



TITRATION LAB REPORT

Abstract

In this report we introduce analytical chemistry through an acid-base titration. Summarising the underlying scientific principles and experimental technique used during the experiment.

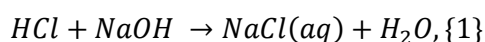
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Introduction

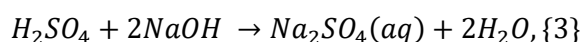
Titration is a cornerstone of practical and analytical chemistry, introducing experimental evidence for the theoretical concepts of stoichiometry and the equivalence point. The purpose of any titration is to determine the molarity of an unknown solution. Achieved by bringing an unknown solution of acid/base to neutral with a known concentration standard solution of a corresponding base/acid. An indicator solution typically used to show the equivalence point which is the point at which there are chemically equivalent quantities of reactants [1]. In the experiment a standard solution of sodium hydroxide was used to neutralise a solution of hydrochloric acid with unknown concentration using phenolphthalein indicator solution. The reaction is as follows:



From the stoichiometry of {1} we have a 1:1 ratio and so the number of moles of NaOH required to create a neutral solution is the number of moles of acid present. The molarity of the acid solution can then be calculated using the number of moles obtained and the known volume of the HCl.

$$M = \frac{n}{V}, \{2\} \text{ Equation of molarity}$$

Here M is the molarity or concentration, n is the number of moles and V is the volume of the solution. Now {2} holds true only for stoichiometric relations of 1:1, if there is a different ratio then we must adjust the equation as below.



$$n_{\text{-OH}} = M_{\text{-OH}}V_{\text{-OH}}, \quad M_{\text{+H}} = \frac{\frac{1}{2}n_{\text{-OH}}}{V_{\text{+H}}}, \{4\}$$

From {3} there are half the number of moles of sulphuric acid as sodium hydroxide present and so we must apply this to the molarity calculation for the acid molarity in {4}. Likewise, if we were titrating an unknown solution of NaOH in {3} then we would have to double the moles of acid to get the hydroxide concentration. This example highlights the need for care to be taken when evaluating the results of a titration, as depending on the chemicals in the experiment and the order in which one is being added to the other an error is easily made.

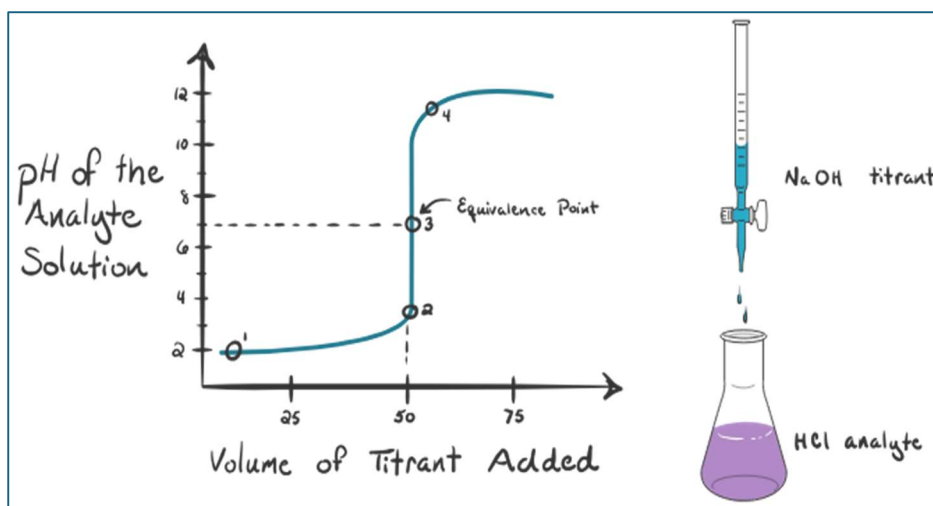


Figure 1:

Titration curve for NaOH and HCl (left), and representation of experimental setup showing phenolphthalein indicator changing colour (right).

Credit:
Kahn Academy [2].

Materials and Methods

Chemicals & Equipment:

- 80 mL of standard NaOH solution, $M_B = 0.1026 \text{ mol/L}$ in a labelled beaker
- 80 mL HCl solution with an unknown concentration
- Phenolphthalein indicator
- 1x Burette, 50 mL attached to a burette stand with clamp
- 1x Pipette, 10 mL, with pipette filler
- 3x Erlenmeyer flasks, 100 mL
- 1x Small funnel to fill the burette
- 1x Plastic pipette
- 1x Wash bottle with deionised water
- 1x Graduated cylinder, 10 mL
- 1x White tile
- 1x Beaker for the waste

Initially all equipment was washed with deionised water with the burette and pipette also rinsed with a small amount of NaOH and HCl respectively. Using the funnel the burette was filled with the known standard solution of NaOH, running some NaOH through the burette gently tapping it to ensure the equipment is free from air bubbles. The initial reading of the burette was taken at this point. 25mL of the HCl was placed in a clean Erlenmeyer flask with a few drops of phenolphthalein indicator. First a rough titration was done to gain a ballpark figure for subsequent titrations, the burette tap was turned on a low stream and the flask stirred until the indicator turned purple. The burette reading was recorded as the 'rough' run of the experiment and the volume of NaOH calculated. The previous steps were repeated with another sample of HCl, with the burette allowed to run stopping a few mL short of the rough volume. Subsequently great care was taken slowly dripping and stirring in the NaOH stopping as soon as the solution changed colour to pink, with the burette reading taken and the volume calculated once more. This was then repeated until there were two volumes of NaOH that were within 0.10mL of each other.

Results and Discussion

	Titration 1 (rough)	Titration 2	Titration 3
Volume HCl, V_A (mL)	25.00	25.00	25.00
Initial Burette Vol, (mL)	0.62	7.70	27.00
Final Burette Vol, (mL)	19.60	27.00	46.40
Δ Volume, V_B (mL)	18.98	19.30	19.40
Moles of NaOH, n_B (mol $\times 10^{-3}$)	1.95	1.98	1.99
Moles of HCl, n_A (mol $\times 10^{-3}$)	1.95	1.98	1.99
Molarity of HCl, M, ([mol/L] $\times 10^{-3}$)	77.8	79.2	79.6

Table 1: Experimental results obtained during titration. Number of moles and molarity of HCl calculated by rearranging equation {2} and substituting in known values shown equation {5} below.

From titration 2:

$$V_B \times M_B = n_B = n_A, \{5\}$$

$$19.3 \times 10^{-3} \times 0.1026 = 1.98 \text{ mmol}, \{6\}$$

Therefore from {2} we find the Molarity of our HCl.

$$M_A = \frac{n}{V} = \frac{1.98 \text{ mmol}}{25 \text{ mL}} = 79.2 \text{ mM} \{7\}$$

The same procedure outlined above was repeated for all titration values in Table 1. The average molarity of titrations 2 and 3 was 79.4 mmol/L. Given how close the values of the titration are in an albeit limited dataset, the author can be confident that I accurate approximation for the molarity of the hydrochloric acid has been determined. Naturally sources of error could have slips in during the experimental process these namely being contamination and parallax error, however following procedures outlined in the method above it is deemed unlikely that these have played significant role. Ideally a greater amount of titrations could be taken in future to mitigate potentially erroneous individual data points. Titrations all I powerful tool in any chemist or engineers belt when it comes to quantifying unknown quantities and can be an OME context by determining the concentration of home made energetic compounds or monitoring the concentration of decomposition precursors with materials in storage.

Conclusion

The titration determined the concentration of hydrochloric acid to be approximately 79.4mM. Without more data or alternative methods to independently confirm this result, it cannot be asserted that's the given concentration is in fact accurate. However, given the values came out within a reasonably the precise range it can be argued that we can have some confidence in our determined concentration.

Bibliography

[1] Yu Kay Law, "Acid-Base Titration Curves," *Pressbooks.pub*, 2020.

<https://iu.pressbooks.pub/virtualchemlabs/chapter/chapter-1/> (accessed Mar. 20, 2024).

[2] "Khan Academy," *Khanacademy.org*, 2023. <https://www.khanacademy.org/test-prep/mcat/chemical-processes/titrations-and-solubility-equilibria/a/acid-base-titration-curves> (accessed Mar. 17, 2024).