Titrations

Students struggle with titrations as they don't understand what is going on experimentally. There is often a lot of information and it can be confusing. **The maths involved is very simple**, so don't think you can't do it because of maths.

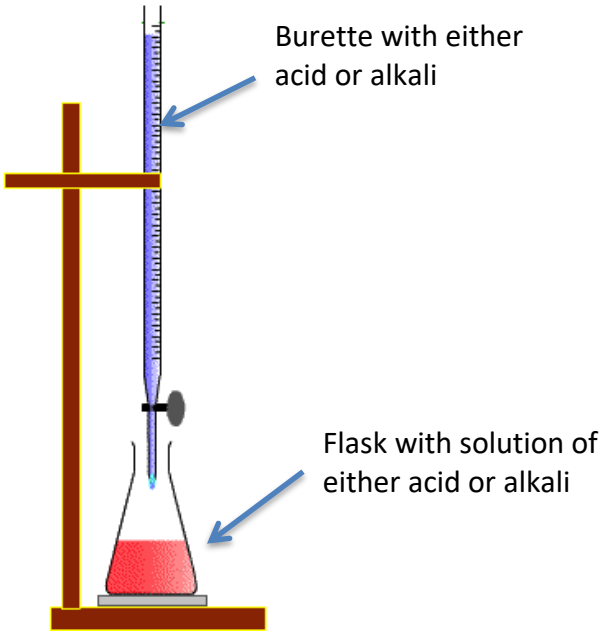
The good news is that they are all just simple **acid-base** neutralisation reactions. And the calculations are very similar to those in the moles tutorials that utilise chemical equations.

General titration explanation

The point of any titration is to allow you to calculate the number of moles of a reactant that you don't know very much about (**the "unknown"**). From there you can work out whatever you want → concentration, percentage, grams etc; exactly like any standard moles calculation.

To be able to do this, the unknown is reacted with something of **known concentration** (we'll call this **the "known"**).

The apparatus is very simple. Just a burette and a conical flask. You can put either solution in the burette but it is **usually** the one of known concentration.



Standard Solution

A standard solution is a solution of **known concentration** and is made up in a **volumetric flask**. You can then take small portions of this to carry out several titrations. This solution is usually in the burette but doesn't have to be, it could go in the conical flask.

Calculate the mass of solid required for desired concentration and accurately weight out this mass in a weighting boat.

Dissolve this solid in deionised (distilled) water and transfer to a volumetric flask e.g. 250 cm^3.

Rinse the weighing boat with the water to ensure no solid is left behind and transfer washings to the volumetric flask.

Make up to the mark with water in the volumetric flask.

Invert or swirl volumetric flask to ensure it is all mixed.

✓ Students often forget about the rinsing and inverting steps.

Rinsing

Rinse the **burette** with whatever solution is going to put in it i.e. acid or alkali. You can rinse it with water first as long as you then rinse it with acid or alkali.

This is a classic exam question….what happens to the titre if you do this rinsing wrong? Any traces of water in the burette will dilute the solution of known concentration, which defeats the purpose of making it accurately in the first place. Therefore the concentration will be slightly less than the solution you made and you have to then add more of the solution to get a complete reaction i.e. the titre is **too high** → inaccurate calculation.

You can rinse the **pipette and conical flask** with **water**. This is ok as the important point is the number of moles in the flask. Adding water does **not affect the moles**. It will of course affect the concentration but isn't important here.

Doing the titration

Accuracy is the key to these experiments. You should use a **pipette** when measuring out volumes.

Ensure that the "tip" of the burette is filled with solution and that there are **no air bubbles**. Otherwise a titre that is **too high** will be obtained.

Add a few drops of an **indicator** to the conical flask:

methyl orange: changes colour in **pH range 2-5** from **orange to red** when adding acid to alkali

phenolphthalein: changes colour in **pH range 8-10** from **colourless to pink** when adding alkali to acid.

All you have to do is open the tap on the burette and look for a colour change. This is the **end point** i.e. where neutralisation occurs. Note down the volume added from the burette (this is the **titre**).

You just have to be careful to **add the solution slowly** towards the end point or you could miss the colour change and your values will be inaccurate. Also, put a **white tile** under the conical flask to enable you to see the colour change more easily. Always give the flask a **swirl** when adding the solution from the burette to ensure that the solutions are actually mixing.

Reliability = repeat

You should do a **rough titration** first so you know approximately when the colour change occurs. You can take your time in the next titration when you get close to the point and drip the solution in carefully to get an accurate value.

You should **repeat** the titration until you get two values within **0.1 cm^3** of each other, known as **concordant values**. The more times you do it, the more **reliable** your result will be. You take the **average** of these concordant titres as your value to use in the calculation. Ignore any values not within 0.1 cm^3.

Calculations

They are ALL simple neutralisation reactions at AS level. So remember the equation:

Acid + Base → Salt + Water

Let's look at an example and work through what is going on and how you would do a calculation like this. They want to know the concentration of ethanoic acid. The equation for the reaction is:

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

25 cm^3 of an ethanoic acid solution was transferred to a 250 cm^3 flask and made up to the mark with water. 25 cm^3 of this solution was then pipetted into a conical flask and titrated against NaOH. It took 28 cm^3 of 0.1 mol dm^{-3} NaOH to neutralise the ethanoic acid. What is the concentration of the ethanoic acid?

Further Explanation

Before we dive into the calculation, I am going to explain what they are doing experimentally, which will help you massively with the calculations. If you understand why they are doing these steps then you won't freak out and start multiplying random numbers together.

Preparation Steps

Think of the first part of the question as "**preparation**". The titration is always at the **end** of the question.

It's rare that you can just take a sample and titrate it straight away. So they usually do a bit of preparation to make it "titrateable" (that's not really a word, I just made it up but you get what I mean right?). The good thing is that the preparation part is always very similar. And very simple.

<u>Dilution</u>

In this case the original ethanoic solution is probably too concentrated to titrate and would require a huge titre. So they dilute it. And that is very common in these questions. They add the solution to a volumetric flask then and **add water** up to 250 cm^3, so it is diluted by a factor of 10. The important point:

Adding water lowers the concentration but does not, I repeat, DOES NOT, affect the number of moles.

So we still have the **same number of moles** in the volumetric flask as we had originally.

✓ Don't confuse this with the standard solution. This ethanoic solution is of unknown concentration whereas a standard solution is of known concentration. They just both happen to use a volumetric flask.

Remove a portion

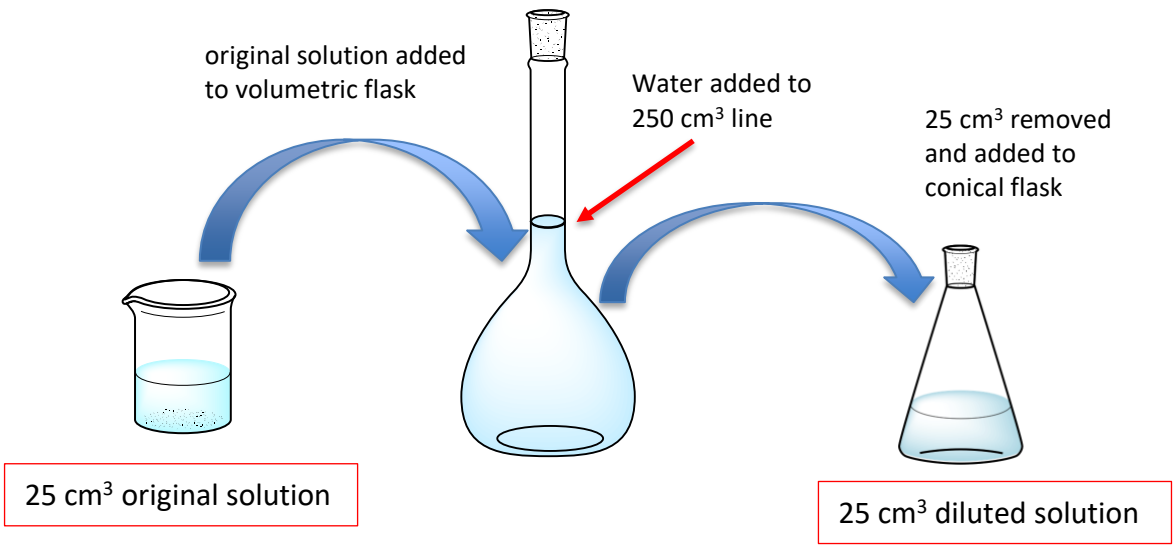
Next they usually but not always, **remove** *some* of the 250 cm^3 solution for the titration and put it in the conical flask. Why? Well, 250 cm^3 is a helluva volume to titrate! You'd need a monster burette if such a thing exists or have to keep refilling it which introduces errors.

So they remove a **small manageable amount**, in this case, 25 cm^3. It doesn't have to be this amount, it could be any amount. Also, by removing only a small amount, it means that they can do multiple titrations from the same "batch" in the volumetric flask. This reduces errors and saves you making up fresh solutions. The important point:

Removing a portion (25 cm^3) is taking 1/10th of the moles from the volumetric flask

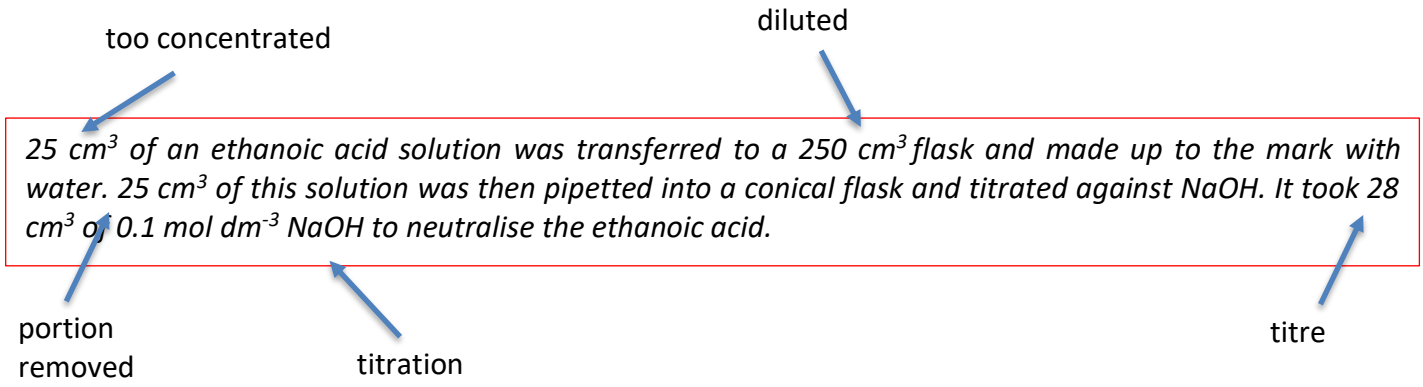
We are only titrating 1/10th of the original moles. We need to take account of this in the calculation because we always want the original amount.

Below is a diagram to illustrate what is going on in this preparation part. It's so easy it's crazy. Add some water and then take some of the solution out! But you need to be aware of this as it makes the questions so much easier. I'm also quite proud of the diagram so wanted an excuse to put it in!



In summary:

Let's look at the question again armed with this knowledge of what is going on. Now all the numbers should make sense. There should be no freaking out and randomly multiplying numbers. You should now know **exactly** what is going on. When you read questions from now on, this is the sort of analysis you need to be doing:



✓ If you understand this, then this will help immensely from now on and into harder calculations in A2. The students that struggle have no idea what's going on experimentally but blame it on the calculations. Understand first ☺

Calculation Structure

To do these calculations you have to **start at the end** of the question and **work backwards**. The titration is the last step and that's where we have to start. Two key points:

- as a general rule, **look for the titre** as you will *usually* use this value **first** (but not always).
- usually the **first number** in the question is used **last** in the calculation.

25 cm^3 of an ethanoic acid solution was transferred to a 250 cm^3 flask and made up to the mark with water. 25 cm^3 of this solution was then pipetted into a conical flask and titrated against NaOH. It took 28 cm^3 of 0.1 mol dm^{-3} NaOH to neutralise the ethanoic acid. What is the concentration of the ethanoic acid?

So now we know **why** they are doing each step and also roughly the **structure** of a standard question. If you understand both these points, then the calculation becomes a joke! If you don't get this, then the calculations will still be a problem….forever ☹

Step 1: moles = conc x vol (known substance)

Always always work out the moles of the substance of **known concentration**. We can see that this is NaOH. We must use the 28 cm^3 as that is the titre associated with the NaOH; the other volumes have nothing to do with NaOH!

moles = 0.1 x 28/1000 = **2.8×10^{-3} moles** of NaOH

Step 2: use the equation for ratios

This is always the 2nd step. We want to know the number of moles of CH_3COOH so we need to use the equation to see how the CH_3COOH is related to the NaOH:

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

We can see that the CH_3COOH:NaOH **ratio is 1:1** therefore the CH_3COOH moles are also:

2.8×10^{-3} moles

Step 3: look for portions

We know they took a 25 cm^3 portion from the 250 cm^3 volumetric flask. We said that they have taken **1/10th of the original moles** to do the titration. **So you need to multiply by 10** → original moles.

$$2.8 \times 10^{-3} \times 10 = \mathbf{2.8 \times 10^{-2}\ moles\ in\ 250\ cm^3}$$

Step 4: Answer the question (convert moles into something else)

The last part should be easy. Just answer the question.

What are they looking for? Are they looking for the mass in grams, a percentage, a concentration etc.

In this question, they are looking for concentration, so you just use moles = conc x vol again. But it's the original concentration. Remember we usually use the first number at the end if you get confused. The 250 cm^3 is irrelevant, that was just a practical measure as discussed above. We use the original 25 cm^3:

$$c = n/v \rightarrow c = 2.8 \times 10^{-2} / 0.025\ (25/1000)$$

$$= \mathbf{1.12\ mol\ dm^{-3}}$$

✓ Concentration is very commonly asked for. Calculating the RFM of a compound is as well. In this case you do **mass/moles** at the end…always! Probably the two most common questions at AS.

Percentage Error

As titrations are all about accuracy, percentage error calculations often appear. The percentage error in using a piece of apparatus is given by:

% error = error in the apparatus/value x 100

Everytime that you read a burette, thermometer, balance etc. there is a small error associated with the value that you read. For example, if you read the volume in a burette as 12.5 cm^3, you can't be certain that it is 100% correct.

Error in the apparatus

For things like volumetric flasks and pipettes where there is only **one reading** possible, the error is **written on the apparatus**.

For things like burettes, thermometers, measuring cylinders and balances the error is **half of the smallest division**. So if the graduation on a burette is 0.1 cm^3 (difference between each line), then you need to half it → **0.05 cm^3**

Examples

250 cm^3 volumetric flask: error given as 0.2 cm^3. Therefore we do:

0.2/250 x 100

= 0.08 %

Burette with 0.1 cm^3 divisions: error is half the smallest division → 0.05 cm^3. Titre obtained = 17.1 cm^3:

0.05/17.1 x 100

= 0.29 %

But…as we need to read the **burette twice** (initial and final volumes) we must multiply by 2 as there will be two errors (one for each reading):

0.29 x 2 = 0.58 %

Reducing % error

Use the above percentage error equation. It's common sense. If you want a smaller error, then either make the value of the bottom line smaller or the top line bigger in the equation.

You can do this by:

1. Use more accurate apparatus to **reduce** the value on the top line.

2. More commonly in questions the answer is to obtain a **larger titre** to increase the value on the bottom line. You can do this by **decreasing the concentration** of the solution in the burette or use **more** of whatever is in the flask.