

Finding out How Much Acid there is in a Solution

When a metal is extracted from its ore, a waste product is often sulphur dioxide (SO_2). This is then converted into sulphuric acid (H_2SO_4) and sold. To sell sulphuric acid, its accurate concentration must be known. In this investigation, I will use a titration method to attempt to find the accurate concentration of a sample of sulphuric acid thought to have a concentration of between 0.05 and 0.15 mol dm^{-3} . The chemicals which I will use to do this will be solid anhydrous sodium carbonate (Na_2CO_3) and an indicator.

Finding out How Much Acid there is in a Solution - Plan

Quantities of Chemicals Required

In this investigation, the chemical reaction used will be the neutralisation reaction between sulphuric acid and sodium carbonate:



As the sodium carbonate has to be aqueous for this reaction, I will need to make up a solution of sodium carbonate before I begin the titration. I will need to choose an appropriate concentration for this solution and an appropriate volume of it to make up. As there is a 1:1 ratio between the moles of sulphuric acid and sodium sulphate used in this reaction, I should make up a solution of sodium carbonate of similar solution to that of the sulphuric acid. Assuming that the concentration of the sulphuric acid is around 0.1 mol dm^{-3} , this will be the concentration of the sodium carbonate solution I make.

Next I must decide what volume of solution to make up. Firstly I need to decide the volume of solution I will require per titration. This volume will be measured using a pipette and filler, so only set volumes, such as 10cm^3 , 25cm^3 and 50cm^3 are available. If I use 10cm^3 , the percentage error will be too large to produce reliable results, and if I use 50cm^3 , it is likely that I will need too much acid for a 50cm^3 burette. Therefore the best volume of solution to use per titration will be 25cm^3 , as it allows the acid to have a concentration greater than or less than the assumed 0.1 mol dm^{-3} . To produce reliable results, I will need to complete three titrations with a range of no more than 0.1cm^3 . Taking into account a rough titration and any anomalous results, I will probably need to do at least six titrations, which will require 150cm^3 of sodium carbonate solution. The solution will be made up in a volumetric flask, which has set volumes. I will make up a solution of volume 250cm^3 , which will allow a maximum of nine titrations plus a small amount of solution to wash out apparatus.

To make up a volume of 250cm^3 at concentration 0.1 mol dm^{-3} , I will need to work out the mass of solid sodium carbonate I will need to add to deionised water. First I will work out the number of moles of sodium carbonate in this solution:

$$\text{moles Na}_2\text{CO}_3 = \frac{\text{concentration Na}_2\text{CO}_3 \times \text{volume of Na}_2\text{CO}_3 \text{ solution}}{1000}$$

$$\text{moles Na}_2\text{CO}_3 = \frac{0.1 \times 250}{1000} = 2.5 \times 10^{-2} \text{ moles}$$

I can use this value for the number of moles to work out the mass of sodium carbonate required for this solution:

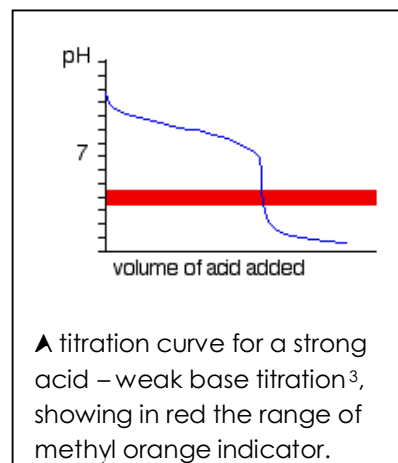
$$\text{mass Na}_2\text{CO}_3 = \text{moles Na}_2\text{CO}_3 \times \text{mass of one mole Na}_2\text{CO}_3$$

$$\text{mass Na}_2\text{CO}_3 = 2.5 \times 10^{-2} \times ((2 \times 23) + 12 + (3 \times 16))$$

$$\text{mass Na}_2\text{CO}_3 = 2.65 \text{ g (to 3 s.f.)}$$

Indicator Choice

I need to decide which indicator will be most suitable to use in my investigation. Using the information sheet on the use of indicators in acid-alkali titrations¹, this is a titration using a strong acid and a weak alkali, so methyl orange should be used as an indicator. This is because methyl orange changes colour in the pH range 3.1 – 4.4². The titration curve shown here³ for a typical strong acid – weak base titration shows that the equivalence point for this titration is below 7 – around the pH range of methyl orange, showing that methyl orange is a good choice for an indicator for my investigation.



Equipment List

This is a list of all the laboratory equipment I will need for my investigation:

- 50cm³ burette
- 25cm³ pipette and filler
- 250cm³ volumetric flask
- Conical Flask
- Safety goggles
- Several small beakers
- Digital balance
- Spatula
- Glass rod
- Filter funnel
- Dropping pipette
- Clamp and clamp stand
- White tile

This is a list of all the chemicals I will need for my investigation:

- Sulphuric acid (H₂SO₄ (aq)), concentration unknown
- Sodium carbonate solid (Na₂CO₃ (s)), mass 2.65g

- Deionised water
- Methyl orange indicator

Risk Assessment⁴

I need to complete a risk assessment for all of the chemicals I will use in my investigation, as well as all those that are produced. I need to assess what I should do if they are spilt in the laboratory, on the skin, or in the eyes.

• Sodium Carbonate (Na_2CO_3)



- Can be irritant to the skin, the eyes and the respiratory system
- Minimal hazard unless ingested in large quantities
- **Spilt in laboratory:** scoop up, add mineral absorbent in solution, rinse area with water
- **Spilt on skin:** wash area with water
- **Spilt in eye:** flood eye with water for at least 10 minutes, seek medical advice

• Sulphuric Acid (H_2SO_4)



- Corrosive at higher concentrations, but irritant if below 0.5 mol dm^{-3}
- **Spilt in laboratory:** wash area with water
- **Spilt on skin:** rinse affected area with plentiful water
- **Spilt in eye:** flood eye with water for at least 10 minutes, seek medical advice

• Methyl orange

- Minimal hazard as only using a few drops
- **Spilt in laboratory:** wash area with water
- **Spilt on skin:** rinse affected area with water
- **Spilt in eye:** flood eye with water for at least 10 minutes, seek medical advice

• Sodium Sulphate (Na_2SO_4)



- Can be harmful but not at the small quantities that will be produced in my investigation
- **Spilt in laboratory:** rinse out area with water
- **Spilt on skin:** wash off with water
- **Spilt in eye:** flood eye with water for at least 10 minutes, seek medical advice

• Water (H_2O)

- No risk with small volumes produced in this investigation

• Carbon Dioxide (CO_2)

- No risk with small volumes produced in this investigation

Method – Making up Sodium Carbonate Solution

This is an outline of how I will prepare a solution of sodium carbonate from the solid I am given. After putting on my safety goggles, I will put a small beaker on a balance and reset the balance to zero. I will then weigh out 2.65g of sodium carbonate solid using a spatula. I will add a small volume of deionised water and dissolve the solid in this, stirring with a glass rod.

When all of the solid is dissolved, I will pour the solution into a 250cm³ volumetric flask using a funnel. I will then wash out the beaker and glass rod I used with deionised water and pour the washings into the volumetric flask. I will repeat this to ensure that all of the solution is in the flask. Using more deionised water, I will fill up the volumetric flask up to the 250cm³ mark, using a dropping pipette to add the final few drops, taking great care not to go over the mark. Finally, I will stopper the flask, label it and then invert the flask fully at least 10 times, to ensure the concentration is equal throughout the solution.

Method – Checking the Indicator Colour Change

I need to check the exact colour which the indicator turns my acid and alkali solutions so I can be sure when the titration is finished and the substance is neutralised. I will pour a little of the acid and the sodium carbonate into beakers and put a little indicator in each solution, keeping these to have a record of the colour change.

Method - Titration⁵

Now I will be ready to complete the main part of my investigation – the titration. After putting on my safety goggles, I will wash out the burette with the sulphuric acid solution and the pipette with the sodium carbonate solution. I will vertically clamp the burette and then fill it with acid solution. I will let some acid run through the jet and then take the initial burette reading, measuring to two decimal places. I will record all measurements in a table with these headings:

Titration	Rough	1	2	3	4	5
Final burette reading (cm ³)						
Initial burette reading (cm ³)						
Titre (cm ³)						

I will then use the pipette and filler to transfer 25cm³ of sodium carbonate solution from the volumetric flask to a conical flask. I will add four drops of methyl orange indicator to this and place it under the burette, on a white tile to make it easier to see the colour change. Then I will complete the first titration, a rough titration – I will add acid from the burette into the conical flask whilst swirling the flask, stopping when I observe the colour change seen when I checked the indicator colour change. I will record the final burette reading to two decimal places in the table. I will then repeat the titration until I have three results, which are within 0.1cm³ of each other, excluding

In the devising of this plan, I have taken several steps to ensure that it will provide reliable results:

- I will wash out all equipment I have used to make up the sodium carbonate solution with deionised water, ensuring that no sodium carbonate is lost in the apparatus and the concentration of the sodium carbonate solution is precise.
- When the sodium carbonate solution is made up, I will invert the flask at least ten times to ensure that the concentration is equal throughout the solution.
- I will have a record of the colour change so that I will know exactly when the solution has been neutralised, so the concentration can be found accurately.
- I will use a white tile during the titration to help me see exactly when the colour change has taken place.
- I will measure all burette readings to two decimal places, to reduce the percentage error of the burette, so I can be sure that my results are accurate.
- I will repeat the titration until I have three results within 0.1 cm^3 of each other. This will ensure that the average titre, and hence, the concentration found is accurate and does not include any anomalous results.

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² <http://www.avogadro.co.uk/chemeqm/acidbase/titration/phcurves.htm>, a page explaining titration curves and the effect of the strength of the acid and base uses, from a website “designed for students who are studying Chemistry at AS-Level and A-Level”, ©2000-2007 Avogadro.co.uk

³ Image of a titration curve (edited) taken from http://www.alevelchemistry.co.uk/Module_4/HTML%20Files/equilibrium/4.3_equilibrium_notes.htm, a page about equilibria from a website providing students with help with A-Level Chemistry, © alevelchemistry.co.uk

⁴ Risk assessment completed using the Hazcards – School Science Service, Brunel University, Uxbridge, UB8 3PH; printed 1989, reprinted 1992

⁵ Method for titration and table of results based on method found on ~~M2. *Winning* *now* *in* *the* *classroom* *is* *in* *solution*, *School Science Service*, Heinemann Educational Publishers, Halley Court, Jordan Hall, Oxford, OX2 8EJ, published 1994 reprinted 2000~~

Finding out How Much Acid there is in a Solution - Results

I have now completed the practical part of my investigation; the results are shown below.

Table of Results

This table shows the results of all the titrations that I completed. All results are to two decimal places and anomalous results written in red.

Titration	Rough	1	2	3	4	5
Final burette reading (cm ³)	1.90	1.55	0.40	0.90	0.70	0.85
Initial burette reading (cm ³)	29.35	28.10	26.75	27.35	26.50	27.25
Titre (cm ³)	27.45	26.65	26.35	26.45	25.80	26.40

Average Titre

Using the three accurate results from my table, I will come up with an average titre to use to calculate the concentration of the acid:

$$\text{Average Titre} = \frac{26.35 + 26.45 + 26.40}{3}$$

$$\text{Average Titre} = 26.40\text{cm}^3$$

Moles of Sulphuric Acid used per Titration

To find the concentration, I must first know the number of moles of sulphuric acid used in each titration. First, I must know how many moles of sodium carbonate were used per titration

$$\text{Total moles Na}_2\text{CO}_3 = \frac{\text{mass Na}_2\text{CO}_3}{\text{mass of one mole Na}_2\text{CO}_3}$$

$$\text{Total moles Na}_2\text{CO}_3 = \frac{2.65}{106} = 2.50 \times 10^{-2} \text{ moles}$$

$$\text{Concentration Na}_2\text{CO}_3 = \frac{\text{moles Na}_2\text{CO}_3 \times 1000}{\text{volume of Na}_2\text{CO}_3 \text{ solution}}$$

$$\text{Concentration Na}_2\text{CO}_3 = \frac{2.5 \times 10^{-2} \times 1000}{250} = 0.10 \text{ mol dm}^{-3}$$

$$\text{moles Na}_2\text{CO}_3 \text{ per titration} = \frac{\text{concentration Na}_2\text{CO}_3 \times \text{volume of Na}_2\text{CO}_3 \text{ solution per titration}}{1000}$$

$$\text{moles Na}_2\text{CO}_3 \text{ per titration} = \frac{0.1 \times 25}{1000} = \mathbf{2.5 \times 10^{-3} \text{ moles}}$$

I can use the chemical equation for the titration to work out the mole ratio between sulphuric acid and sodium carbonate:



As there is a 1:1 ratio between the moles of sulphuric acid used and the moles of sodium carbonate used, I can conclude that 2.5×10^{-3} moles of sulphuric acid are used per titration.

Concentration of Sulphuric Acid

Using the data for the average titre (26.40cm^3) and the moles of sulphuric acid (2.5×10^{-3} moles), I can work out the concentration of the acid:

$$\text{Concentration H}_2\text{SO}_4 = \frac{\text{moles H}_2\text{SO}_4 \times 1000}{\text{average titre}}$$

$$\text{Concentration H}_2\text{SO}_4 = \frac{2.5 \times 10^{-3} \times 1000}{26.40}$$

$$\text{Concentration H}_2\text{SO}_4 = \mathbf{9.47 \times 10^{-2} \text{ mol dm}^{-3} \text{ (to 3 s.f.)}}$$

Therefore I have completed my investigation, as I have found out how much acid there is in the solution, as the acid has a concentration of $9.47 \times 10^{-2} \text{ mol dm}^{-3}$, to three significant figures. This is a similar result to the expected result, which was between 0.05 and 0.15 mol dm^{-3} .

Finding out How Much Acid there is in a Solution - Evaluation

Overall, I believe that my investigation has been a success. The concentration that I have found for the sulphuric acid is almost exactly in the middle of the range given, showing that my method worked well and the procedure was a success.

Anomalous Results

Two of the titrations that I carried out were anomalous; they fell outside of a range of 0.1cm^3 for the three successful titrations. The anomalous titrations were the first and fourth titrations following the rough titration:

Titration	1	4
Final burette reading (cm^3)	1.55	0.70
Initial burette reading (cm^3)	28.10	26.50
Titre (cm^3)	26.65	25.80

There are several possible causes for these anomalies. For the first titration, as it was the first to be completed after the rough titration, we still were not certain exactly where the end point fell. As our rough titration turned out to have a slightly greater volume than our final average titre, it is possible that this is why we went a little too far on the first titration as well. The fourth titration was the first to be carried out on the second day of our investigation, so it is possible that we could have been less sure of the end point than on the first day. The anomalies could also be down to various experimental errors, such as an equipment error.

Procedural Errors

The most likely cause of error to my experiment will have been errors in the procedure. I will examine some of these and the possible effect they could have had on the final outcome of my investigation.

- Firstly, there is the possibility of error on deciding upon the correct end point for the titration. This is entirely down to human judgement, based on the colour of the indicator in the acid and alkali solution. Even though I took many precautions to reduce end point errors, there is still a high possibility of making errors here. If, for example, I suggest that I, on average, overshoot the end point by 1cm^3 , making my average titre 27.40cm^3 , I can work through the calculations to see what effect this had on my final result for the concentration of the acid:

$$\text{Concentration H}_2\text{SO}_4 = \frac{\text{moles H}_2\text{SO}_4 \times 1000}{\text{average titre}}$$

$$\text{Concentration H}_2\text{SO}_4 = \frac{2.5 \times 10^{-3} \times 1000}{27.40}$$

$$\text{Concentration H}_2\text{SO}_4 = 9.12 \times 10^{-2} \text{ mol dm}^{-3} \text{ (to 3 s.f.)}$$

This is a change of $3.5 \times 10^{-3} \text{ mol dm}^{-3}$, which is a 3.70% change from the actual calculation, which is quite a large change for a difference of just 1 cm^3 .

- Another potential procedural error could have occurred in the making up of the sodium hydroxide solution. I took several washings to ensure that all the sodium carbonate was in the solution, but there is a possibility that the concentration was not quite accurate. I will examine the effect of a very slight change in concentration of $0.001 \text{ mol dm}^{-3}$, reworking the calculations using a concentration of $0.099 \text{ mol dm}^{-3}$:

$$\text{moles Na}_2\text{CO}_3 \text{ per titration} = \frac{\text{concentration Na}_2\text{CO}_3 \times \text{volume of Na}_2\text{CO}_3 \text{ solution per titration}}{1000}$$

$$\text{moles Na}_2\text{CO}_3 \text{ per titration} = \frac{0.099 \times 25}{1000} = 2.48 \times 10^{-3} \text{ moles}$$

This means that 2.48×10^{-3} moles of sulphuric acid were used, meaning that:

$$\text{Concentration H}_2\text{SO}_4 = \frac{\text{moles H}_2\text{SO}_4 \times 1000}{\text{average titre}}$$

$$\text{Concentration H}_2\text{SO}_4 = \frac{2.48 \times 10^{-3} \times 1000}{26.40}$$

$$\text{Concentration H}_2\text{SO}_4 = 9.38 \times 10^{-2} \text{ mol dm}^{-3} \text{ (to 3 s.f.)}$$

This very slight concentration change gives a difference of 0.95%, which is still fairly large for such a small concentration change. However, the steps I took make this small change the very maximum that could have feasibly occurred.

Equipment Errors

Four pieces of equipment that I used could have had an error associated with them: the balance, the volumetric flask, the pipette and the burette. I will examine these errors in the form of percentage errors to look at their potential effect on the final value for the acid concentration (all to 2d.p.).

- **Balance:** $\% \text{ error} = \frac{0.005 \times 2}{2.65} \times 100 = 0.38\%$
- **Volumetric Flask (250cm³):** $\% \text{ error} = \frac{0.15}{250} \times 100 = 0.06\%$
- **Pipette (25cm³):** $\% \text{ error} = \frac{0.06}{25} \times 100 = 0.24\%$
- **Burette:** $\% \text{ error} = \frac{0.05 \times 2}{26.40} \times 100 = 0.38\%$

From this, I can calculate that the total percentage error for the equipment of my investigation is **1.06%**. The effect of this on my final value for the acid concentration is

given by the percentage of the final value that this percentage corresponds to. My final value for concentration was $9.47 \times 10^{-2} \text{ mol dm}^{-3}$, and 1.06% of this is 1.00×10^{-3} , so therefore the final value for my concentration can be written as **$9.47 \times 10^{-2} \pm 1.00 \times 10^{-3} \text{ mol dm}^{-3}$** .

Overall greatest sources of errors

Looking at the sources of error, the possibility of having an incorrect end point appears to be the greatest potential source of error – overshooting the end point by 1 cm^3 would yield a percentage change three times the size of the total percentage equipment error. Of the equipment errors, the greatest was the balance, but this was still only 0.38%.

Procedures that reduced errors

During the course of my investigation, I took several precautions that should have reduced errors, which should have contributed to my final result being accurate. These procedures included:

- I repeated the titration until I had completed three titrations that had results within 0.1 cm^3 of each other. This should have ensured that my average titre was an accurate value for the acid used and not affected by any anomalous results.
- When I was making up the solution of sodium carbonate, I washed out all equipment used more than once, adding the washings to my solution. This should have meant that as much of the sodium carbonate as possible became part of the solution, so I knew very accurately the concentration of the sodium carbonate solution.
- I ensured the colour for the end point was correct by keeping a sample of the initial titration, so that I measured to the same end point every time
- I completed the titration using a white tile so that it was easier for me to tell exactly when the end point for the titration occurred, reducing the chance of overshooting, for example.
- All values for the titration were measured to two decimal places, providing greater accuracy, and hence less potential for error in the calculations

Improvements that could reduce errors

Whilst I have assessed that my investigation will have produced reliable results, it was not perfect, and there are improvements that could be made to reduce errors, such as:

- I could use a pH meter instead of indicator to more accurately find the end point of the titration, which would remove all procedural error from the process of finding the end point and would leave only the equipment error of the pH meter, which would be a lot smaller.

- Similarly, I could use a data logger to measure the pH of my solution, which would not only give me a more accurate end point, but would show me the progression of the pH of the solution, making further titrations more accurate
- To reduce the equipment error of the balance and the volumetric flask, I could weigh out more sodium carbonate and hence make a greater volume of sodium carbonate solution. Although this would probably leave some waste solution at the end, it would serve to reduce the percentage error of the balance and the volumetric flask, which are inversely proportional to the mass and the volume measured respectively.

In summary, I have assessed that my investigation produced reliable results, identified and attempted to explain any anomalous results, investigated potential sources of error and looked at their possible effect on the final concentration of acid, and I have suggested some improvements that could be made to reduce these errors if I were to do this investigation again.