

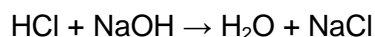
EXPERIMENT 1

ACID BASE TITRATION

Objective : To determine pH curve for titration of strong acid-strong base and weak acid-strong base.

1 Theory

The process of adding acid to a base (or vice versa) to produce a salt and water is called **neutralization**. In the neutralization of hydrochloric acid with sodium hydroxide, the reaction that occurs is:



When an acid and a base are present in stoichiometric amount, for example one mole of hydrochloric acid is added to one mole of sodium hydroxide as in the above neutralization process, this means that the equivalent point has been reached in the acid-base system. Suitable indicators (refer to table 1) can be used to monitor the end-point (the point at which an indicator changes colour) of the titration.

Indicator	Colour in Acid	End-point Colour	Colour in Alkali	pK _{ind}	Titration to Use For
Litmus	RED	PURPLE	BLUE	6.5	Strong Acid - Strong Base
Methyl Orange	RED	ORANGE	YELLOW	3.1-4.4	Strong Acid - Strong Base or Strong acid - Weak Base
Phenolphthalein	COLOURLESS	PALER PINK	PINK	9.3	Weak acid and Strong Base
Bromothymol Blue	YELLOW	GREEN	BLUE	6.2-7.6	Strong acid - Strong base

Table 1: Common Indicators and Their Colours

In the above example, pH of the hydrochloric acid will change with increasing amount of sodium hydroxide added. At the end-point the solution will become neutral with pH of 7. This is only true for strong acid-strong base titration. For strong acid-weak base, weak acid-strong base and weak acid-weak base titration, the salt that is formed may undergo hydrolysis and this may cause the pH not to be equal to 7 at the end-point. Graph of pH versus volume of base that is added to the acid of constant volume or otherwise is called the **pH titration curve**.

2 Materials

Chemicals: Sodium carbonate, hydrochloric acid, sodium hydroxide and acetic acid.

Indicators: Methyl orange, phenolphthalein and bromothymol blue.

3 Experimental Procedure

3.1 Standardization of acidic and basic solution

3.1.1 Preparation of Standard Sodium Carbonate Solution

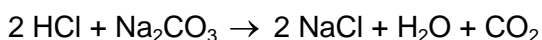
- (i) Weigh accurately about 1.3 g of the primary standard, anhydrous sodium carbonate in a weighing bottle.
- (ii) Transfer the materials quantitatively into a 250 ml volumetric flask; Then add about 80 mL distilled water and shake the flask until all the sodium carbonate dissolve. Add distilled water to the volumetric mark and then shake the flask well.
- (iii) Calculate the concentration of the standard sodium carbonate solution in unit of mol L⁻¹. (Report the concentration to four significant figures from the decimal point).

3.1.2 Standardization of hydrochloric acid

- (i) Pipette 25 mL of standard sodium carbonate solution into a 250 mL conical flask, add 2 drops of methyl orange.
- (ii) Titrate with the given hydrochloric acid until the solution starts becoming red.
- (iii) Repeat the titration, this time immediately add the acid until it is short of the titer value obtained in (ii) by 0.5 mL. Then titrate slowly until the end-point is reached.
- (iv) Repeat the titration until the titer value does not differ by more than 0.05 mL from the previous titration.
- (v) Calculate the concentration of the hydrochloric acid given. (Report your result to four significant figures from the decimal point).

Method of calculating the concentration of hydrochloric acid solution

Consider the concentration of the standard sodium carbonate solution is c mol dm⁻³, and v mL of hydrochloric acid is required to neutralize 25 mL of this standard sodium carbonate solution. The reaction that occurs is:



From the above reaction it is clear that the neutralization of 1 mole of sodium carbonate requires 2 moles of hydrochloric acid.

25 mL of the carbonate solution used contains $\frac{25c}{1000}$ moles of sodium carbonate.

Therefore the neutralization requires $2 \times \frac{25c}{1000}$ moles hydrochloric acid. Hence, the concentration of the hydrochloric acid solution is

$$\frac{2 \times 25c}{1000} \times \frac{1000}{v} = 2 \times \frac{25c}{v} \text{ mol dm}^{-3} \text{ (or M)}$$

3.1.3 Standardization of sodium hydroxide solution

- (i) Pipette 10 mL of hydrochloric acid that has been standardized into a 250 mL conical flask.

- (ii) Add 2 drops of methyl orange.
- (iii) Titrate with the given sodium hydroxide solution.
- (iv) Repeat the titration; this time simply add the sodium hydroxide until it 0.5 mL less than the titer value obtained in step (ii). Then titrate slowly until the end-point is reached.
- (v) Repeat the titration until the titer value does not differ by more than 0.05 mL.
- (vi) Repeat steps (i) to (iv) using phenolphthalein as the indicator.
- (vii) Calculate the concentration of the sodium hydroxide solution obtained from the titrations using methyl orange and phenolphthalein as indicators.

3.1.4 Standardization of acetic acid solution

- (i) Pipette 10 mL of acetic acid into a 250 mL conical flask and add 2 drops of phenolphthalein indicator.
- (ii) Titrate with sodium hydroxide solution that has been previously standardized.
- (iii) Repeat the titration; this time the sodium hydroxide is quickly added until it is 0.5 mL short of the titer value obtained in step (ii), then titrate slowly until the end-point is reached.
- (iv) Repeat the titration until the titer value obtained does not differ by more than 0.05 mL.
- (v) Repeat steps (i) to (iv) but using bromothymol blue as the indicator.
- (vi) Calculate the concentration of acetic acid from the titration using phenolphthalein and bromothymol blue as the indicators.

3.2 pH curve for strong acid-strong base titration

- (i) Prepare a solution of sodium hydroxide that is diluted 10 times by pipetting 25 mL of sodium hydroxide into a 250 mL volumetric flask and adding distilled water to the mark.
- (ii) Prepare two burettes, one is filled with the **standard** sodium hydroxide solution and the other is filled with sodium hydroxide that has been diluted 10 times.

(You are required to share your burette with your neighbour in this part)

In order to fill the burette with diluted sodium hydroxide, empty the standard sodium hydroxide from the burette and rinse the burette with distilled water several times. Finally, rinse the burette with 3 parts (each part 10 mL) of dilute sodium hydroxide, then the burette is filled with the diluted sodium hydroxide solution.

- (iii) Pipette 10 mL of standard hydrochloric acid into a 50 mL beaker.
- (iv) Titrate this acid with standard sodium hydroxide until the pH is 2.8 or 2.9 (you may use a calibrated pH meter to determine the pH of the solution). Record the volume of the required standard sodium hydroxide solution.
- (v) Continue the titration with dilute sodium hydroxide in order to get pH 4, 5, 6, 7, 8, 9 and 10.
- (vi) Continue the titration with standard sodium hydroxide in order to get pH 11 and 12.
- (vii) Plot a graph of pH of solution against the volume of standard sodium hydroxide used.

(Attention: 10 mL of dilute sodium hydroxide equals 1 mL of standard sodium hydroxide).

3.3 pH curve for weak acid-strong base titration

- (i) Pipette 10 mL of standard acetic acid into a 100 mL conical flask.
- (ii) Titrate the acetic acid with standard sodium hydroxide in order to get pH 3, 4, 5 and 6.
- (iii) Continue the titration with dilute sodium hydroxide in order to get pH 7, 8, 9 and 10.
- (iv) Continue the titration with standard sodium hydroxide in order to get pH 11 and 12.
- (v) Plot a graph of pH of solution against the volume of standard sodium hydroxide used.

4 Questions

1. Calculate the theoretical pH of the solution at the following instances when 0.00, 5.00, 9.00, 9.50, 9.80, 9.92, 10.00, 10.02, 10.20 and 15.00 mL of standard sodium hydroxide solution are added to 10 mL hydrochloric acid. Plot the pH curve obtained by theoretical calculations together with the pH curve obtained in **Part 3.2** of this experiment. Compare the two pH curves and give your comments on the differences that are observed.
2. Calculate the theoretical pH of the solution at the following instances when 0.00, 5.00, 9.00, 9.50, 9.80, 9.92, 10.00, 10.02, 10.20 and 15.00 mL of standard sodium hydroxide solution are added to 10 mL acetic acid. Assume K_a (acetic acid) to be 1×10^{-5} . Plot the pH curve obtained by theoretical calculation together with the pH curve obtained in **Part 3.3** of this experiment. Compare the two pH curves and give your comments on the differences that are observed.
3. The colour changes corresponding to the pH intervals for the three indicators are given in the following table:-

Indicators	Colour change	pH Interval
Methyl orange	Red to orange	3 - 5
Bromothymol blue	Yellow to blue	6 - 8
Phenolphthalein	Colourless to pink	8 - 10

Comment on the end-point reached by using different indicators in **Part 3.1.3** and **Part 3.1.4**. Confirm the suitability of each indicator by referring to the pH curves determined in this experiment.

5 Reference

1. Mahan, B.H. (1987). *University Chemistry*, 4th edn. Addison-Wesley.
2. <http://www.usetute.com.au/indicata.html>
3. Vogel, Athur T. (1994). *Vogel's Textbook of Quantitative Chemical Analysis*. Longman.

nb: Please quote experimental error estimates for all your data presented.

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Some of these experiments are carried out in groups of usually a pair of students. Therefore expectedly, each member of a group followed an identical procedure in the laboratory and has the same set of raw data. Members of a group are allowed to discuss the analysis of data with one another. However, preparation of the report including data analysis, interpretation and discussion must be prepared by the individual student submitting the report.

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