SECTION 10.2

Neutralization Reactions and Acid-base Titrations

Key Terms

neutralization reaction
salt
titration
titrant
burette
end point
equivalence point

Acids and bases are found in air, soil, oceans, and waterways. Lactic acid is found in spoiled milk and in tired muscles after a strenuous workout. Acids and bases are also used widely in industry. The manufacturing of fertilizers, fabrics, soaps, plastics, pesticides, and numerous other chemicals relies on chemical reactions that involve acids and bases. **Table 10.5** lists some acids and bases that are manufactured in million tonne quantities worldwide.

Table 10.5 Ranking of Some Acids and Bases by Quantity Manufactured

Acid or Base	Uses
1. Sulfuric acid, $H_2SO_4(aq)$	Making phosphate fertilizers and car batteries
2. Ammonia, $NH_3(aq)$	Making fertilizers and explosives
3. Sodium hydroxide, $NaOH(aq)$	Making soaps and detergents; refining vegetable oil; peeling fruits and vegetables
4. Phosphoric acid, $H_3PO_4(aq)$	Making detergents, food additives, and phosphate fertilizers
5. Nitric acid, $HNO_3(aq)$	Making fertilizers, explosives, and plastics
6. Hydrochloric acid, $HCl(aq)$	Cleaning metal products; making chlorides, fertilizers, and dyes
7. Acetic acid, $CH_3COOH(aq)$	Preserving food; making vinegar, adhesives, and paints

Neutralization Reactions

neutralization reaction a reaction between an acid and a base

salt a compound composed of a metal cation from a base and an anion from an acid

The reaction between an acid and base is often called a **neutralization reaction**. In a neutralization reaction, the acid counteracts (or neutralizes) the properties of the base, and the base counteracts the properties of the acid.

During a neutralization reaction between an Arrhenius acid and an Arrhenius base, the hydrogen ions, $H^+(aq)$, from the acid combine with the hydroxide ions, $OH^-(aq)$, from the base to form water, $H_2O(\ell)$. The metal cation from the base and the anion from the acid combine to form an ionic compound called a **salt**.

In everyday language, the word *salt* is used to mean table salt, which is the ionic compound sodium chloride, $NaCl(s)$. In chemistry, however, a salt can be made of different cations and anions. For example, the neutralization reaction between hydrochloric acid, $HCl(aq)$, and potassium hydroxide, $KOH(aq)$, produces an aqueous solution of the salt potassium chloride, $KCl(aq)$, as shown in **Figure 10.10**.

$$\mathrm{HCl(aq) + KOH(aq) \rightarrow HOH(\ell) + KCl(aq)}$$

Figure 10.10 The cation from the base and the anion from the acid form a salt. The formula for water, $H_2O(\ell)$, can be written as $HOH(\ell)$ to see more clearly how the hydrogen ion from the acid combines with the hydroxide ion from the base to form water.

The reaction between any aqueous solution of a strong acid and any aqueous solution of a strong base forms a neutral salt. If the molar amounts are balanced—that is, if there are equal numbers of aqueous hydrogen ions and aqueous hydroxide ions—all the acid and all the base will be neutralized, leaving a solution with a pH of 7. However, reactions that involve equal amounts of a weak acid and/or a weak base usually produce a solution with a pH that is not 7 because the ions react with water. Such reactions are still called neutralization reactions even though the solution at the end of the reaction is not neutral. Using the term *neutralization reaction* in this situation may be confusing because the term *neutral* in everyday language implies that everything is balanced and unreactive.

Calculations That Involve Neutralization Reactions

When solutions of an acid and a base undergo a neutralization reaction, the concentration of one can be determined if the concentration of the other is known, and if the volumes of both have been accurately measured. You can apply the solution stoichiometry techniques you learned in Chapter 9 to solve problems that involve neutralization reactions. Recall that you need to begin with the appropriate balanced chemical equation. This equation tells you the molar ratios of the reactants and products. For example, you could determine the concentration of an aqueous solution of potassium hydroxide by measuring the volume of a solution that neutralizes a measured volume of hydrochloric acid with a known concentration.

Sample Problem

Determining the Concentration of a Base in a Neutralization Reaction

Problem

A technician performed three trial reactions to determine the concentration of a solution of potassium hydroxide, KOH(aq). In each trial, a 0.1250 mol/L solution of hydrochloric acid, HCl(aq), was used to neutralize a 25.00 mL sample of the potassium hydroxide solution. The average volume of hydrochloric acid required was 32.86 mL. Determine the concentration of the potassium hydroxide solution.

What Is Required?

You need to determine the concentration of the potassium hydroxide solution.

What Is Given?

You know each of the following:

Volume of KOH(aq): 25.00 mL

Volume of HCl(aq): 32.86 mL

Concentration of HCl(aq): 0.1250 mol/L

Plan Your Strategy	Act on Your Strategy
Write the balanced chemical equation for the reaction.	$HCl(aq) + KOH(aq) \longrightarrow KCl(aq) + H_2O(\ell)$
Use the formula $n = cV$ to determine the amount of hydrochloric acid. Remember to convert the volume in mL to L.	$n = cV$ $n = 0.1250\ \text{mol/L} \times 32.86\ \text{mL}$ $= 0.1250\ \text{mol}/\cancel{\text{L}} \times 0.032\ 86\ \cancel{\text{L}}$ $= 4.1075 \times 10^{-3}\ \text{mol}$
Determine the amount of potassium hydroxide needed to neutralize the hydrochloric acid.	The OH^- in KOH(aq) reacts with the H^+ in HCl(aq) in a 1:1 ratio, so the amount of KOH(aq) is 4.1075×10^{-3} mol.
Substitute values into the formula for molar concentration to calculate the concentration of the potassium hydroxide solution. Remember to convert the volume in mL to L.	$c = \frac{n}{V}$ $c = \frac{4.1075 \times 10^{-3}\ \text{mol}}{0.025\ 00\ \text{L}} = 0.1643\ \text{mol/L}$ The concentration of the potassium hydroxide solution is 0.1643 mol/L.

Check Your Solution

The balanced equation shows that hydrochloric acid and potassium hydroxide react in a 1:1 ratio. The volume of potassium hydroxide required for neutralization was less than the volume of hydrochloric acid. Therefore, the concentration of potassium hydroxide must be greater than the concentration of hydrochloric acid. The answer is reasonable.

Practice Problems

1. Hydrochloric acid was slowly added to an Erlenmeyer flask that contained 50.0 mL of 1.50 mol/L sodium hydroxide, $NaOH(aq)$, and a pH meter. The pH meter read 7.0 after the addition of 35.3 mL of hydrochloric acid. Calculate the concentration of the hydrochloric acid.
2. What volume of 0.400 mol/L sodium hydroxide, $NaOH(aq)$, is needed to neutralize 26.8 mL of 0.504 mol/L sulfuric acid, $H_2SO_4(aq)$, completely? **Hint:** Sulfuric acid loses two hydrogen ions during this neutralization reaction.
3. A 25.00 mL sample of a nitric acid solution, $HNO_3(aq)$, is neutralized by 18.55 mL of a 0.1750 mol/L sodium hydroxide, $NaOH(aq)$. What is the concentration of the nitric acid solution?
4. What volume of 1.25 mol/L hydrobromic acid, $HBr(aq)$, will neutralize 75.0 mL of 0.895 mol/L magnesium hydroxide, $Mg(OH)_2(aq)$?
5. A solution of sodium hydroxide was prepared by dissolving 4.0 g of sodium hydroxide, $NaOH(s)$, in 250 mL of water. It was found that 20.0 mL of the sodium hydroxide solution neutralizes 25.0 mL of vinegar. Determine the concentration of acetic acid, $CH_3COOH(aq)$, in the sample of vinegar. Assume that acetic acid is the only acidic substance in the vinegar.
6. Phosphoric acid, $H_3PO_4(aq)$, is a triprotic acid. If 15.0 mL of phosphoric acid completely neutralizes 38.5 mL of 0.150 mol/L sodium hydroxide, $NaOH(aq)$, what is the concentration of the phosphoric acid?
7. The acidity of a water sample can be measured by a neutralization reaction with a solution of sodium hydroxide, $NaOH(aq)$. What is the concentration of hydrogen ions in a water sample if 100 mL of the sample is neutralized by the addition of 8.0 mL of 2.5×10^{-3} mol/L sodium hydroxide?
8. Citric acid, $H_3C_6H_5O_7(aq)$, is a weak, triprotic acid that occurs naturally in many fruits and vegetables, especially the citrus fruits from which it gets its name. What volume of 0.165 mol/L sodium hydroxide, $NaOH(aq)$, will completely react with 40.0 mL of 0.120 mol/L citric acid? For this calculation, assume that all the hydrogen ions are released by the citric acid.
9. Phosphoric acid, $H_3PO_4(aq)$, is a weak, triprotic acid. When phosphoric acid reacts with a base, different salts can be prepared, depending on how many hydrogen ions are replaced by cations. For example, potassium hydrogen phosphate, $K_2HPO_4(aq)$, can be prepared in an aqueous solution by adding just enough potassium hydroxide, $KOH(aq)$, to replace two hydrogen ions:
$$H_3PO_4(aq) + 2KOH(aq) \rightarrow K_2HPO_4(aq) + 2H_2O(\ell)$$
What volume of 0.185 mol/L potassium hydroxide should be added to 80.0 mL of 0.137 mol/L phosphoric acid to form a solution of potassium hydrogen phosphate?
10. What volume of 0.150 mol/L calcium hydroxide, $Ca(OH)_2(aq)$, is needed to completely neutralize 20 mL of 0.185 mol/L sulfuric acid, $H_2SO_4(aq)$?

Acid-base Titration Can Determine the Concentration of a Solution

To determine the concentration of an acid or a base, chemists perform a neutralization reaction while doing a procedure called **titration**. In a titration, the concentration of a solution is determined by reacting a known volume of that solution with a measured volume of a solution with a known concentration.

For example, suppose that you wanted to find the concentration of an acid solution. You would gradually add a basic solution with a known concentration to an accurately measured volume of the acid solution to find the volume of basic solution that would completely react. Then you would use stoichiometry to calculate the concentration of the acid. In a titration, the solution with the known concentration is the **titrant**.

Special glassware is used in a titration experiment. A sample of the solution with an unknown concentration is drawn into a volumetric pipette or a graduated pipette. The titrant is poured into a **burette**. A burette is a long, narrow graduated tube, with a tap on the bottom end. The burette is used to measure the volume of titrant that is added to the sample. The procedure on pages 468 to 469 describes how to do a titration and how to use these pieces of glassware during the titration.

titration a procedure that is used to determine the concentration of a solution by reacting a known volume of that solution with a measured volume of a solution that has a known concentration

titrant in a titration, the solution with a known concentration

burette a clear tube with volume markings along its length and a tap at the bottom

Acid-base Indicators and Titration

In a titration of hydrochloric acid with aqueous potassium hydroxide, the two clear, colourless solutions react to form another clear, colourless solution. The temperature of the solution rises because the reaction is exothermic, but there is no other visible sign that a reaction took place. To know when a neutralization reaction is complete, chemists often use an acid-base indicator.

Figure 10.11 shows the colour change of phenolphthalein, a common indicator that is used in titrations. Phenolphthalein is colourless between pH 0 and pH 8. It turns pink between pH 8 and pH 10, and it is red in more basic solutions. This pH range may seem too large to be useful for a titration, but the colour change is quite abrupt. A single drop of titrant is usually enough to cause a vivid colour change. The point when the indicator changes colour is called the **end point** of the titration.

end point the point at which the indicator in a titration changes colour

equivalence point the point at which the amount of titrant is just enough to react with all of the reactant in the sample

Go to **scienceontario** to find out more

Figure 10.11 Like all indicators, phenolphthalein has a distinct colour change. The solution on the left is acidic, and the solution on the right is basic.

The aim of a titration is to know when the amount of titrant that has been added to the sample is just enough to react with all the acid or base the sample contains. This point is referred to as the **equivalence point**. Ideally, the end point and equivalence point should coincide—that is, the indicator should be in the middle of its colour change at the pH of the equivalence point.

Phenolphthalein is often used for titrating a strong acid with a strong base, even though the equivalence point, which is at pH 7.0, is not in phenolphthalein's range. However, near the equivalence point, the pH of a titrated solution changes *very* rapidly. A fraction of a drop of titrant beyond the equivalence point will place the pH of the solution in the range where phenolphthalein changes colour.

Phenolphthalein is also perfect for titrating a weak acid with a strong base. The salt solution formed is mildly basic, so phenolphthalein changes colour very close to the equivalence point. However, methyl orange is more suitable for titrating a weak base with a strong acid because the solution at the equivalence point is mildly acidic.

Suggested Investigation

Inquiry Investigation 10-B, The Concentration of Acetic Acid in Vinegar

Inquiry Investigation 10-C, The Percent (m/m) of Ascorbic Acid in a Vitamin C Tablet

Learning Check

13. What are the products of a neutralization reaction?

14. How does the pH of the equivalence point in a titration between a strong acid and a strong base compare with the pH of the equivalence point in a titration involving either a weak base or a weak acid?

15. Why must you know the concentration of one of the solutions used during titration?

16. Explain why a solution that is produced by a neutralization reaction may not have a pH of 7.

17. Explain the difference between the end point and the equivalence point of a titration.

18. A student performed a titration between hydrochloric acid and a weak base. The student used a pipette to add hydrochloric acid to the reaction flask and then added a few drops of phenolphthalein as the indicator. At the equivalence point, the solution in the reaction flask was acidic. Explain how the choice of indicator could cause an error in determining the concentration of hydrochloric acid.

Procedure for an Acid-base Titration

The following steps describe how to prepare for and perform a titration.

Figure A Squeeze the pipette bulb as you put it on the stem of the pipette.

Rinsing the Volumetric or Graduated Pipette

Rinse a pipette with the solution whose volume you are measuring. This will ensure that any drops remaining inside the pipette will form part of the measured volume.

1. Put the pipette bulb on the pipette, as shown in **Figure A**. Place the tip of the pipette into a beaker of distilled water.
2. Relax your grip on the bulb to draw up a small volume of distilled water.
3. Remove the bulb, and discard the water by letting it drain out.
4. Pour a sample of the solution with the unknown concentration into a clean, dry beaker.
5. Rinse the pipette by drawing several millilitres of the solution with the unknown concentration from the beaker into the pipette. Coat the inner surface with the solution, as shown in **Figure B**. Discard the rinse. Rinse the pipette twice in this way. The pipette is now ready to be filled with the solution that has the unknown concentration.

Figure B Cover the ends of the pipette so that none of the solution spills out as you rock the pipette back and forth to coat its inner surface with solution.

Filling the Pipette

6. Place the tip of the pipette below the surface of the solution with the unknown concentration.
7. Hold the suction bulb loosely on the end of the glass stem. Use the suction bulb to draw the solution up to the point shown in **Figure C**.
8. As quickly and smoothly as you can, slide the bulb off the glass stem and place your index finger over the end.
9. Roll your finger slightly away from end of the stem to let the solution slowly drain out.
10. When the bottom of the meniscus aligns with the etched mark, as in **Figure D**, press your finger back over the end of the stem. This will prevent more solution from draining out.
11. Touch the tip of the pipette to the side of the beaker to remove any clinging drops. The measured volume inside the pipette is now ready to transfer to an Erlenmeyer flask.

Transferring the Solution

12. Place the tip of the pipette against the inside glass wall of the flask, as shown in **Figure E.** Let the solution drain slowly, by removing your finger from the stem.
13. After the solution drains, wait several seconds and then touch the tip to the inside wall of the flask to remove any drops on the end. Note: Do not remove the small amount of solution shown in **Figure F**.

Figure C Start with more of the unknown solution than you need. You will drain out the excess solution in the next two steps.

Figure D Always read the volume of the solution at the bottom of the meniscus.

Figure E Draining the pipette with the tip against the wall of the flask will prevent splashing.

Adding the Indicator

14. Add two or three drops of the indicator to the flask and its contents. Do not add too much indicator. Using more indicator does not make the colour change easier to see. Also, most indicators are weak acids. Too much indicator can change the amount of base needed for the neutralization. You are now ready to prepare the apparatus for the titration.

Rinsing the Burette

15. To rinse the burette, close the tap and add about 10 mL of distilled water from a wash bottle.

16. Tip the burette to one side, and roll it gently back and forth so that the water comes in contact with all the inner surfaces.

17. Hold the burette over a sink. Let the water drain out, as shown in **Figure G**. While you do this, check that the tap does not leak. Make sure that the tap turns smoothly and easily.

18. Rinse the burette twice, with 5 to 10 mL of the titrant. Remember to open the tap to rinse the lower portion of the burette. Discard the rinse solution each time.

Filling the Burette

19. Assemble a retort stand and burette clamp to hold the burette. Place a funnel in the top of the burette, and put a beaker under the burette.

20. With the tap closed, add the solution until it is above the zero mark. Remove the funnel. Carefully open the tap. Drain the solution into the beaker until the bottom of the meniscus is at or below the zero mark.

21. Touch the tip of the burette against the beaker to remove any clinging drops. Check that the part of the burette below the tap is filled with solution and contains no air bubbles. **Figure H** shows the air bubbles that you should avoid.

22. Find the initial burette reading using a meniscus reader, as shown in **Figure I**. Record the initial volume to the nearest 0.05 mL.

Titrating the Unknown Solution

23. Replace the beaker with the Erlenmeyer flask that contains the solution you want to titrate. Place a sheet of white paper under the flask to help you see the colour change.

24. Add titrant from the burette to the Erlenmeyer flask by opening the tap, as shown in **Figure J**. You may start by adding the titrant quickly, but slow down when you start to see a colour change in the solution in the flask.

25. At first, the colour change will disappear as you mix the solution in the flask. Add a small amount of titrant, and swirl thoroughly before adding any more. Stop adding titrant when the solution in the Erlenmeyer flask has a persistent colour change. If you are using phenolphthalein as an indicator, stop when the solution is a faint pink colour.

26. Use the meniscus reader to read the final volume. Record this volume, and subtract the initial volume from it to find the volume of the titrant needed to reach the end point.

Figure J Always swirl the flask as you add the titrant. If you have trouble swirling and adding titrant at the same time, use a magnetic stirrer or have your laboratory partner swirl the flask as you add the titrant.

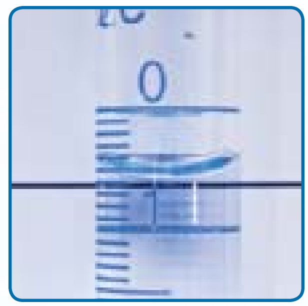

Figure I Hold the meniscus reader so that the line is under the meniscus.

Figure F A small amount of solution will always remain in the tip of the pipette. Do not remove this.

Figure G The tap is fully open when the handle on the tap is parallel to the burette and the solution inside the burette comes out quickly.

Figure H Do NOT start a titration if you have air bubbles like these in the tip of the burette. They will cause errors in your measurements.

Section 10.2 REVIEW

Section Summary

- A neutralization reaction between an acid and a base in aqueous solution forms a salt and water.
- An acid-base titration is a quantitative technique in which a neutralization reaction is used to determine the concentration of one solution.
- The end point of a titration occurs when the indicator changes colour.
- An indicator must be chosen to change colour near the equivalence point of the titration, when equal amounts of acid and base have reacted.

Review Questions

1. C Write a general word equation that describes all neutralization reactions between an Arrhenius acid and an Arrhenius base.
2. K/U Write a chemical equation for each neutralization reaction in aqueous solution.
   a. sulfuric acid with potassium hydroxide
   b. hydroiodic acid with magnesium hydroxide.
3. K/U Why might the equivalence point in a titration be different from the end point?
4. T/I Before using a pipette to draw up a standard solution of an acid, a student rinses the pipette with distilled water but not with the acid solution. How will the concentration of the acid determined by the titration be affected? Explain your answer.
5. T/I During a titration, a student uses the pipette bulb to force out the final drop of the basic solution from the pipette. How will this mistake affect the calculated concentration of the basic solution?
6. A Explain why some people rinse their hair with vinegar after washing it with shampoo.
7. K/U Explain why phenolphthalein, which changes colour in the pH range of 8.2 to 10.0, is used as the indicator for a titration that forms a solution with pH 7 at the equivalence point.
8. T/I What amount of calcium hydroxide, $Ca(OH)_2(aq)$, will be neutralized by 1 mol of hydrochloric acid?
9. A Heartburn is a condition that is caused when fluid from the stomach moves up into the esophagus, causing irritation. Some people use milk of magnesia, which contains magnesium hydroxide, $Mg(OH)_2(s)$, to relieve the symptoms of heartburn. Explain why this medicine works.
10. T/I What volume of 0.996 mol/L barium hydroxide, $Ba(OH)_2(aq)$, is needed to neutralize 25.0 mL of 1.70 mol/L nitric acid, $HNO_3(aq)$?
11. K/U Explain how a salt is formed during a neutralization reaction.
12. C Describe how you would design and perform a titration in which you use 0.250 mol/L sulfuric acid, $H_2SO_4(aq)$, to determine the concentration of a strontium hydroxide, $Sr(OH)_2(aq)$, solution. Include the equation for the reaction, as well as an outline of the calculations you would make.
13. T/I Methanoic acid, HCOOH(aq), is a weak monoprotic acid. A 25.00 mL sample of methanoic acid was titrated with a standard solution of 0.1004 mol/L sodium hydroxide, NaOH(aq). Three trials were conducted. The average volume of sodium hydroxide solution that was required to reach the end point was 16.32 mL. What is the concentration of the methanoic acid solution?
14. T/I Suppose that you titrate 25.0 mL of 0.100 mol/L sodium hydroxide, NaOH(aq), with two different acids. In the first titration, you use 0.150 mol/L hydrochloric acid, HCl(aq), which is a strong acid. In the second titration, you use 0.150 mol/L acetic acid, $CH_3COOH(aq)$, which is a weak acid. How will the volume of acid used in each titration compare? Explain your answer.
15. K/U Explain how the meaning of the term *salt* differs in chemistry and in everyday language.
16. T/I Suppose that you are going to use the solutions in the bottles shown for a titration. Which solution will you put in the burette? Which solution will you measure with a graduated pipette? Explain your answers.

STSE

CHEMISTRY Connections

Acid-base Reactions on the Rise

Have you ever watched a vinegar baking soda volcano erupt? The bubbles of carbon dioxide, $CO_2(g)$, result from a decomposition reaction that quickly follows an acid-base reaction between the vinegar $CH_3COOH(aq)$, an acid, and baking soda, $NaHCO_3(aq)$, a base, as shown below.

Acid-base Reaction

$$CH_3COOH(aq) + NaHCO_3(aq) \rightarrow NaCH_3COO(aq) + H_2CO_3(aq)$$

Decomposition

$$H_2CO_3(aq) \rightarrow CO_2(g) + H_2O(\ell)$$

The release of carbon dioxide as a result of the chemical reaction between an acid and a base, as shown in the photograph to the right, is part of the reason why baked goods rise. An ingredient that causes a batter to rise when baked is called a *leavening agent*. The two main chemical leavening agents are baking soda and baking powder.

BAKING SODA The chemical name for baking soda is sodium hydrogen carbonate. When used in cooking, baking soda reacts with mildly acidic liquids, forming carbon dioxide bubbles. Mildly acidic liquids include vinegar, molasses, honey, citrus juice, and buttermilk.

Baking soda must be mixed with other dry ingredients and added last to a batter so that the release of carbon dioxide is uniform throughout the batter. This acid-base reaction happens quickly. If baking soda is the only leavening agent in a recipe, the batter must be baked immediately before the bubbles have a chance to escape. Baking causes the bubbles to expand, and the batter rises. As the batter firms, the bubbles are trapped, as shown below.

Baking traps the bubbles that are formed during the reaction between an acid and a base, resulting in a light, airy cake.

BAKING POWDER If a recipe does not include an acidic liquid, baking powder is used. Most baking powder is a mixture of baking soda and two dry acids. One of the acids reacts with the baking soda when it dissolves in the batter, and the other acid reacts with the baking soda when it is heated.

Like baking soda, baking powder is mixed with other dry ingredients and added last to a batter. However, batter does not have to be baked immediately.

Sometimes, a batter made with mildly acidic liquid(s) includes both baking powder and baking soda. Excess acid can disrupt the action of the baking powder. The baking soda helps to neutralize the acid, and the baking powder provides a reliable source of carbon dioxide.

Carbon dioxide forms bubbles when baking soda, a base, is added to vinegar, an acid.

Connect to Society

Research to find out more about the chemical properties and relative costs of baking soda and baking powder. Create a chart summarizing your results. Based on your findings, write a recommendation to a recipe developer for an industrial bakery, suggesting which of the two bases would be best to include in recipes for mass-produced baked goods.