A First Year blog on Titration: 9th March 2019

(This is a **revision** blog. It is merely a *summary* of some of the things that you need to know. *Everything in this blog can be found in the relevant Chapters of the Book.*)

A consideration of Titration

- The normal way of conducting an exercise in titration is to place in a conical[1] flask a given volume (usually 25 cm^3 = 0.025 dm^3) of a solution of known concentration of **substance A** together with two drops of an appropriate indicator (a substance that changes colour when it is neutralised).

- A 50 cm^3 burette is then filled[2] with a solution of **substance B,** and then B is tritrated against **A** drop by drop whilst shaking the conical flask gently until A has been *exactly* neutralised by B. This point in the process is called the **Equivalence Point**, and it is usually indicated by the change in the colour of the indicator (the point where the colour changes is called the **End Point**). If an appropriate indicator has been used, then to all intents and purposes the End Point will be exactly the same as the Equivalence Point. As it happens, at the Equivalence Point a tiny increase in the amount of B in the mixture of A and B will change the pH value of the mixture by a very large amount (the change could be from pH=4 to pH=10 or vice versa, and a pH differential of 6 is a **1,000,000 fold difference** in the concentration of H^+ ions in the solution!) Therefore, even though the End Point of the indicator and the Equivalence Point of the two substances being titrated may not be exactly the same, nevertheless **the change in colour of the indicator (if chosen appropriately) will signal that the Equivalence Point of the titration has been reached**.

- Now, since both the concentration and the volume of substance A are known, then the number of moles of substance A can be calculated from the equation N = C x V (i.e. the number of moles = the concentration x the volume); and, since the volume of substance B can be ascertained by deducting the **initial** volume from the **final** volume of the titration exercise, then if the concentration of solution of B is known, the number of moles involved in the titration can be calculated (and hence the stoichiometric ratio of the reaction can be ascertained); or, if the reaction equation for A and B is known, then from the stoichiometric ratios involved, the concentration of solution B can be ascertained.

- That is all that there is to titration – but having said that, in practice the actual exercise is a seemingly complicated one for the novice titrator, and it is always a fiddly process even for the experienced titrator.

- One thing that I will stress and I cannot stress it too strongly is that **in the vast majority of cases the substances that are being titrated are highly aggressive substances and the wearing of safety goggles during the titration is not just necessary but it is COMPULSORY**. In an acid-base reaction the H^+ ions and OH^- ions are massively aggressive species and if they get into your eyes, then they could easily cause you to go blind.

- **PLEASE always wear you safety goggles when conducting experiments in the lab.**

In my lab you will **not** be allowed to conduct **ANY** experiments without the appropriate protection.

---

[1] Also called an "Erlenmeyer" flask.

[2] But the bottom of the curvature of the miniscus **must** be below the topmost graduated marking on the side of the burette – otherwise you will not be able to ascertain how much of solution B has been used in the titration exercise.

The Procedures in Titration

NB The solution of the substance whose concentration is known is called the "titrant" and is usually placed in the burette, and the one whose concentration is not known is called the "analyte" and is usually placed in a flat-bottomed conical "Erlenmeyer" flask. The following is the procedure for doing Titrations. Follow it, and you will never make a mistake in Titration.

- "Titration" is the process by which an unknown characteristic of a solution (e.g. its Concentration) can be ascertained by "titrating" that solution against (or dripping small amounts of that solution into) a solution whose characteristics ARE known, until the **exact** point is reached where one solution has just COMPLETELY reacted with the other solution.

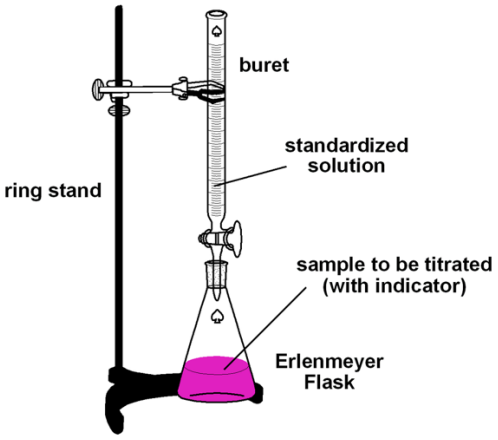
1) **START by writing down the Equation for the Reaction of A with B.**
   You cannot do any calculations without knowing the stoichiometry involved.
2) Derive the **MOLE Reaction Ratios** from the Reaction Equation.
3) Wash out your conical flask carefully[3] with some of the solution whose characteristics are known (let's call this solution A), and wash out your burette with some of the solution whose characteristics are unknown[4] (let's call this solution B).
4) Place a given volume (e.g. 25 cm^3) of solution A in the conical flask that you will place under your burette, and add two/three drops of the appropriately chosen **Indicator** to solution A in the conical flask.
5) Almost fill your burette with solution B. DO NOT FILL THE BURETTE ABOVE THE ZERO MARK![5] If you do so, then the invigilator in your Assessed Practical (i.e. your official State Practical exam) **will almost certainly have to (and SHOULD) fail you (because you will not be able to perform any of the calculations that are required)**!
6) Now take the reading of the burette from the bottom of the miniscus (this being your "**Initial Reading**"), and then titrate B against A until you obtain the exact point at which the reaction has fully taken place (this being indicated by the change in colour of the Indicator which you added to the conical flask). The last few drops of solution B **must** be dropped very slowly one drop at a time into the conical flask, while the flask is being shaken gently.
7) Take the reading of the burette from the bottom of the miniscus again (this now being your "**Final Reading**"), and the difference between the initial reading and the final reading is the exact amount of solution B that was needed to react with the 25 cm^3 of solution A.
8) Repeat steps (4) to (7) until you obtain two readings that are within 0.2 cm^3 of each other. [You should be able to achieve this within three trials i.e. one "ranging" shot, and two accurate ones, and these latter two trials should give answers that are within 0.2 cm^3 of each other – and preferably within 0.1 cm^3 of each other.]
9) From the known Concentration of A work out how many **moles of A** there are in the volume of solution A (i.e. in 25 cm^3).
10) From the Reaction Ratio, work out how many **moles of solution B** there must have been in the average volume of B, established by the Titration.
11) You now know (i) the volume of B and (ii) the number of moles of B, therefore you can now calculate the **Concentration of B** because $C = N \div V$ where

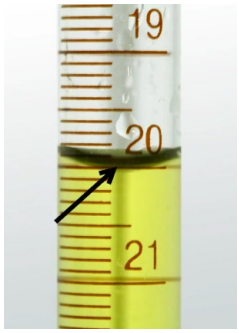
$$\textbf{Concentration} = \frac{\textbf{The AMOUNT of something, N (expressed in moles)}}{\textbf{The VOLUME involved, V (expressed in dm}^3\textbf{)}}$$

$$= \textbf{X mol dm}^{-3} \qquad [\text{X mol dm}^{-3} = \text{X moles per dm}^3]$$

[3] This reduces the risk of the flask being contaminated.
[4] This reduces the risk of the burette being contaminated.
[5] If the liquid rises above the zero mark, then you cannot calculate how much will have been used in the titration.

- Let me remind you of the set-up of for a burette/conical flask/clamp/etc. (The diagram below is from Sachi's Acids and Bases.) In England the burette is spelt differently from that below.





- An "**Erlenmeyer**" flask is just a name for a conical flask.

- If you see a conical flask with a tube in the neck of the flask, then it is one that can be used to create a vacuum in the flask to speed up *filtration*. It is called a **Büchner** flask/vacuum flask/filtration flask. You should not use a vacuum flask for titration. You could break the little protruding tube at the neck of the flask when you are swirling the flask and then you will cut yourself.

- Please take your readings from the bottom of the miniscus (as in the right hand diagram).

Inaccuracies in Measurement

In Titration,
- the term "ranging shot" establishes a close estimate of the value that is being established, while
- the term "precision" refers to how close successive measurements are to each other, and
- the term "accuracy" refers to how close a measurement comes to its true/known to be true value.

- Let us say that a manufacturer of burettes makes his glass tube and puts the marks on it showing 0 cm^3 to 50 cm^3, but finds that sometimes the readings are 0.05 cm^3 more than the true amount and at other times they are 0.05 cm^3 less than the true amount. He would then mark his burette as $\pm$ 0.05 cm^3.

- Please note that each time someone takes a measurement from the above instrument, there could be an error of $\pm$ 0.05 cm^3, therefore if a titre had been 17.3 cm^3 (from an initial reading of let's say 0.3 cm^3 and a final reading let's say 17.6 cm^3) then **the percentage error involved** would be two lots of $\pm$ 0.05 cm^3 (*one error in taking the 0.3 cm^3 measurement, and one error in taking the 17.6 cm^3 measurement*) in relation to 17.3 cm^3 expressed as a percentage

$$= \frac{2 \times (\pm\ 0.05\ cm^3)}{17.3\ cm^3} \times 100\% = 0.58\%.$$

If I wanted to reduce the error involved I could (i) **halve** the *concentration* of the solution in the burette (the titrant) and then I would need twice the volume of the titrant (=34.6 cm^3) and my error would halve to 0.29%, or I could (ii) **double** the *concentration* of the solution in the conical flask (the analyte) and again I would need 34.6 cm^3 of the titrant in the burette to neutralise the analyte in the conical flask, and my percentage error would halve to 0.29%.

Please remember that 1 metre = 10 decimetres (dm) = 100 centimetres (cm) = 1,000 millimetres (mm), and (by definition)

1 litre = 1dm x 1dm x 1dm = **1 cubic decimetre (1 dm^3)**
and since 10 centimetres = 1 decimetre, then 1 litre also
= 10cm x 10cm x 10cm = **1,000 cm^3 (1,000 cubic centimetres)**
= 1,000 millilitres (**1,000 ml**)
ALL these units represent EXACTLY the same volume (and 1 litre of Water has a mass of 1kg).

1 litre = 1 dm^3 (= 1,000 cm^3 = 1,000 ml)

and that $\text{Concentration} = \frac{\text{Number or Amount of things}}{\text{Volume of liquid in dm}^3}$

$$\mathbf{C = \frac{N}{V \text{ in } dm^3}} \quad \textbf{, and } \mathbf{N = C \times V \text{ in } dm^3}$$

NB The unit of concentration is "dm^3 ", therefore **if you are given a volume in cm^3 you MUST convert it into dm^3 by dividing the cm^3 number by 1000**
e.g. 17.3 cm^3 = $\frac{17.3}{1000}$ dm^3 = 0.0173 dm^3

This is the sort of calculation involved in a typical acid/base reaction titration.

Question : If 25.0 cm^3 of 0.100 mol dm^{-3} $Na_2CO_3(aq)$ react with 35.0 cm^3 of HCl, what is the Concentration of the Acid?

Answer
No. of moles of Na_2CO_3 = C x V in dm^3 = 0.100 mol ~~dm^{-3}~~ x (25.0 x 10^{-3} ~~dm^3~~) = 2.5 x 10^{-3} mol
Reaction equation $Na_2CO_3(aq) + 2HCl(aq) \longrightarrow 2NaCl(aq) + CO_2(g) + H_2O(l)$
MOLE reaction ratio 1 of $Na_2CO_3(aq)$: 2 of HCl(aq
therefore 0.0025 moles of $Na_2CO_3(aq)$: 0.005 mol of HCl(aq)
C of HCl(aq) = $\frac{N}{V \text{ in } dm^3}$ = $\frac{0.005 \text{ mol}}{0.035 \text{ } dm^3}$ ≈ 0.143 mol dm^{-3}

The numbers in front of any of the substances (i.e. 1/2/3/4/etc) will affect your MOLE calculations, **but you do NOT need them after that**!

Once you have calculated the MOLE Reaction Ratio, the number in front of the substance does not come into the calculation again!