

# Classic Titration

## pH Curves and an Unknown

### Introduction

One of the most common questions chemists have to answer is how much of something is present in a sample or a product. If the product contains an acid or base, these questions are usually answered by titration. Acid–base titrations can be used to measure the concentration of an acid or base in solution and to calculate the formula (molar) mass of an unknown acid or base.

### Concepts

- Acids and bases
- Neutralization reactions
- Titration curve
- Titration
- Equivalence point

### Background

Titration is a method of *calorimetric analysis*—the use of volume measurements to analyze an unknown. In acid–base chemistry, titration is most often used to analyze the amount of acid or base in a sample or solution. Consider a solution containing an unknown amount of hydrochloric acid. In a titration experiment, a known volume of the hydrochloric acid solution would be “titrated” by slowly adding (dropwise) a standard solution of a strong base, such as sodium hydroxide. (A standard solution is one with a concentration that is accurately known.) The titrant, sodium hydroxide (in this case), reacts with and consumes the acid via a neutralization reaction (Equation 1). The exact volume of base needed to react completely with the acid is measured. This is called the equivalence point of the titration—the point at which stoichiometric amounts of the acid and base have combined.



Knowing the exact concentration and volume added of the titrant gives the number of moles of sodium hydroxide. The latter, in turn, is related by stoichiometry to the number of moles of hydrochloric acid initially present in the unknown.

Indicators are usually added to acid–base titrations to detect the equivalence point. The end point of the titration is the point at which the indicator changes color and signals that the equivalence point has indeed been reached. For example, in the case of the neutralization reaction shown in Reaction 1, the pH of the solution would be acidic (< 7) before the equivalence point and basic (> 7) after the equivalence point. The pH at the equivalence point should be exactly 7, corresponding to the neutral products (sodium chloride and water). An indicator that changes color around pH 7 is therefore a suitable indicator for the titration of a strong acid with a strong base.

The progress of an acid–base titration can also be followed by measuring the pH of the solution being analyzed as a function of the volume of titrant added. A plot of the resulting data is called a pH curve or titration curve. Titration curves allow a precise determination of the equivalence point of the titration without the use of an indicator.

### Experiment Overview

The purpose of this experiment is to analyze the shape of the titration curve for neutralization of hydrochloric acid with sodium hydroxide and to determine the concentration of an unknown hydrochloric acid solution. In Part A, pH measurements may be made electronically as a function of time using a pH sensor attached to a computer- or calculator-based interface system, or manually as a function of volume added using a pH meter. In Part B, the titration will be repeated with the use of an indicator to assess the endpoint.

### Pre-Lab Questions

The approximate concentration of a hydrochloric acid solution is 0.5 M. The exact concentration of this solution is to be determined by titration with 0.215 M sodium hydroxide solution.

1. A 10.00-mL sample of the HCl solution was transferred by pipet to an Erlenmeyer flask and then diluted by adding about 40 mL of distilled water. What is the approximate  $\text{H}_3\text{O}^+$  concentration and pH of the solution in the flask before the titration begins?
2. Thymolphthalein indicator was added, and the solution in the flask was titrated with 0.215 M NaOH until the indicator just turned pink (pH = 8–9). The exact volume of NaOH required was 22.75 mL. Use the following equation to calculate the concentration of HCl in the original 10.00-mL sample.

$$M_1 \times V_1 = nM_2 \times V_2$$

- $M_1$  = molarity of standard base solution
- $M_2$  = unknown molarity of acid solution
- $V_1$  = volume of base added
- $V_2$  = initial volume of acid solution
- $n$  = mole ratio (number of moles of base that react with one mole of acid)

3. One student accidentally “overshot” the endpoint and added 23.50 mL of 0.215 M NaOH. Is the calculated concentration of HCl likely to be too high or too low as a result of this error?

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