cm}^3 \times 100]}{40 \text{ cm}^3} \% = 0.25 \% .$$

Therefore, the overall percentage error in any titration can be **halved** by halving the **concentration** of the liquid inside the burette, because **double** the amount of liquid would then be needed to react with whatever is in the conical flask.

Equally, the percentage error could also be halved by **doubling** the the concentration of the liquid **inside the conical flask** because then, once again, 40 cm<sup>3</sup> of the liquid in the burette would be needed to react with the liquid in the conical flask.

In English that may sound a bit complicated, but if you are more comfortable with Maths than with English, then I will do the Maths involved – and let me do the Maths for a straightforward *Acid:Base* reaction – let us say a reaction of Sulphuric Acid against Sodium Hydroxide.

We will need to start with a Reaction Equation – **and every titration MUST always start with a Reaction Equation because it is the proportions (or ratios) in a reaction that control the number of molecules (or the number of moles of molecules) that react with each other.**

## A Sample Acid-Alkali Reaction: Titration 1

Let us say that 20cm<sup>3</sup> of Sulphuric Acid at a concentration of 0.100 mol dm<sup>-3</sup> in the burette are needed to react with 40 cm<sup>3</sup> of Sodium Hydroxide at a concentration of 0.100 mol dm<sup>-3</sup> of the Alkali<sup>2</sup> in the conical flask – and please remember that

$$\text{Concentration, } C = \frac{\text{Number of Moles (N) of the substance}}{\text{Volume (V) of the containing liquid}} \quad \frac{\text{in moles}}{\text{in dm}^3}$$

$$\text{or} \quad C = \frac{N}{V} \quad (\text{in mol dm}^{-3}) \quad \text{and} \quad V = \frac{N}{C} \quad (\text{in dm}^3) \quad \text{and} \quad N = C \times V \quad (\text{in mol})$$

where **cm<sup>3</sup>** must always be converted into **dm<sup>3</sup>** (because that is the appropriate unit for Titrations).

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<sup>2</sup> ... or the **Base** in the conical flask – an alkali being a *soluble* base.

**First a word about Significant Figures: A chain is as strong only as its weakest link!**

- In Chemistry, an expression such as  $0.100 \text{ mol dm}^{-3}$  Sulphuric Acid gives the value to **THREE** significant figures. In Maths one can re-write “0.100” as “0.1” – but this **cannot** be done in Chemistry because that would change the number of significant figures given from **three** to just **one**.
- The expression  $20 \text{ cm}^3$  gives **two** significant figures, therefore when it is converted into  $\text{dm}^3$ , the expression must be written as  $0.020 \text{ dm}^3$  and not  $0.0200 \text{ dm}^3$  or  $0.02 \text{ dm}^3$  because the latter would give different numbers of significant figures.

When doing the mathematics of a calculation, one should use whatever number of decimal places that there are in the calculator that is being used – but the **final** answer must always be given to the **LOWEST** number of significant figures that was given in the original context (the exam question/the experiment/whatever).

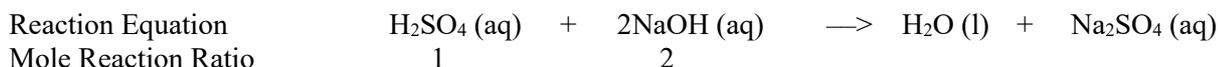
OK, let us get back to the Titration.

In this Titration, of **Sulphuric Acid** against **Sodium Hydroxide**

The Number of moles of NaOH,  $N = C \times V = 0.100 \text{ mol dm}^{-3} \times 40 \text{ cm}^3 (= \frac{40}{1,000} \text{ dm}^3 \text{ or } 0.040 \text{ dm}^3)$ .

**N of NaOH** therefore =  $C \times V = 0.100 \text{ mol dm}^{-3} \times 0.040 \text{ dm}^3 = 0.0040 \text{ mol}$  (because the  $\text{dm}^{-3}$  and  $\text{dm}^3$  cancel out).

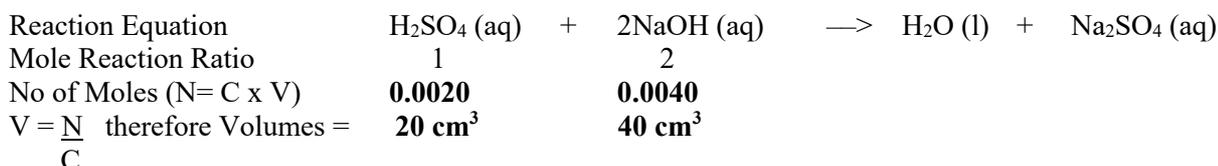
In Chemistry you **must** write this value as **0.0040 mol** to indicate that **two** significant figures are needed. *A chain is as strong only as its weakest link, and although C was given to 3 Sig Figs, V was given to 2 Sig Figs.*



therefore the number of moles of  $\text{H}_2\text{SO}_4 (\text{aq})$  required to react with **0.0040** moles of  $\text{NaOH} (\text{aq}) = 0.0020$ , and, since  $V = \frac{N}{C} = \frac{0.0020 \text{ mol}}{0.100 \text{ mol dm}^{-3}}$  this gives **0.020 dm<sup>3</sup>** to 2 Sig.Figs. (and this =  $20 \text{ cm}^3$  of  $\text{H}_2\text{SO}_4$ ).

We have therefore established the reacting Volumes of the substances concerned for the Concentrations that are involved viz. **20 cm<sup>3</sup>** of  $0.100 \text{ mol dm}^{-3}$  the Sulphuric will react exactly with **40 cm<sup>3</sup>** of the Sodium Hydroxide.

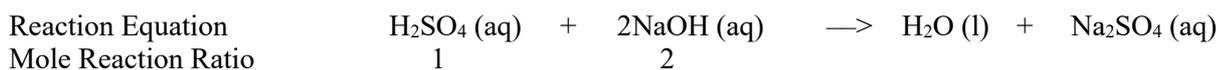
In an exam, to save time and space you can write the calculation as follows.



**The Sample Acid-Alkali Reaction: Titration 2**

However, if we were to **HALVE** the concentration of Sulphuric Acid to **0.0500 mol dm<sup>-3</sup>**, then the number of moles of Sulphuric Acid required to exactly react with **40 cm<sup>3</sup> of Sodium Hydroxide** (which is still at a concentration of  $0.100 \text{ mol dm}^{-3}$  of the Alkali/the Base) would still be given by

**N of NaOH** therefore =  $C \times V = 0.100 \text{ mol dm}^{-3} \times 0.040 \text{ dm}^3 = 0.0040 \text{ mol}$  (because the  $\text{dm}^{-3}$  and  $\text{dm}^3$  cancel out).



but the number of moles of  $\text{H}_2\text{SO}_4 (\text{aq})$  that are required to react with **0.0040** moles of  $\text{NaOH} (\text{aq})$   
 = **0.0020** moles of  $\text{H}_2\text{SO}_4 (\text{aq})$ , (because the Mole Reaction Ratio = 1:2)

and the Volume of  $\text{H}_2\text{SO}_4 (\text{aq})$  that is required will thus be

$$V = \frac{N}{C} = \frac{0.0020 \text{ mol}}{0.0500 \text{ mol dm}^{-3}} = 0.040 \text{ dm}^3 \quad (\text{and this} = 40 \text{ cm}^3 \text{ of } \text{H}_2\text{SO}_4).$$

By **halving** the Concentration of the Sulphuric Acid, the Volume that is required to exactly neutralise the  $\text{NaOH}$  has **doubled** to  $40 \text{ cm}^3$ , and this will halve the possible % error involved from 0.5% to 0.25%.

Equally, if the concentration of the  $\text{NaOH}$  in the conical flask were **doubled** (but the Concentration of the Sulphuric Acid were left unaltered), then **twice** the volume of the Sulphuric Acid would be needed (i.e.  $40 \text{ cm}^3$  instead of  $20 \text{ cm}^3$ , to exactly neutralise the  $20 \text{ cm}^3$  of  $\text{NaOH}$  in the conical flask) – and the possible % error would then be halved.

Let me sum it up simply by saying that in Titration, the possible overall percentage error can be halved

- by halving the concentration of the liquid in the **burette**, or
- by doubling the concentration of the liquid in the **conical flask**

**because each of these solutions would double the volume of the liquid needed from the burette.**

In my view neither the English nor the Maths in this exercise is *easy*, but it is important grasp BOTH. It is important to get to grips with the Maths because all Titrations in the ‘A’ Level exams are based on similar calculations; and, it is important to get to grips with the English because all languages are merely a means of communicating the things in one person’s head into another person’s head and, order to do so, there must be widely-acknowledged rules as to what a certain way of saying something will be understood by the hearer or by the reader of what has been said – and those rules **must be obeyed**, otherwise there is no guarantee that the hearer/reader will understand the speaker/writer correctly. **The rules of Grammar in every language ARE important and MUST be obeyed** just as the rules of Mathematics are important and must be obeyed.

The more that one reads and hears correctly-spoken language, then the better that human beings become at understanding each other.

It is also vital that you understand that, in Titrations, it is the **proportions** in a Reaction Equation that control the amounts of the reacting liquids. You must therefore start every Titration with a Reaction Equation – otherwise you are likely to get the *wrong* answer. I therefore show you a couple of classic Titrations overleaf.

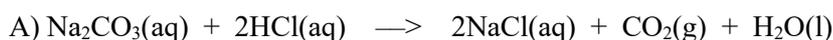


- OK, those are the basic GCSE calculations involving C, N, and V , so let us now do some slightly more difficult stuff.
- One of the most fundamental reactions in Chemistry is the neutralisation of an Acid by a Base – and (hardly surprisingly) this is called an Acid-Base Reaction. You may know it as  

$$\text{Acid} + \text{Alkali} \longrightarrow \text{Salt} + \text{Water}$$
and the neutralisation of an acid by an Alkali is one of the most simple of all titrations.
- However, there are many other titrations that can be executed (e.g. the titration of a Thiosulphate solution against a solution of Iodine, and so on), but in all the titrations that you will do, what you will be trying to do is **to find out the EXACT point at which the full extent of the reaction (that is involved) has just taken place (this being called the “equivalence point”)**. NB The “**end-point**” of the reaction is where the indicator that has been used changes colour. The ‘end-point’ indicates that the ‘equivalence point’ has been reached.
- As I said earlier, one of the IMMUTABLE rules of Chemistry is that **you cannot do ANY calculation involving a reaction without first knowing what the REACTION EQUATION is** therefore, whenever a reaction equation is involved, always start by writing down the reaction equation (and please make sure that your reaction equations are correctly balanced). In “solid” calculations the reaction equation is very often given, whereas in “liquid” calculations this is **not** the case.
- **Question : If 25.0 cm<sup>3</sup> of 0.100 mol dm<sup>-3</sup> Na<sub>2</sub>CO<sub>3</sub>(aq) react with 35.0 cm<sup>3</sup> of HCl, what is the Concentration of the Acid?**
- Always
  - (A) Start with the **reaction equation**.
  - (B) Work out the **MOLE Reaction Ratio**.
  - (C) Work out the **number of moles used in the known solution**.
  - (D) Use the reaction ratio to work out the **number of moles of the unknown solution**.
  - (E) Use the **volume** of the unknown solution and the **number of moles** to work out its concentration.

**You MUST start with a balanced reaction equation. Write it out!**

*(Hint: what happens when you react a Carbonate with an Acid?)*



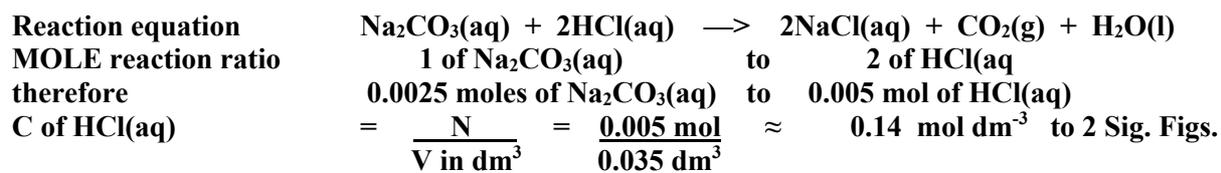
B) Reaction Ratio = 1 of Na<sub>2</sub>CO<sub>3</sub>(aq) : 2 of HCl(aq)

C) No. of moles of Na<sub>2</sub>CO<sub>3</sub>(aq) = Concentration (in mol dm<sup>-3</sup>) x Volume (in dm<sup>3</sup>)  
= 0.100 mol dm<sup>-3</sup> x 0.025 dm<sup>3</sup>  
= 0.0025 mol of Na<sub>2</sub>CO<sub>3</sub>(aq)

D) Reaction Ratio = 1 : 2 , therefore No of moles of HCl = 0.0025 mol x 2 = 0.0050 mol of HCl

E) There are 0.0050 moles of HCl in 35.0 cm<sup>3</sup> of solution, therefore  
Concentration = N (in mol) ÷ V (in dm<sup>3</sup>) , therefore C = 0.005 mol ÷ 0.035 dm<sup>3</sup>  
≈ 0.14286 mol dm<sup>-3</sup>  
= 0.14 mol dm<sup>-3</sup> (to 2 significant figures).

- Let me now show you a way of setting out the whole calculation – and I would advise you to ALWAYS set out your liquid calculations in this manner.



- That's it! You can do the whole calculation in FOUR lines, and in the exams that is exactly what you should do.