10.3 Acid–Base Stoichiometry

An important application of neutralization reactions is finding the concentration of an unknown solution. The dairy industry, for example, uses a neutralization reaction to monitor the concentration of lactic acid in stored milk. Fresh milk has no lactic acid. Bacteria in milk consume lactose, a milk sugar, and produce lactic acid as a waste product. The older the milk, the more lactic acid is present. Finding the concentration of an unknown solute like lactic acid in milk involves an analysis called titration.

CAREER LINK

Food scientists are concerned about the freshness of the ingredients used in food products. To learn more about the work of a food scientist,

GO TO NELSON SCIENCE

titration a procedure used to determine the concentration of a solution using a standardized solution

titrant the solution in the burette during a titration

burette a calibrated tube used to deliver variable known volumes of a liquid during a titration

Titration

A **titration** is an analytical procedure used to determine the concentration of a solution. During a titration, a measured volume of a standardized solution called the **titrant** is gradually added to a flask containing a measured volume of a solution of unknown concentration. Recall that a standardized solution is a solution whose concentration is precisely known. A burette is used to accurately measure the volume of titrant used. A **burette** is a tube that is typically calibrated with 0.1 mL divisions. The valve at one end of the burette is used to control the volume of titrant that leaves the burette. **Figure 1** shows the titration of sulfuric acid (in the flask) with sodium hydroxide (in the burette), using bromothymol blue indicator. The chemical equation for this neutralization reaction is

$$2\ NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2\ H_2O(l)$$

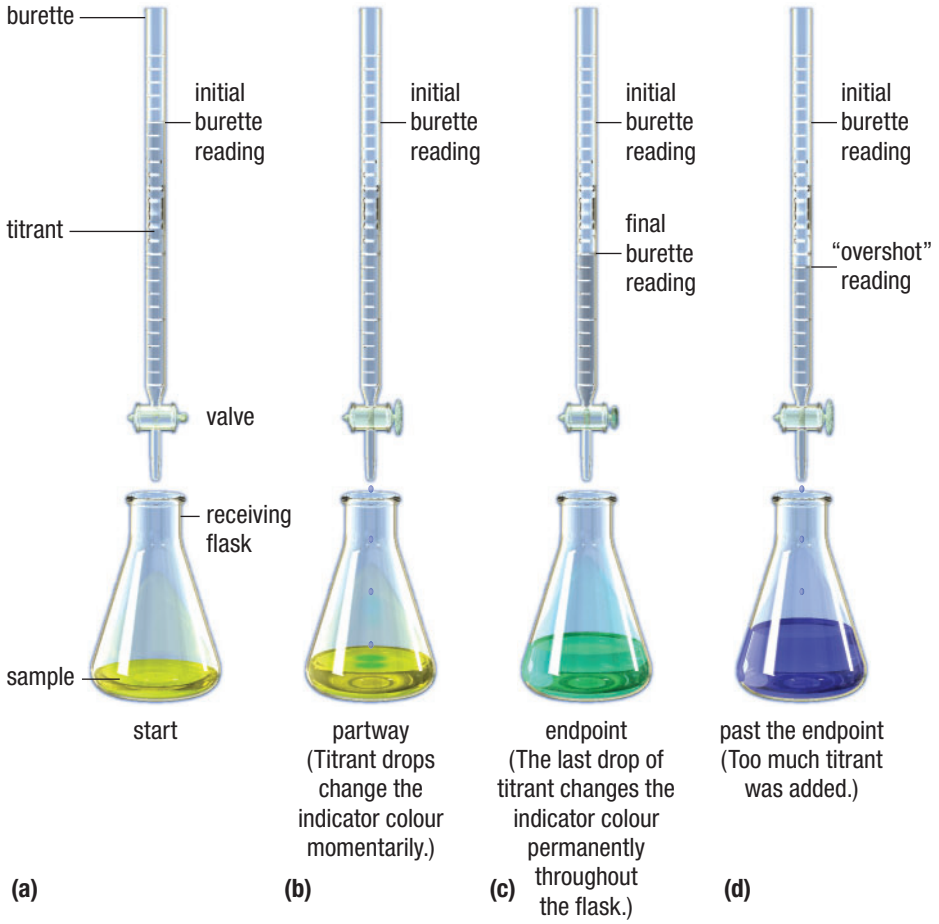

Figure 1 The titration of sulfuric acid with sodium hydroxide (a) A few drops of bromothymol blue indicator make the sulfuric acid solution initially yellow. (b) Adding sodium hydroxide produces a temporary colour change that fades as the flask is swirled. (c) The last drop of titrant changes bromothymol blue permanently to green at the endpoint. (d) The blue colour indicates that the titration has gone past the desired endpoint. This means too much base has been added.

WEB LINK

There are many online videos showing how to perform a titration and how to identify when the endpoint has been reached. To view some of these videos,

GO TO NELSON SCIENCE

The point during an acid–base titration when neutralization is complete is called the **equivalence point**. At this point, the amount of acid and base added to the flask exactly match their mole ratio in the chemical equation. They are stoichiometrically equal. To identify when the equivalence point is reached, chemists look for a sudden change in an observable property of the solution. This could be a sudden change in pH (detected using a pH meter) or the colour of an acid–base indicator. The point in a titration at which this sudden observable change occurs is called the **endpoint**. The endpoint and equivalence point are not the same (**Table 1**). In most cases, we chose an acid–base indicator that changes colour at the equivalence point of the reaction. The colour change is a visual clue that the equivalence point has been reached. The indicator changes colour permanently when a very slight excess of titrant is added. For example, the endpoint of the titration in Figure 1 occurs when bromothymol blue changes permanently to green.

You might wonder whether titrations are done in "real life." Indeed they are! For example, companies that process pickles have to be sure that the liquid in the pickle jar is acidic enough to preserve the pickles. Usually they use an electronic pH meter to monitor the acidity, but they also periodically check the accuracy of the pH meter by performing a titration.

When titrating, it is important to make accurate and precise observations. You need to ensure that you accurately determine the volume of titrant required to reach the equivalence point. To improve the quality of your data, you should repeat the procedure several times, determining the volume used each time. Three trials with results that are within 0.2 mL are normally required for a reliable analysis.

Although the equivalence point and the endpoint are not the same, they coincide if the correct indicator is selected. Thus, the endpoint tells us that the equivalence point has been reached.

equivalence point the point in a titration when neutralization is complete

endpoint the point during a titration when a sudden change in an observable property of the solution occurs; usually a change in colour of an acid–base indicator or a significant change in pH

Table 1 Comparing Equivalence Points and Endpoints

Endpoint	Equivalence point
sudden observable change (e.g., in colour, pH, or conductivity)	theoretical: cannot be observed
occurs when an acid–base indicator reacts with the titrant to produce a lasting colour change	occurs when an acid and a base have completely neutralized each other
depends on the type of indicator used	depends on the stoichiometry of the neutralization reaction

Choosing an Acid–Base Indicator

A titration gives accurate results only when its endpoint coincides with its equivalence point. This means that the acid–base indicator chosen for the titration must change colour when the equivalence point is reached. How do you know which is an appropriate indicator for a titration? To answer this question, we can first examine the pH changes that occur during the neutralization of a strong acid with a strong base (**Figure 2**). Imagine that a pH meter is recording the pH of the solution in the flask throughout the titration.

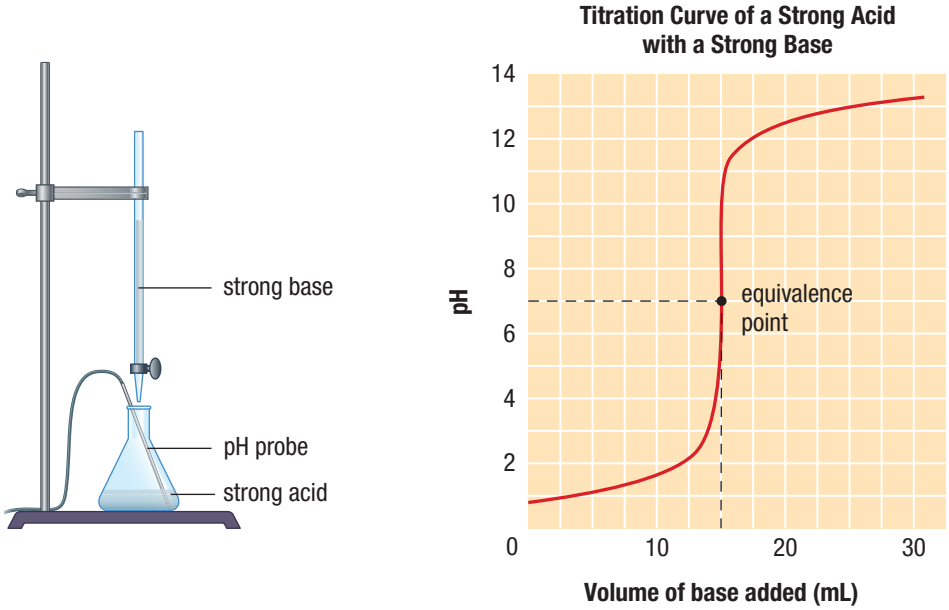

Figure 2 When a strong acid is titrated with a strong base, the pH increases quickly near the equivalence point. In this titration, the equivalence point is at pH 7.

The pH of the solution starts off very low, when there is only acid in the flask. When the base is dripped into the flask the pH increases, slowly at first, as the base neutralizes the acid. Then a rapid increase in pH occurs as the base neutralizes the last traces of the acid. This occurs during the steep portion of the graph that includes the equivalence point.

Acid–base indicators do not change colour at a specific pH value. Rather, we can observe the colour changes over a small pH range. **Table 2** lists some common acid–base indicators and pH ranges over which they change colour. The ideal acid–base indicator should change colour during the steep portion of the titration graph.

Table 2 Common Acid–Base Indicators and the pH Range of Their Colour Changes

Indicator	Acid colour	Base colour	pH range of colour change
thymol blue	red	yellow	1.2 to 2.8
	yellow	blue	8.0 to 9.6
methyl red	red	yellow	4.4 to 6.2
bromothymol blue	yellow	blue	6.0 to 7.6
phenolphthalein	colourless	pink	8.0 to 9.6
alizarin yellow	yellow	red	10.1 to 12.0

LEARNING **TIP**

Indicator Solvents

Some indicators are compounds dissolved in ethanol, which makes them flammable. They should be handled with care and kept away from flames and sparks.

Which of these indicators would be suitable to mark the equivalence point of the titration represented in Figure 2? Methyl red, bromothymol blue, and phenolphthalein are all suitable because their colour changes occur during the steep portion of the curve that includes the equivalence point (**Figure 3**). Thymol blue and alizarin yellow are not suitable because their colour changes do not occur in the steep portion of the graph.

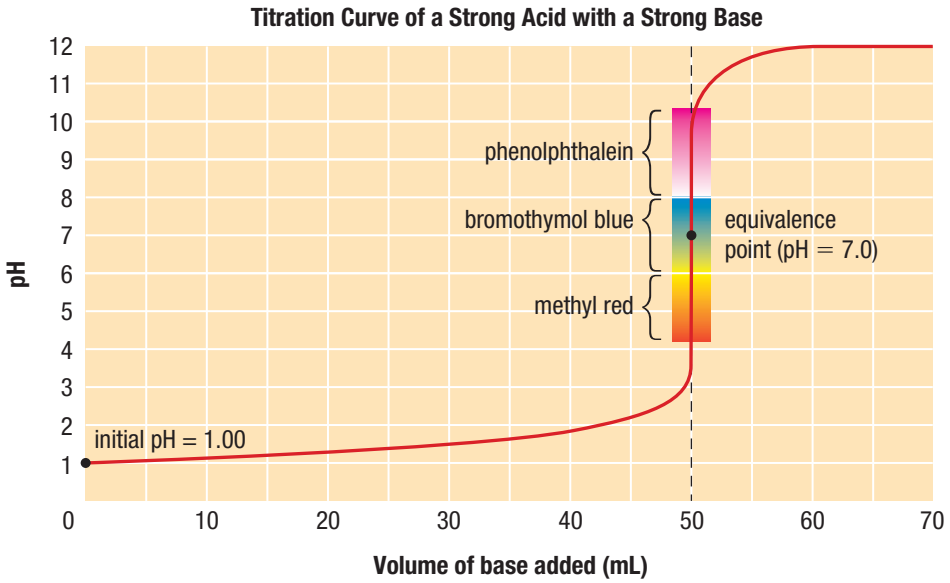

Figure 3 Methyl red, bromothymol blue, and phenolphthalein are appropriate indicators for the titration of a strong acid with a strong base because their colour changes occur along the steepest portion of the pH curve.

UNIT TASK **BOOKMARK**

In the Unit Task (page 498) you will have a chance to apply the skills that you learn in this section and in the investigations. The task involves a titration and a series of stoichiometric calculations. Your aim is to test a water-softening process that you have designed.

Tutorial 1 Performing a Titration

Titration is an analytical procedure used to determine the concentration of a solution. Titration involves using a burette to add a precise volume of one solution (the titrant) to a sample of another solution in an Erlenmeyer flask. Titrant is added until an observable endpoint is reached. The volume of titrant used is then determined by subtracting initial volume from final volume.

The following steps outline how to perform a titration.

Preparing the Burette

SKILLS HANDBOOK A3. 5

1. Assemble the required equipment: burette; burette brush; retort stand and burette clamp; small funnel; beaker; Erlenmeyer flask; meniscus reader; wash bottle of distilled water (**Figure 4**). You will also need the titrant solution and the sample solution. Check that the tip of the burette is not chipped or broken. Replace the tip or the entire burette if necessary. Also check that the valve turns easily and smoothly and that liquid flows easily out of the tip.
2. Close the valve. Rinse the inside of the burette with small volumes of distilled water from the wash bottle. Open the valve to allow the water to drain through the tip (**Figure 5**). If water clings to the inner wall of the burette, carefully scrub the inside of the burette with the brush and rinse again.
3. Close the valve. Clamp the burette to the stand. Pour about 10 mL of the titrant solution into the burette using the funnel (**Figure 6**). You may have to raise the funnel slightly to allow the solution to flow into the burette.
4. Carefully remove the burette from its clamp. Tilt and roll the burette in your fingers so that its inner surface is coated with titrant (**Figure 7**). Allow some titrant to flow out through the tip into a waste beaker. Pour the remaining titrant out of the top of the burette into the beaker. Replace the burette in its clamp and repeat Steps 3 and 4.

Figure 7 Roll the burette in your fingers to wet the inner surface with titrant.

5. Using the funnel, fill the burette with titrant until the volume level is about 2 mL higher than the 0.00 mL mark.
6. Place the waste beaker under the burette. Briefly open the valve fully to clear any air bubbles that may have formed in the valve and tip. Then almost close the valve (**Figure 8**, on next page) and allow liquid to run slowly out of the burette until the level is between 0.00 mL and 1.00 mL. Record this reading to the nearest 0.02 mL. Always read a liquid level at the bottom of the meniscus. Read at eye level rather than on an angle (**Figure 9**, on next page). You might find it helpful to use a meniscus reader: a card with a black stripe that makes it easier to see the meniscus.
7. Touch the tip of the burette to the side of the waste beaker, if necessary, to remove any clinging drops. Use the wash bottle of distilled water to rinse the outside of the tip, to ensure that there are no drops of titrant on the outside that might affect your readings.

Figure 4 Titration equipment

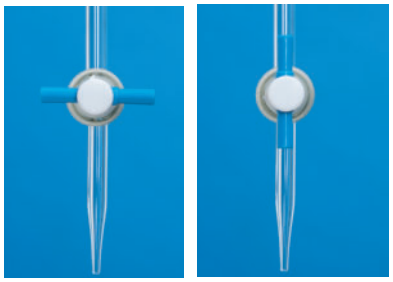

Figure 5 The valve is closed when it is perpendicular to the burette and open when it is parallel to the burette.

Figure 6 A tight-fitting funnel prevents the titrant from draining into the flask. Raise the funnel slightly if necessary.

LEARNING TIP

Titration Precautions

When you are preparing a burette for a titration, follow all the usual safety precautions for handling toxic or corrosive solutions. In addition, take particular care with laboratory glassware. Always wear chemical safety goggles, an apron, and protective gloves.

Figure 8 Air bubbles in the tip and valve must be cleared before starting a titration.

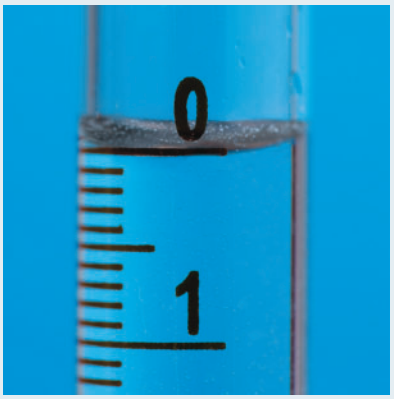

Figure 9 Always take your reading from the bottom of the meniscus.

Figure 10 The burette tip should remain in the flask throughout the titration.

Preparing the Sample

SKILLS HANDBOOK A3. 4

8. You will need a beaker containing the sample solution and a pipette fitted with a pipette pump. Pipette the required volume of sample solution into the Erlenmeyer flask. (Instructions for pipetting are given in the Skills Handbook, A3.4.)
9. Add an appropriate indicator to the flask, if necessary. Use the smallest possible volume of indicator. Since many indicators are weak acids, using too much of the indicator solution will give an inaccurate endpoint.
10. Place a sheet of white paper under the flask to make colour changes easier to see.

Titrating an Unknown Solution

11. Record the initial burette volume.
12. Raise the titration flask or lower the burette so that the burette tip is in the mouth of the flask (**Figure 10**). This will help prevent spillage of titrant.
13. Add titrant to the flask, quickly at first (**Figure 11**). Then, as you near the endpoint (when you see the first signs of a change in colour), add titrant drop by drop, swirling the flask gently between drops.
14. Use a wash bottle to rinse the sides of the flask and the tip of the burette to ensure that all the titrant has reacted (**Figure 12**).
15. Stop titrating when the endpoint is reached: when the addition of a drop of titrant produces a permanent colour change. If you are unsure whether you have reached the endpoint, record the burette volume. Then add another drop of titrant. If a dramatic colour change occurs, you were likely at the endpoint and have now passed it.

LEARNING **TIP**

Cleaning Equipment

All laboratory equipment should be cleaned before being stored. To clean burettes and pipettes, first drain them, then rinse them with acetic acid (vinegar), and then with distilled water.

Figure 11 Allow the titrant to flow into the flask. Prepare to reduce the flow as you approach the endpoint.

Figure 12 Rinse the tip and sides of the titration flask with distilled water prior to the endpoint.

Questions

1. Why is it necessary to keep the volume of indicator used in an acid–base titration to a minimum? K/U
2. Why do you think it is necessary to rinse a burette with titrant after it is rinsed with water? T/I
3. A student forgot to clear an air bubble prior to starting a titration. What effect would this have on the measured volume of titration required to reach the endpoint? T/I
4. Why do you think rinsing the sides of a titration flask near the end of the titration is recommended? T/I
5. A student titrates a hydrochloric acid solution (in the flask) with a sodium hydroxide solution (from the burette) to a phenolphthalein endpoint. The solution in the flask turns pink momentarily when a small volume of sodium hydroxide is added (**Figure 13**). The pink then disappears as the flask is swirled. Explain why these colour changes are observed. T/I

Figure 13 The endpoint has been reached when a single drop of titrant results in a colour change that does not fade when the mixture is swirled.

Investigation **10.3.1**

Standardization of a Sodium Hydroxide Solution (p. 487)
In this investigation you will learn and practise the techniques of titration as you determine the concentration of a basic solution.

Standardizing Solutions of Acids and Bases

Titrations require standard solutions. Recall, from Section 8.6, that we know the precise concentration of a standard solution. Standard solutions are sometimes prepared using a precisely measured mass of a solid solute or volume of a concentrated solution. In practice, however, this is not always possible. Some solids are difficult to obtain in a highly pure form. Others change once they are exposed to air. Sodium hydroxide, for example, absorbs water vapour from the air. As a result, the mass of solid sodium hydroxide increases slightly as it is being weighed. Sodium hydroxide solutions also react with carbon dioxide in the air to form carbonate compounds (**Figure 14**). This decreases the sodium hydroxide concentration in the solution.

Because of these problems, chemists must first standardize the solution of sodium hydroxide needed for an analysis. Standardizing a solution involves accurately determining its concentration. This is done by titrating the solution against a primary standard. A **primary standard** is a chemical that is highly pure and chemically stable. It should react completely and almost instantaneously and should ideally have a large molecular mass. Potassium hydrogen phthalate, $KHC_8H_4O_4$ (abbreviated to KHP), is a common primary standard used to standardize basic solutions. During this process, a measured mass of KHP is titrated by the basic solution. Chemists use the amount of KHP present to determine the amount concentration of base in the titrant. The equation for the neutralization of sodium hydroxide with KHP is

$$KHP(aq) + NaOH(aq) \rightarrow H_2O(l) + NaKP(aq)$$

Sodium carbonate is a common primary standard used to standardize acidic solutions. The equation for the neutralization of hydrochloric acid with sodium carbonate is

$$2\ HCl(aq) + Na_2CO_3(aq) \rightarrow H_2O(l) + CO_2(g) + 2\ NaCl(aq)$$

Figure 14 Sodium hydroxide reacts with carbon dioxide to form solid sodium carbonate. If the mouth and glass stopper of this bottle are not cleaned regularly, sodium carbonate can glue the stopper to the bottle, making it almost impossible to open.

primary standard a highly pure and stable chemical used to determine the precise concentration of acids or bases

Tutorial 2 Titration Calculations

Titrations involve several trials to ensure that the data collected are reliable. Data from the three or four most consistent trials are then averaged and used to determine the concentration of the solution being analyzed. Calculations involving titration data follow the same general process as any stoichiometry calculation.

Investigation **10.3.2**

What Is the Amount Concentration of Ethanoic Acid in Vinegar? (p. 488)
Put your skills and understanding to the test to experimentally determine the concentration of a common solution.

Sample Problem 1: Determining Concentration Using Titration Data

Several 10.00 mL samples of sulfuric acid solution of unknown concentration are titrated with a 0.100 mol/L solution of sodium hydroxide. (Note that the burette contains the sodium hydroxide solution.) The endpoint was determined using phenolphthalein indicator. The acceptable observations from the titration are summarized in **Table 3**. Use these data to determine the amount concentration of the acid solution.

Table 3 Titration Data for the Titration of Sulfuric Acid with Sodium Hydroxide Solution

Trial	1	2	3
final burette volume reading (mL)	12.52	24.98	37.62
initial burette volume reading (mL)	00.10	12.52	25.10
volume of base (titrant) added (mL)	12.42	12.46	12.52

Given: volume of acid, $V_{H_2SO_4} = 10.00$ mL; concentration of base, $c_{NaOH} = 0.100$ mol/L; volume of base, V_{NaOH} (see Table 3)

Required: concentration of acid, $c_{H_2SO_4}$

Analysis: $c = \frac{n}{V}$

Solution:

Step 1. Calculate the average volume of titrant used. In this example, the titrant is the base.

$$V_{NaOH\,(average)} = \frac{12.42\text{ mL} + 12.46\text{ mL} + 12.52\text{ mL}}{3}$$

$$V_{NaOH\,(average)} = 12.4667\text{ mL}\,[2\text{ extra digits carried}]$$

Step 2. Convert the given volumes to litres.

$$V_{NaOH\,(average)} = 12.47\ \cancel{\text{mL}} \times \frac{1\text{ L}}{1000\ \cancel{\text{mL}}}$$

$$V_{NaOH\,(average)} = 0.01247\text{ L}$$

$$V_{H_2SO_4} = 10.00\ \cancel{\text{mL}} \times \frac{1\text{ L}}{1000\ \cancel{\text{mL}}}$$

$$V_{H_2SO_4} = 0.01000\text{ L}$$

Step 3. Write a balanced equation for the reaction, listing the given value(s), required value(s), molar masses of solids (if any), and amount concentrations of solutions below the entities being considered in the problem.

$$2\,NaOH(aq) + H_2SO_4(aq) \rightarrow 2\,H_2O(l) + Na_2SO_4(aq)$$

0.01247 L 0.01000 L

0.100 mol/L $c_{H_2SO_4}$

Step 4. Use the concentration equation to determine the amount of the substance whose volume and concentration are given, n_{NaOH}.

$$c_{NaOH} = \frac{n_{NaOH}}{V_{NaOH}}$$

$$n_{NaOH} = c_{NaOH} V_{NaOH}$$

$$= 0.100\,\frac{\text{mol}}{\cancel{\text{L}}} \times 0.01247\ \cancel{\text{L}}$$

$$n_{NaOH} = 1.247 \times 10^{-3}\text{ mol}$$

Step 5. Use the amount of the substance determined in Step 4 and the mole ratio in the balanced equation to determine the amount of the required entity, $n_{H_2SO_4}$.

$$n_{H_2SO_4} = 1.247 \times 10^{-3}\ \cancel{\text{mol}}_{NaOH} \times \frac{1\ \text{mol}_{H_2SO_4}}{2\ \cancel{\text{mol}}_{NaOH}}$$

$$= 6.233 \times 10^{-4}\text{mol}\ [1\ \text{extra digit carried}]$$

Step 6. Use the amount of the required substance (determined in Step 5), the volume of the required substance, and the concentration equation to determine the amount concentration of the required substance, $c_{H_2SO_4}$.

$$c_{H_2SO_4} = \frac{n_{H_2SO_4}}{V_{H_2SO_4}}$$

$$= \frac{6.233 \times 10^{-4}\ \text{mol}}{1.000 \times 10^{-2}\ \text{L}}$$

$$c_{H_2SO_4} = 6.233 \times 10^{-2}\ \frac{\text{mol}}{\text{L}}$$

Statement: The amount concentration of the sulfuric acid solution is 6.233×10^{-2} mol/L.

The strategy used in Sample Problem 1 also applies to problems involving a primary standard. The only difference is that the mass of the primary standard is used rather than a volume.

Sample Problem 2: Determining Concentration Using a Primary Standard

A 0.306 g mass of potassium hydrogen phthalate, $KHC_8H_4O_4(s)$ or KHP, is dissolved in 50 mL of water. This solution is titrated with a sodium hydroxide solution of unknown concentration. It is determined that 14.80 mL of the sodium hydroxide solution is needed to titrate the potassium hydrogen phthalate sample to the endpoint at which phenolphthalein indicator changes colour. Calculate the amount concentration of the sodium hydroxide solution.

Given: mass of acid, $m_{KHP} = 0.306$ g; volume of base, $V_{NaOH} = 14.80$ mL

Required: amount concentration of sodium hydroxide solution, c_{NaOH}

Solution:

Step 1. Convert the volume of titrant used to litres.

$$V_{NaOH} = 14.80\ \cancel{\text{mL}} \times \frac{1\ \text{L}}{1000\ \cancel{\text{mL}}}$$

$$V_{NaOH} = 0.01480\ \text{L}$$

Step 2. Write a balanced equation for the reaction, listing the given value(s), required value(s), molar masses of solid(s), and amount concentrations below the entities being considered in the problem.

KHP(aq)	+	$NaOH(aq) \rightarrow H_2O(l) + NaKP(aq)$
0.306 g		0.01480 L
204.23 g/mol		c_{NaOH}

Step 3. Convert mass of given substance (KHP) into amount of given substance.

$$n_{KHP} = 0.306\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{204.23\ \cancel{\text{g}}}$$

$$n_{KHP} = 1.4983 \times 10^{-3}\ \text{mol}\ [2\ \text{extra digits carried}]$$

Step 4. Use the amount of the substance (determined in Step 3) and the mole ratio in the balanced equation to determine the amount of the required entity, n_{NaOH}.

$$n_{NaOH} = 1.4983 \times 10^{-3}\ \cancel{\text{mol}}_{KHP} \times \frac{1\ \text{mol}_{NaOH}}{1\ \cancel{\text{mol}}_{KHP}}$$

$$n_{NaOH} = 1.4983 \times 10^{-3}\ \text{mol}\ [2\ \text{extra digits carried}]$$

LEARNING **TIP**

Masses and Moles

The amount of a pure substance in a sample is determined only by its mass and not by the volume of water it is dissolved in. That is why the 50 mL volume of water was not involved in calculating the amount of KHP in this problem.

Step 5. Use the amount of the required substance (determined in Step 4), the volume of the required substance, and the concentration equation to determine the amount concentration of the required substance, c_{NaOH}.

$$c_{NaOH} = \frac{n_{NaOH}}{V_{NaOH}}$$

$$= \frac{1.4983 \times 10^{-3}\ \text{mol}}{1.480 \times 10^{-2}\ \text{L}}$$

$$c_{NaOH} = 0.101\ \text{mol/L}$$

Statement: The amount concentration of the sodium hydroxide solution is 0.101 mol/L.

Practice

SKILLS HANDBOOK A6.2, A6.5

1. A 0.500 mol/L solution of nitric acid, $HNO_3(aq)$, is used to titrate several 25.00 mL samples of a sodium hydroxide solution, NaOH(aq). The concentration of the sodium hydroxide solution is unknown. The endpoint is detected using bromothymol blue indicator. The acceptable data collected from this titration are summarized in **Table 4**. T/I

Table 4 Titration Data for the Titration of Sodium Hydroxide Solution with Nitric Acid

Trial	1	2	3
final burette volume reading (mL)	8.00	16.12	24.62
initial burette volume reading (mL)	0.00	8.06	16.50

   (a) What is the average volume of nitric acid used? [ans: 8.06 mL]
   (b) Calculate the amount concentration of the sodium hydroxide solution. [ans: 0.161 mol/L]

2. A 1.00 g sample of potassium carbonate is dissolved in 100 mL water. A titration of this sample requires 24.20 mL of nitric acid to reach the endpoint. Calculate the amount concentration of the nitric acid. T/I [ans: 0.598 mol/L]
3. 15.52 mL of 0.100 mol/L hydrochloric acid is required to titrate 25.00 mL of a barium hydroxide solution to a bromothymol blue endpoint. Calculate the amount concentration of the barium hydroxide solution. T/I [ans: 3.10×10^{-2} mol/L]

WEB LINK

If you would like more practice solving problems involving titration,

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10.3 Summary

- Titrations are used to determine the concentration of a solution using another solution whose precise concentration is known. A measured volume of the titrant is added, from a burette, to a measured volume of the other solution until an endpoint is reached.
- The equivalence point of a titration is reached when the sample in the titration flask is completely neutralized.
- The endpoint of a titration is a visual cue used to indicate when the equivalence point is reached. The endpoint can be a sudden change in colour of an acid-base indicator or a sudden change in pH detected using a pH meter.
- A primary standard is a highly pure and stable chemical used in a titration to determine accurately the concentration of a solution.
- The pH of the solution in the flask changes sharply near the equivalence point.
- Appropriate acid-base indicators for a titration change colour during the steepest portion of the pH curve (near the equivalence point).

10.3 Questions

1. What is the purpose of an acid–base indicator in a titration? K/U
2. Distinguish between equivalence point and endpoint. K/U
3. A technician carefully measures 4.00 g of solid sodium hydroxide and dissolves it in water to prepare 100.0 mL of solution. T/I
   (a) Calculate the predicted concentration of this solution.
   (b) Titrating this solution against a primary standard showed that the sodium hydroxide concentration was actually 0.940 mol/L. Suggest why this concentration differs from the predicted concentration.
4. Why do you think it is necessary to re-standardize a solution after it has been stored for quite some time? K/U T/I
5. Why do you think storage bottles of solutions are often filled right to the top? K/U
6. Potassium hydrogen phthalate samples are generally dried in an oven before being used to standardize a base. Why do you think this step is necessary? What effect might omitting this step have on the final concentration of the base? Why? K/U T/I
7. Describe how a pH meter can be used to determine the equivalence point of a titration. What factor determines whether or not an acid–base indicator is suitable for a titration? K/U
8. Two different acids of the same concentration were titrated with sodium hydroxide solution. The pH changes occurring during these titrations were measured and plotted on the same graph (**Figure 15**). T/I

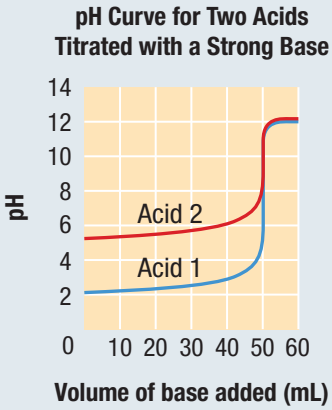

Figure 15

   (a) Suggest an explanation to account for the different initial pH of the acids.
   (b) What volume of base is required to neutralize these acids? Why is this volume the same for both acids?
   (c) Select an acid–base indicator from Table 2 that could be used for the titration of one of the acids but not for the other.
9. Calcium carbonate is the active ingredient in many antacid medications. An antacid tablet neutralizes 25.20 mL of 0.50 mol/L hydrochloric acid. T/I C
   (a) Write the chemical equation for the neutralization reaction.
   (b) What is the mass of calcium carbonate in the tablet? (Assume that calcium carbonate is the only ingredient that reacts with acid.)
10. An ammonium hydroxide solution is standardized by titrating 10.00 mL samples with a 0.25 mol/L solution of sulfuric acid. The titration data, indicating the volumes of acid added from the burette, are recorded in **Table 5**. T/I C

Table 5 Titration Data for the Titration of Ammonium Hydroxide Solution with 0.25 mol/L Sulfuric Acid

Trial	1	2	3
final burette volume reading (mL)	8.46	17.00	25.66
initial burette volume reading (mL)	0.00	8.50	17.00

   (a) Write the balanced chemical equation for the neutralization reaction.
   (b) Calculate the average volume of acid used.
   (c) Determine the amount concentration of the ammonium hydroxide solution.
   (d) Why do you think it is necessary to conduct more than one trial when standardizing a solution?
11. The chemical equation for the neutralization of lactic acid, $HC_3H_5O_2$, with sodium hydroxide is

$$HC_3H_5O_2(aq) + NaOH(aq) \rightarrow H_2O(l) + NaC_3H_5O_2(aq)$$

    During a titration analysis for lactic acid in milk, several 10.00 mL milk samples were titrated using a 0.100 mol/L sodium hydroxide solution. Phenolphthalein indicator was used to indicate the endpoint. On average, 1.60 mL of the sodium hydroxide solution was required for each titration. Calculate the lactic acid concentration in the milk. T/I
12. Potassium hydrogen iodate, $KH(IO_3)_2$, is a primary standard sometimes used to standardize base solutions. For each of several trials, 1.50 g of $KH(IO_3)_2$ was first dissolved in water in the titration flask. An average of 25.42 mL of a potassium hydroxide solution was required to reach the endpoint. T/I C
    (a) Write a balanced equation for the neutralization of a solution of potassium hydroxide by a solution of potassium hydrogen iodate.
    (b) Determine the amount concentration of the potassium hydroxide solution.
13. Potassium carbonate, K_2CO_3, is a primary standard used to standardize a hydrochloric acid solution. 1.00 g samples of solid potassium carbonate are first dissolved in water to make 100 mL of solution. Over several trials, an average of 24.20 mL of hydrochloric acid is required to titrate the potassium carbonate solution to a phenolphthalein endpoint. Determine the amount concentration of the hydrochloric acid solution. T/I