

1.11.8: Acid–Base Titration

Note from Dr. B.

We will be looking at acids and bases more in Module 7, but a titration is a quantitative way of looking at solution concentrations of acids and bases, so we can look at it in Module 6.

Don't worry about the very end of the page (Indicator Selection for Titrations) because we haven't covered pH yet.

Learning Objectives

- Understand the basics of acid-base titrations.
- Understand the use of indicators.
- Perform a titration calculation correctly.

The reaction of an acid with a base to make a salt and water is a common reaction in the laboratory, partly because so many compounds can act as acids or bases. Another reason that acid-base reactions are so prevalent is because they are often used to determine quantitative amounts of one or the other. Performing chemical reactions quantitatively to determine the exact amount of a reagent is called a **titration**. A titration can be performed with almost any chemical reaction for which the balanced chemical equation is known. Here, we will consider titrations that involve acid-base reactions.

During an acid-base titration, an acid with a known concentration (a **standard solution**) is slowly added to a base with an unknown concentration (or vice versa). A few drops of indicator solution are added to the base. The indicator will signal, by color change, when the base has been neutralized (when $[H^+] = [OH^-]$). At that point—called the **equivalence point**, or **end point**—the titration is stopped. By knowing the volumes of acid and base used, and the concentration of the standard solution, calculations allow us to determine the concentration of the other solution.

It is important to accurately measure volumes when doing titrations. The instrument you would use is called a **burette** (or buret).

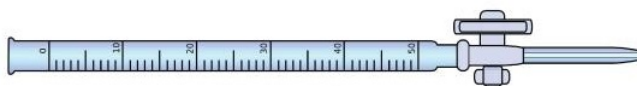
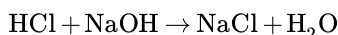


Figure 1.11.8.1: Equipment for Titrations. A burette is a type of liquid dispensing system that can accurately indicate the volume of liquid dispensed.

For example, suppose 25.66 mL (or 0.02566 L) of 0.1078 M HCl was used to titrate an unknown sample of NaOH. What mass of NaOH was in the sample? We can calculate the number of moles of HCl reacted:

$$\# \text{ mol HCl} = (0.02566 \text{ L})(0.1078 \text{ M}) = 0.002766 \text{ mol HCl}$$

We also have the balanced chemical reaction between HCl and NaOH:



So we can construct a conversion factor to convert to number of moles of NaOH reacted:

$$0.002766 \text{ mol } \cancel{\text{HCl}} \times \frac{1 \text{ mol NaOH}}{1 \text{ mol } \cancel{\text{HCl}}} = 0.002766 \text{ mol NaOH}$$

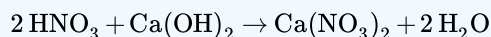
Then we convert this amount to mass, using the molar mass of NaOH (40.00 g/mol):

$$0.002766 \text{ mol } \cancel{\text{HCl}} \times \frac{40.00 \text{ g NaOH}}{1 \text{ mol } \cancel{\text{HCl}}} = 0.1106 \text{ g NaOH}$$

This type of calculation is performed as part of a titration.

✓ Example 1.11.8.1: Equivalence Point

What mass of Ca(OH)_2 is present in a sample if it is titrated to its equivalence point with 44.02 mL of 0.0885 M HNO_3 ? The balanced chemical equation is as follows:



Solution

In liters, the volume is 0.04402 L. We calculate the number of moles of titrant:

$$\# \text{ moles HNO}_3 = (0.04402 \text{ L})(0.0885 \text{ M}) = 0.00390 \text{ mol HNO}_3$$

Using the balanced chemical equation, we can determine the number of moles of Ca(OH)_2 present in the analyte:

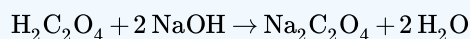
$$0.00390 \text{ mol } \cancel{\text{HNO}_3} \times \frac{1 \text{ mol } \text{Ca(OH)}_2}{2 \text{ mol } \cancel{\text{HNO}_3}} = 0.00195 \text{ mol } \text{Ca(OH)}_2$$

Then we convert this to a mass using the molar mass of Ca(OH)_2 :

$$0.00195 \text{ mol } \cancel{\text{Ca(OH)}_2} \times \frac{74.1 \text{ g } \text{Ca(OH)}_2}{\cancel{\text{mol } \text{Ca(OH)}_2}} = 0.144 \text{ g } \text{Ca(OH)}_2$$

? Exercise 1.11.8.1

What mass of $\text{H}_2\text{C}_2\text{O}_4$ is present in a sample if it is titrated to its equivalence point with 18.09 mL of 0.2235 M NaOH ? The balanced chemical reaction is as follows:



Answer

0.182 g

? Exercise 1.11.8.2

If 25.00 mL of HCl solution with a concentration of 0.1234 M is neutralized by 23.45 mL of NaOH , what is the concentration of the base?

Answer

0.1316 M NaOH

? Exercise 1.11.8.3

A 20.0 mL solution of strontium hydroxide, Sr(OH)_2 , is placed in a flask and a drop of indicator is added. The solution turns color after 25.0 mL of a standard 0.0500 M HCl solution is added. What was the original concentration of the Sr(OH)_2 solution?

Answer

$3.12 \times 10^{-2} \text{ M } \text{Sr(OH)}_2$

Indicator Selection for Titrations

The indicator used depends on the type of titration performed. The indicator of choice should change color when enough of one substance (acid or base) has been added to exactly use up the other substance. Only when a strong acid and a strong base are produced will the resulting solution be neutral. The three main types of acid-base titrations, and suggested indicators, are:

The three main types of acid-base titrations, suggested indicators, and explanations

Titration between . . .	Indicator	Explanation
strong acid and strong base	any	
strong acid and weak base	methyl orange	changes color in the acidic range (3.2 - 4.4)
weak acid and strong base	phenolphthalein	changes color in the basic range (8.2 - 10.6)

Summary

A titration is the quantitative reaction of an acid and a base. Indicators are used to show that all the analyte has reacted with the titrant.

Contributions & Attributions

- [Peggy Lawson](#) (Oxbow Prairie Heights School). Funded by Saskatchewan Educational Technology Consortium.

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