10.3 Acid-Base Reactions

Section Preview/ Specific Expectations

In this section, you will

- **perform** calculations involving neutralization reactions
- **determine** the concentration of an acid in solution by conducting a titration
- **communicate** your understanding of the following terms: *neutralization reaction, salt, acid-base indicator, titration, equivalence point, end-point*

Is there a box of baking soda in your refrigerator at home? Baking soda is sodium hydrogen carbonate. (It is also commonly called sodium bicarbonate.) Baking soda removes the odours caused by spoiling foods. The smelly breakdown products of many foods are acids. Baking soda, a base, eliminates the odours by neutralizing the characteristic properties of the acids.

Adding a base to an acid neutralizes the acid's acidic properties. This type of reaction is called a **neutralization reaction.**

There are many different acids and bases. Being able to predict the results of reactions between them is important. Bakers, for example, depend on neutralization reactions to create light, fluffy baked goods. Gardeners and farmers depend on these reactions to modify the characteristics of the soil. Industrial chemists rely on these reactions to produce the raw materials that are used to make a wide variety of chemicals and chemical products. (See Figure 10.13.)

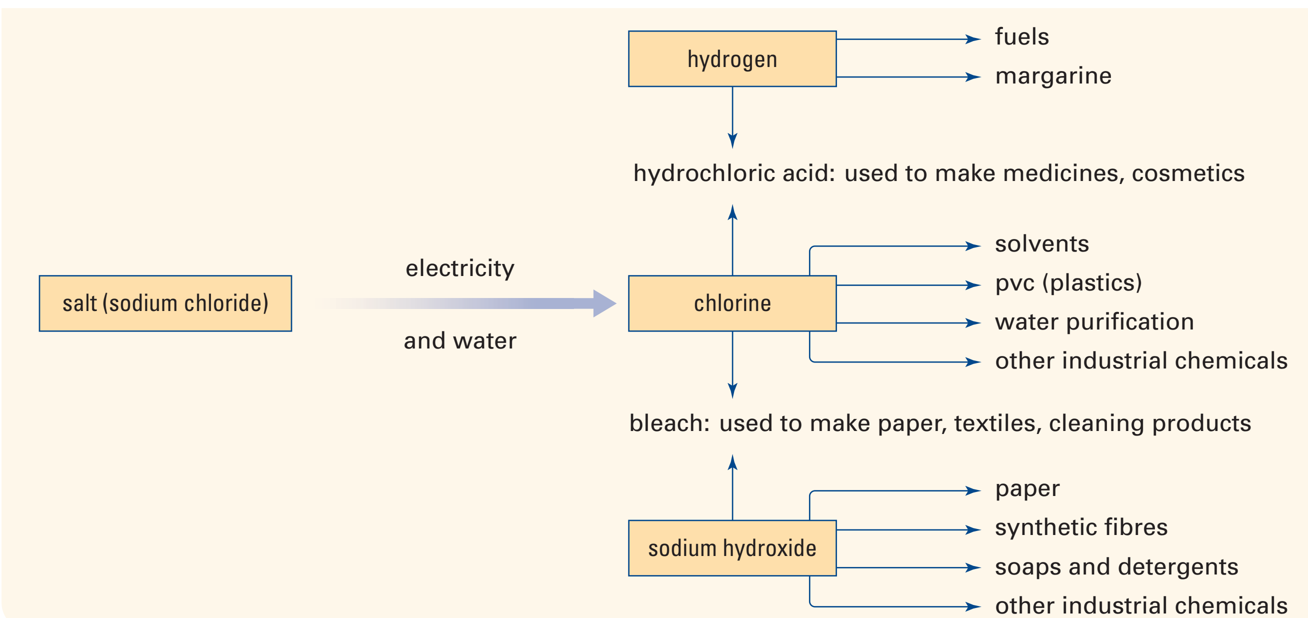

Figure 10.13 You know sodium chloride as common table salt. As you can see here, however, sodium chloride is anything but common. Sodium chloride is a product of an acid-base reaction between hydrochloric acid and sodium hydroxide.

Neutralization Reactions

The reaction between an acid and a base produces an ionic compound (a salt) and water.

$$\text{acid} + \text{base} \rightarrow \text{a salt} + \text{water}$$

A **salt** is an ionic compound that is composed of the anion from an acid and a cation from a base. For example, sodium nitrate is a salt that is found in many kitchens. It is often added to processed meat to preserve the colour and to slow the rate of spoiling by inhibiting bacterial growth. Sodium nitrate can be prepared in a laboratory by reacting nitric acid with sodium hydroxide, as shown on the next page.

cation from base

$$HNO_{3(aq)} + NaOH_{(aq)} \rightarrow NaNO_{3(aq)} + H_2O_{(\ell)}$$

anion from acid

The balanced chemical equation for this reaction shows that 1 mol of nitric acid reacts with 1 mol of sodium hydroxide. If equal molar quantities of nitric acid and sodium hydroxide are used, the result is a neutral (pH 7) aqueous solution of sodium nitrate. In fact, *when any strong acid reacts with any strong base in the mole ratio from the balanced chemical equation, a neutral aqueous solution of a salt is formed.* Reactions between acids and bases of different strengths usually do not result in neutral solutions.

For most neutralization reactions, there are no visible signs that a reaction is occurring. How can you determine that a neutralization reaction is taking place? One way is to use an **acid-base indicator.** This is a substance that changes colour in acidic and basic solutions. Most acid-base indicators are weak, monoprotic acids. The undissociated weak acid is one colour. Its conjugate base is a different colour.

colour 1

$$\text{H (indicator)}_{(aq)} \rightleftharpoons H^+ + \text{(indicator)}^-_{(aq)}$$

colour 2

In an acidic solution, the indicator does not dissociate very much. It appears as colour 1. In a basic solution, the indicator dissociates much more. It appears as colour 2. Often a single drop of indicator causes a dramatic change in colour. For example, phenolphthalein is an indicator that chemists often use for reactions between a strong acid and a strong base. It is colourless between pH 0 and pH 8. It turns pink between pH 8 and pH 10. (See Figure 10.14.)

Figure 10.14 **A good indicator, such as the phenolphthalein shown here, must give a vivid colour change.**

CHEM FACT

If a small quantity of an acid or a base is spilled in a laboratory, you can use a neutralization reaction to minimize the hazard. To neutralize a basic solution spill, you can add solid sodium hydrogen sulfate or citric acid. For an acidic solution spill, you can use sodium hydrogen carbonate (baking soda). Note that you cannot use a strong acid or base to clean up a spill. This would result in another hazardous spill. As well, the neutralization reaction would generate a lot of heat, and thus produce a very hot solution.

CHECKPOINT

Show that the net ionic equation for the reaction between HNO_3 (a strong acid) and NaOH (a strong base) results in the formation of water.

Calculations Involving Neutralization Reactions

Suppose that a solution of an acid reacts with a solution of a base. You can determine the concentration of one solution if you know the concentration of the other. (This assumes that the volumes of both are accurately measured.) Use the concentration and volume of one solution to determine the amount (in moles) of reactant that it contains. The balanced chemical equation for the reaction describes the mole ratio in which the compounds combine. In the following Sample Problems and Practice Problems, you will see how to do these calculations.

CHEM FACT

An old remedy to relieve the prickly sting of a nettle plant is to rub the area with the leaf of a dock plant. The sting contains an acid. This acid is neutralized by a base that is present in the dock leaf. Bees and ants also have an acidic sting. You can wash the sting with soap, because soap is basic. You can also apply baking soda (a base) to the skin for more effective relief. If you are stung by a wasp, however, you should apply vinegar. The sting of a wasp contains a base.

Sample Problem

Finding Concentration

Problem

13.84 mL of hydrochloric acid, $HCl_{(aq)}$, just neutralizes 25.00 mL of a 0.1000 mol/L solution of sodium hydroxide, $NaOH_{(aq)}$. What is the concentration of the hydrochloric acid?

What Is Required?

You need to find the concentration of the hydrochloric acid.

What Is Given?

Volume of hydrochloric acid, HCl = 13.84 mL
Volume of sodium hydroxide, NaOH = 25.00 mL
Concentration of sodium hydroxide, NaOH = 0.1000 mol/L

Plan Your Strategy

Step 1 Write the balanced chemical equation for the reaction.
Step 2 Calculate the amount (in mol) of sodium hydroxide added, based on the volume and concentration of the sodium hydroxide solution.
Step 3 Determine the amount (in mol) of hydrochloric acid needed to neutralize the sodium hydroxide.
Step 4 Find $[HCl_{(aq)}]$, based on the amount and volume of hydrochloric acid solution needed.

Act on Your Strategy

Step 1 The balanced chemical equation is

$$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(\ell)}$$

Step 2

$$\text{Amount (in mol)} = \text{Concentration (in mol/L)} \times \text{Volume (in L)}$$

$$\text{Amount NaOH (in mol) added} = 0.1000\ \text{mol}/\cancel{\text{L}} \times 0.02500\ \cancel{\text{L}} = 2.500 \times 10^{-3}\ \text{mol}$$

PROBLEM TIPS

1. Make sure that the values you use in your calculations refer to the same reactant. For example, you can use the concentration and volume of sodium hydroxide to find the amount of sodium hydroxide in this problem. You cannot use the concentration of sodium hydroxide and the volume of hydrochloric acid.
2. In the solution, the volumes are converted to litres. If all the volumes are expressed in the same unit, the conversion step is not necessary.
3. Do not drop significant digits, even zeros, during your calculations.

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FROM PAGE 396

Step 3 HCl reacts with NaOH in a 1:1 ratio, so there must be 2.500×10^{-3} mol HCl.

Step 4 $\text{Concentration (in mol/L)} = \dfrac{\text{Amount (in mol)}}{\text{Volume (in L)}}$

$$[HCl_{(aq)}] = \frac{2.500 \times 10^{3}\ \text{mol}}{0.01384\ \text{L}}$$
$$= 0.1806\ \text{mol/L}$$

Therefore, the concentration of hydrochloric acid is 0.1806 mol/L.

Check Your Solution

$[HCl_{(aq)}]$ is greater than $[NaOH_{(aq)}]$. This is reasonable because a smaller volume of hydrochloric acid was required. As well, the balanced equation shows a 1:1 mole ratio between these reactants.

Sample Problem

Finding Volume

Problem

What volume of 0.250 mol/L sulfuric acid, $H_2SO_{4(aq)}$, is needed to react completely with 37.2 mL of 0.650 mol/L potassium hydroxide, $KOH_{(aq)}$?

What Is Required?

You need to find the volume of sulfuric acid.

What Is Given?

Concentration of sulfuric acid, H_2SO_4 = 0.250 mol/L
Concentration of potassium hydroxide, KOH = 0.650 mol/L
Volume of potassium hydroxide, KOH = 37.2 mL.

Plan Your Strategy

Step 1 Write the balanced chemical equation for the reaction.

Step 2 Calculate the amount (in mol) of potassium hydroxide, based on the volume and concentration of the potassium hydroxide solution.

Step 3 Determine the amount (in mol) of sulfuric acid that is needed to neutralize the potassium hydroxide.

Step 4 Find the volume of the sulfuric acid, based on the amount and concentration of sulfuric acid needed.

Act on Your Strategy

Step 1 The balanced chemical equation is

$$H_2SO_{4(aq)} + 2KOH \rightarrow K_2SO_{4(aq)} + 2H_2O_{(\ell)}$$

Step 2 $\text{Amount (in mol) of KOH} = 0.650\ \text{mol}/\cancel{\text{L}} \times 0.0372\ \cancel{\text{L}}$
$$= 0.02418\ \text{mol}$$

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Step 3 H_2SO_4 reacts with KOH in a 1:2 mole ratio. The amount of H_2SO_4 needed is

$$\frac{1 \text{ mol } H_2SO_4}{2 \text{ mol KOH}} = \frac{0.01209 \text{ mol } H_2SO_4}{0.2418 \text{ mol KOH}}$$

$$\frac{0.024\,18 \text{ mol KOH} \times 1 \text{ mol } H_2SO_4}{2 \text{ mol KOH}} = 0.01209 \text{ mol } H_2SO_4$$

Step 4 Amount (in mol) H_2SO_4

$$= 0.01209 \text{ mol}$$

$$= 0.250 \text{ mol/L} \times \text{Volume } H_2SO_{4(aq)} \text{ (in L)}$$

$$\text{Volume } H_2SO_{4(aq)} = \frac{0.01209 \cancel{\text{mol}}}{0.250 \cancel{\text{mol}}/\text{L}}$$

$$= 0.04836 \text{ L}$$

Therefore, the volume of sulfuric acid that is needed is 48.4 mL.

Check Your Solution

The balanced chemical equation shows that half the amount of sulfuric acid will neutralize a given amount of potassium hydroxide. The concentration of sulfuric acid, however, is less than half the concentration of potassium hydroxide. Therefore, the volume of sulfuric acid should be greater than the volume of potassium hydroxide.

Practice Problems

10. 17.85 mL of nitric acid neutralizes 25.00 mL of 0.150 mol/L $NaOH_{(aq)}$. What is the concentration of the nitric acid?

11. What volume of 1.015 mol/L magnesium hydroxide is needed to neutralize 40.0 mL of 1.60 mol/L hydrochloric acid?

12. What volume of 0.150 mol/L hydrochloric acid is needed to neutralize each solution below?

- **(a)** 25.0 mL of 0.135 mol/L sodium hydroxide
- **(b)** 20.0 mL of 0.185 mol/L ammonia solution
- **(c)** 80 mL of 0.0045 mol/L calcium hydroxide

13. What concentration of sodium hydroxide solution is needed for each neutralization reaction?

- **(a)** 37.82 mL of sodium hydroxide neutralizes 15.00 mL of 0.250 mol/L hydrofluoric acid.
- **(b)** 21.56 mL of sodium hydroxide neutralizes 20.00 mL of 0.145 mol/L sulfuric acid.
- **(c)** 14.27 mL of sodium hydroxide neutralizes 25.00 mL of 0.105 mol/L phosphoric acid.

Acid-Base Titration

In the previous Sample Problems and Practice Problems, you were given the concentrations and volumes you needed to solve the problems. What if you did not have some of this information? Chemists often need to know the concentration of an acidic or basic solution. To acquire this information, they use an experimental procedure called a titration. In a **titration**, the concentration of one solution is determined by quantitatively observing its reaction with a solution of known concentration. The solution of known concentration is called a *standard solution*. The aim of a titration is to find the point at which the number of moles of the standard solution is stoichiometrically equal to the original number of moles of the unknown solution. This point is referred to as the **equivalence point.** At the equivalence point, all the moles of hydrogen ions that were present in the original volume of one solution have reacted with an equal number of moles of hydroxide ions from the other solution.

Precise volume measurements are needed when you perform a titration. Chemists use special glass apparatus to collect these measurements. (See Figure 10.15.) As well, an acid-base indicator is needed to monitor changes in pH during the titration.

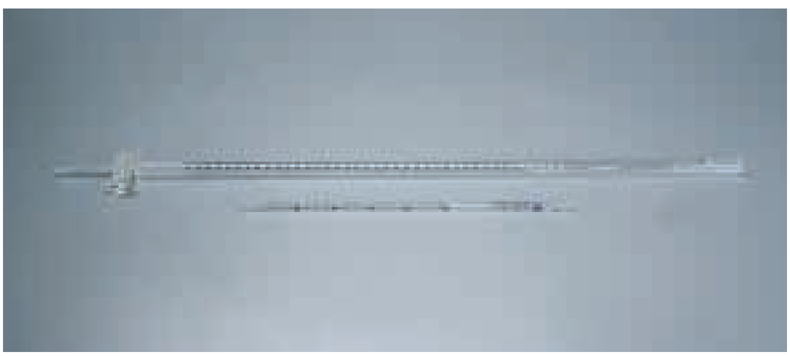

Figure 10.15 A transfer pipette (bottom) measures a fixed volume of liquid, such as 10.00 mL, 25.00 mL, or 50.0 mL. A burette (top) measures a variable volume of liquid.

In a titration, a pipette is used to measure a precise volume of standard solution into a flask. The flask sits under a burette that contains the solution of unknown concentration. After adding a few drops of indicator, you take an initial burette reading. Then you start adding the known solution, slowly, to the flask. The **end-point** of the titration occurs when the indicator changes colour. The indicator is chosen so that it matches its equivalance point.

Titration Step by Step

The following pages outline the steps that you need to follow to prepare for a titration. Review these steps carefully. Then observe as your teacher demonstrates them for you. At the end of this section, in Investigation 10-B, you will perform your own titration of a common substance: vinegar.

Web LINK

www.school.mcgrawhill.ca/resources

Sometimes a moving picture is worth a thousand words. To enhance your understanding, your teacher will demonstrate the titration procedure described in this textbook. In addition, some web sites provide downloadable or real-time titration movies to help students visualize the procedure and its techniques. Go to the web site above, then to Science Resources and to Chemistry 11 to see where to go next. Compare the different demonstrations you can find and observe, including your teacher's. Prepare your own set of "Titration Tips" to help you recall important details.

PROBEWARE

If you have access to probeware, try the Chemistry 11 lab, Titrating an Unknown, or a similar lab from a probeware company

TITRATION TIP

Never use your mouth in place of a suction bulb to draw a liquid into a pipette. The liquid could be corrosive or poisonous. As well, you will contaminate the glass stem.

TITRATION TIP

Practice removing the bulb and replacing it with your index finger. You need to be able to perform this action quickly and smoothly.

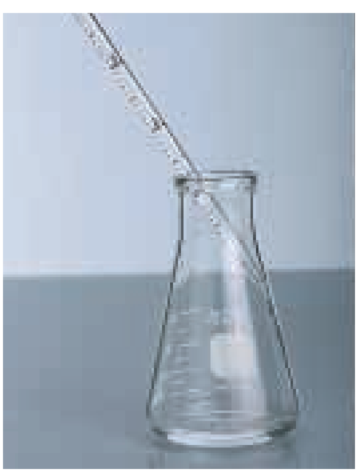

Figure 10.18 You can prevent a "stubborn" drop from clinging to the pipette tip by touching the tip to the inside of the glass surface.

Rinsing the Pipette

A pipette is used to measure and transfer a precise volume of liquid. You rinse a pipette with the solution whose volume you are measuring. This ensures that any drops that remain inside the pipette will form part of the measured volume.

1. Pour a sample of standard solution into a clean, dry beaker.
2. Place the pipette tip in a beaker of distilled water. Squeeze the suction bulb. Maintain your grip while placing it over the stem of the pipette. (If your suction bulbs have valves, your teacher will show you how to use them.)
3. Relax your grip on the bulb to draw up a small volume of distilled water.
4. Remove the bulb and discard the water by letting it drain out.
5. Rinse the pipette by drawing several millilitres of solution from the beaker into it. Rotate and rock the pipette to coat the inner surface with solution. Discard the rinse. Rinse the pipette twice in this way. It is now ready to fill with standard solution.

Filling the Pipette

6. Place the tip of the pipette below the surface of the solution.
7. Hold the suction bulb loosely on the end of the glass stem. Use the suction bulb to draw liquid up just past the etched volume mark. (See Figure 10.16.)
8. As quickly and smoothly as you can, slide the bulb off and place your index finger over the end of the glass stem.
9. Gently roll your finger slightly away from end of the stem to let solution drain slowly out.
10. When the bottom of the meniscus aligns with the etched mark, as in Figure 10.17, 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 drop. See Figure 10.18. The measured volume inside the pipette is now ready to transfer to an Erlenmeyer flask or a volumetric flask.

Transferring the Solution

12. Place the tip of the pipette against the inside glass wall of the flask. Let the solution drain slowly, by removing your finger from the stem.
13. After the solution drains, wait several seconds, then touch the tip to the inside wall of the flask to remove any drop on the end. Note: You may notice a small amount of liquid remaining in the tip. The pipette was calibrated to retain this amount. Do not try to remove it.

Figure 10.16 Draw a bit more liquid than you need into the pipette. It is easier to reduce this volume than it is to add more solution to the pipettes.

Figure 10.17 The bottom of the meniscus must align exactly with the etched mark.

Adding the Indicator

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

Rinsing the Burette

A burette is used to accurately measure the volume of liquid added during a titration experiment. It is a graduated glass tube with a tap at one end.

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 inner surfaces.

17. Hold the burette over a sink. Open the tap, and let the water drain out. While you do this, check that the tap does not leak. Make sure that it turns smoothly and easily.

18. Rinse the burette with 5 mL to 10 mL of the solution that will be measured. Remember to open the tap to rinse the lower portion of the burette. Rinse the burette twice, discarding the liquid each time.

TITRATION TIP

If you are right-handed, the tap should be on your right as you face the burette. Use your left hand to operate the tap. Use your right hand to swirl the liquid in the Erlenmeyer flask. If you are left-handed, reverse this arrangement.

Filling the Burette

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

20. With the tap closed, add solution until the liquid is above the zero mark. Remove the funnel. Carefully open the tap. Drain the liquid into a 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 drop. Check that the portion of the burette that is below the tap is filled with liquid and contains no air bubbles.

22. Record the initial burette reading in your notebook.

23. Replace the beaker with the Erlenmeyer flask that you prepared earlier. Place a sheet of white paper under the Erlenmeyer to help you see the indicator colour change that will occur near the end-point.

TITRATION TIP

Near the end-point, when you see the indicator change colour as liquid enters the flask from the burette, slow the addition of liquid. The end-point can occur very quickly.

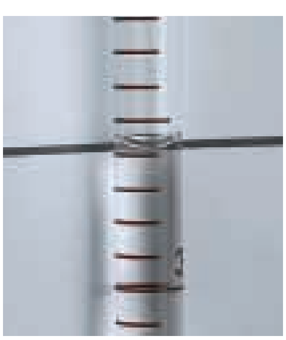

Figure 10.20 A meniscus reader helps you read the volume of liquid in the burette more easily

Reading the Burette

24. A meniscus reader is a small white card with a thick black line on it. Hold the card behind the burette, with the black line just under the meniscus, as in Figure 10.20. Record the volume added from the burette to the nearest 0.05 mL.

TITRATION TIP

Observe the level of solution in the burette so that your eye is level with the bottom of the meniscus.

SKILL FOCUS

Predicting

Performing and recording

Analyzing and interpreting

Investigation 10-B

The Concentration of Acetic Acid in Vinegar

Vinegar is a dilute solution of acetic acid, CH_3COOH. Only the hydrogen atom that is attached to an oxygen atom is acidic. Thus, acetic acid is monoprotic. As a consumer, you can buy vinegar with different concentrations. For example, the concentration of table vinegar is different from the concentration of the vinegar that is used for pickling foods. To maintain consistency and quality, manufacturers of vinegar need to determine the percent concentration of acetic acid in the vinegar. In this investigation, you will determine the concentration of acetic acid in a sample of vinegar.

Prediction

Which do you predict has the greater concentration of acetic acid: table vinegar or pickling vinegar? Give reasons for your prediction.

Materials

pipette
suction bulb
retort stand
burette
burette clamp
3 beakers (250 mL)
3 Erlenmeyer flasks (250 mL)
labels
meniscus reader
sheet of white paper
funnel
table vinegar
pickling vinegar
sodium hydroxide solution
distilled water
dropper bottle containing phenolphthalein

Safety Precautions

Both vinegar and sodium hydroxide solutions are corrosive. Wash any spills on skin or clothing with plenty of water. Inform your teacher immediately.

Procedure

1. Record the following information in your notebook. Your teacher will tell you the concentration of the sodium hydroxide solution.
   - concentration of $NaOH_{(aq)}$ (in mol/L)
   - type of vinegar solution
   - volume of pipette (in mL)
2. Copy the table below into your notebook, to record your observations.

Burette Readings for the Titration of Acetic Acid

Reading (mL)	Trial 1	Trial 2	Trial 3
final reading			
initial reading			
volume added			

3. Label a clean, dry beaker for each liquid: $NaOH_{(aq)}$, vinegar, and distilled water. Obtain each liquid. Record the type of vinegar you will be testing.
4. Obtain a pipette and a suction bulb. Record the volume of the pipette for trial 1. Rinse it with distilled water, and then with vinegar.
5. Pipette some vinegar into the first Erlenmeyer flask. Record this amount. Add approximately 50 mL of water. Also add two or three drops of phenolphthalein indicator.
6. Set up a retort stand, burette clamp, burette, and funnel. Rinse the burette first with distilled water. Then rinse it with sodium hydroxide solution. Make sure that there are no air bubbles in the burette. Also make sure

that the liquid fills the tube below the glass tap. Remove the funnel before beginning the titration.

7. Place a sheet of white paper under the Erlenmeyer flask. Titrate sodium hydroxide into the Erlenmeyer flask while swirling the contents. The end-point of the titration is reached when a permanent pale pink colour appears. If you are not sure whether you have reached the end-point, take the burette reading. Add one drop of sodium hydroxide, or part of a drop. Observe the colour of the solution. If you go past the end-point, the solution will become quite pink.
8. Repeat the titration twice more. Record your results for each of these trials.
9. When you have finished all three trials, dispose of the chemicals as directed by your teacher. Rinse the pipette and burette with distilled water. Leave the burette tap open.

Analysis

1. Average the two closest burette readings. Average all three readings if they agree within about ±0.2 mL.
2. Write the chemical equation for the reaction of acetic acid with sodium hydroxide.
3. Calculate the concentration of acetic acid in your vinegar sample. Use the average volume and concentration of sodium hydroxide, and the volume of vinegar.
4. Find the molar mass of acetic acid. Then calculate the mass of acid in the volume of vinegar you used.
5. The density of vinegar is 1.01 g/mL. (The density of the more concentrated vinegar solution is greater than the density of the less concentrated solution. You can ignore the difference, however.) Calculate the mass of the vinegar sample. Find the percent by mass of acetic acid in the sample.

Conclusions

6. Compare your results with the results of other students who used the same type of vinegar. Then compare the concentration of acetic acid in table vinegar with the concentration in pickling vinegar. How did your results compare with your prediction?
7. List several possible sources of error in this investigation.

Application

8. Most shampoos are basic. Why do some people rinse their hair with vinegar after washing it?

Section Wrap-up

In this section, as in much of Unit 3, you combined liquid solutes and liquid solvents. You have learned how to describe the concentration of ions that determine the acidic or basic nature of a solution. As well, you performed calculations to determine the concentration of acids and many other substances in solution. In the upcoming unit, you will apply your understanding of stoichiometry and solutions by examining the nature and interactions of substances in the gaseous state.

Section Review

1 K/U Write a generalized word equation to describe what happens during a neutralization reaction.

2 K/U Write a chemical equation for each neutralization reaction.

(a) KOH with HNO_3

(b) HBr with $Ca(OH)_2$

(c) H_3PO_4 with NaOH

(d) $Mg(OH)_2$ (the active ingredient in milk of magnesia, an antacid) with HCl (the acid in your stomach)

3 K/U Distinguish between the equivalence point and the end-point in a titration. Why might they be different? How would this affect the result of a titration?

4 I A 25.0 mL sample of sulfuric acid is completely neutralized by adding 32.8 mL of 0.116 mol/L ammonia solution. Ammonium sulfate, $(NH_4)_2SO_4$, and water are formed. What is the concentration of the sulfuric acid?

5 I The following data were collected during a titration. Calculate the concentration of the sodium hydroxide solution.

Titration Data

Volume of $HCl_{(aq)}$	10.00 mL
Final volume of $NaOH_{(aq)}$	23.08 mL
Initial volume of $NaOH_{(aq)}$	1.06 mL
Concentration of $HCl_{(aq)}$	0.235 mol/L

6 I You should always put the two solutions for a titration experiment in clean, dry beakers. You do not need to dry the Erlenmeyer flask to which you add the solutions, however, if it has been thoroughly rinsed with distilled water. Explain the difference in these procedures.

7 C Suppose that a laboratory technician accidentally spills a dilute solution of a strong acid on her hands, the sleeve of her lab coat, and the laboratory bench. Explain how she would deal with this spill.