8.5 Acid–Base Reactions

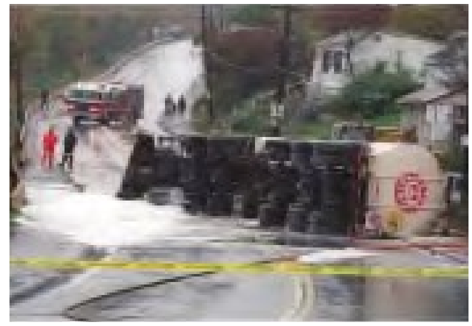

Figure 1
The first thing the emergency crew must do is identify the spilled chemical.

Acids take part in several characteristic reactions, including their reaction with bases. This means that we can sometimes get clues about an unknown substance by observing how it reacts, and what the products of the reaction are. For example, imagine an emergency response team arriving at the site of a traffic accident involving a chemical spill (**Figure 1**). At first team members are not sure what the spilled chemical is, but then they notice that it seems to be reacting with the magnesium/aluminum wheels of a car. They know that all acids react with active metals (e.g., $Mg_{(s)}$, $Zn_{(s)}$, and $Al_{(s)}$) to produce hydrogen gas, so it is likely that the spilled chemical is an acid. They could confirm their suspicions with a strip of litmus paper or by doing a quick test of a sample of the gas being produced. The acid/metal reaction is a single displacement reaction in which the hydrogen in the acid behaves like a metal.

$$\text{active metal} + \text{acid} \rightarrow H_{2(g)} + \text{ionic compound}$$

For example:

$$Mg_{(s)} + 2\ HCl_{(aq)} \rightarrow H_{2(g)} + MgCl_{2(aq)}$$

All acids also react at various rates with carbonates (e.g., $Na_2CO_{3(aq)}$ and $CaCO_{3(s)}$) in a double displacement reaction producing carbonic acid, $H_2CO_{3(aq)}$. To clean up spills, emergency response teams often use a mixture of chemicals that includes sodium carbonate. The carbonic acid produced is unstable and decomposes quickly to form carbon dioxide and water.

$$\begin{aligned}\text{acid} + \text{carbonate} &\rightarrow H_2CO_{3(aq)} + \text{ionic compound}\\ &\rightarrow CO_{2(g)} + H_2O_{(l)} + \text{ionic compound}\end{aligned}$$

For example:

$$2\ HNO_{3(aq)} + Na_2CO_{3(s)} \rightarrow CO_{2(g)} + H_2O_{(l)} + 2\ NaNO_{3(aq)}$$

Finally, acids react with bases in another double displacement reaction often called a neutralization reaction. The result of this reaction is that the pH of the solution moves toward seven. Emergency response teams might test the pH of a spill with pH paper and then monitor the neutralization of the spill by repeated tests with pH paper.

$$\text{strong acid} + \text{strong base} \rightarrow \text{water} + \text{ionic compound}$$

For example:

$$HCl_{(aq)} + KOH_{(aq)} \rightarrow H_2O_{(l)} + KCl_{(aq)}$$
$$H^+_{(aq)} + Cl^-_{(aq)} + K^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)} + K^+_{(aq)} + Cl^-_{(aq)}\ \text{(total ionic equation)}$$

$$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}\ \text{(net ionic equation)}$$
or $$H_3O^+_{(aq)} + OH^-_{(aq)} \rightarrow 2\ H_2O_{(l)}\ \text{(net ionic equation)}$$

Although not a defining property of acids, acids also undergo precipitation reactions with some ionic compounds. Instead of a double displacement neutralization reaction, a double displacement precipitation reaction occurs. For example, we can determine the concentration of hydroiodic acid by reacting the solution with aqueous lead(II) nitrate and measuring the mass of precipitate formed.

acid + ionic compound → precipitate + acid

For example:

$$2\ HI_{(aq)} + Pb(NO_3)_{2(aq)} \rightarrow PbI_{2(s)} + 2\ HNO_{3(aq)}$$

Practice

Understanding Concepts

1. List three chemical-reaction properties that are characteristic of acids.
2. (a) A spill of hydrobromic acid can be neutralized by reacting it with zinc, lye, and/or washing soda. Write chemical equations to represent each of these reactions individually. (Common names are given in Appendix C.)
   (b) List advantages and/or disadvantages for each of the neutralization methods above.
3. Oxalic acid, $H_2C_2O_{4(aq)}$, found in foods such as rhubarb, undergoes typical acid reactions. Write chemical equations to represent the following reactions.
   (a) Oxalic acid reacts with an aluminum cooking pot to make it look shinier. Assume that one product is an oxalate (low solubility).
   (b) Oxalic acid reacts with aqueous calcium, say, calcium chloride in your blood, to produce insoluble crystals of calcium oxalate that can grow to become bladder or kidney stones.
   (c) The high oxalic acid content in spinach reacts with the high iron content in spinach to precipitate iron (III) ions out during digestion as, for example, iron(III) oxalate. Assume that the iron in the stomach appears as iron(III) chloride. How does this destroy the myth of Popeye getting his strength from iron in spinach?

Figure 2
An initial volume reading is taken before any titrant is added to the sample solution. Then titrant is added until the reaction is complete, that is, when a drop of titrant changes the colour of the sample. The final buret reading is then taken. The difference in buret readings is the volume of titrant added.

Acid–Base Titration

titration: a laboratory procedure involving the carefully measured and controlled addition of a solution from a buret into a measured volume of a sample solution

titrant: the solution in the buret during a titration

Titration is a common laboratory technique used to determine the concentration of substances in solution. It is a reliable, efficient, and economical technology that is simple to use. A known volume of the sample to be analyzed is usually transferred into an Erlenmeyer flask (**Figure 2**). The solution in the buret (called the **titrant**) is added, drop by drop, to the sample. Alternatively, the standard solution could be in the flask, so the solution of unknown concentration would be the titrant. The titrant is added drop by drop until the reaction is judged to be complete. To help us identify this point, we select an indicator that changes colour when the reaction is complete (**Table 1**). The point at which the indicator

Table 1: Indicator Colour Change as the Endpoint of Titration

Indicator	Acidic	Basic
litmus	red	blue
methyl orange	red	yellow
bromothymol blue	yellow	blue
phenolphthalein	colourless	red

changes colour is called the **endpoint**. This is at, or close to, the point at which the titrant and sample have completely reacted.

Reactions between aqueous reactants are generally fast. If this reaction involves acids and bases, and at least one of these is strong, then the reaction will normally proceed as in a balanced chemical equation (be stoichiometric), require no special conditions (be spontaneous), and be complete (quantitative). These are the necessary requirements for the use of titration for chemical analysis. Typical chemical analyses include analysis of acids in the environment (acid deposition studies), quality control in industrial and commercial operations, and scientific research. A typical practice titration is the chemical analysis of acetic acid in a sample of vinegar, using a sodium hydroxide solution in a buret as the titrant.

When you perform a chemical analysis by titration, you will use a number of volumetric techniques such as using a pipet to transfer portions of the sample for analysis, the titration using a buret, and measuring solution volumes. In order to obtain precise and reliable results, you must know the concentration of one of the reactants; that is, you must use a **standard solution**.

When doing a titration, you come to a point when the reaction is complete and the indicator suddenly changes colour: the endpoint. At the endpoint you stop the titration and record the volume of titrant used. Chemically equivalent amounts of reactants, as determined by the mole ratio in the balanced chemical equation, have now been combined.

A titration procedure should involve several trials, using different samples of the unknown solution to improve the reliability of the answer. A typical requirement is to repeat measurements until three trials result in volumes within 0.1 mL to 0.2 mL of each other. These three results are then averaged before carrying out further calculations.

endpoint: the point in a titration at which a sharp change in a property occurs (e.g., a colour change)

standard solution: a solution of precisely and accurately known concentration

SUMMARY Titration Requirements

For titration, a chemical reaction must be

- spontaneous—chemicals react on their own without a continuous addition of energy
- fast—chemicals react instantaneously when mixed
- quantitative—the reaction is more than 99% complete
- stoichiometric—there is a single, whole number mole ratio of amounts of reactants and products

Example: Titration of Hydrochloric Acid

Manufacturers of commercial chemicals must ensure that their products meet certain standards. Quality control technicians are responsible for checking samples of product to ensure that they are acceptable. For aqueous solutions of acids such as muriatic acid (hydrochloric acid), the concentration of the product must be within certain limits. Titration is an excellent technique to test concentration. A sodium carbonate solution can be used as the reagent to analyze the hydrochloric acid. Suppose 1.59 g of anhydrous sodium carbonate, $Na_2CO_{3(s)}$, is dissolved to make 100.0 mL of a standard solution. Samples (10.00 mL) of this standard solution are then taken and titrated with the $HCl_{(aq)}$ product, which has been diluted by a factor of 10. The titration evidence collected is shown in **Table 2** (page 396).

Table 2: The Titration of $Na_2CO_{3(aq)}$ with $HCl_{(aq)}$

Trial	**1**	**2**	**3**	**4**
final buret reading (mL)	13.3	26.0	38.8	13.4
initial buret reading (mL)	0.2	13.3	26.0	0.6
volume of $HCl_{(aq)}$ added (mL)	13.1	12.7	12.8	12.8

To analyze this evidence you first need to calculate the molar concentration of the sodium carbonate solution.

$v_{Na_2CO_3} = 100\ \text{mL}$

$M_{Na_2CO_3} = 105.99\ \text{g/mol}$

$m_{Na_2CO_3} = 1.59\ \text{g}$

$$n_{Na_2CO_3} = 1.59\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{105.99\ \cancel{\text{g}}}$$

$$= 0.0150\ \text{mol}$$

$$C_{Na_2CO_3} = \frac{0.0150\ \text{mol}}{0.10000\ \text{L}}$$

$$C_{Na_2CO_3} = 0.150\ \text{mol/L}$$

Now you can start the stoichiometry procedure by writing the balanced chemical equation. Notice in **Table 2** that four trials were done, and the volume in the first trial is significantly higher than in the others. The volume you use for $HCl_{(aq)}$ should be your best average, typically three results within ±0.1 mL of each other. The value, 12.8 mL, is the average of trials 2, 3, and 4. (Keep the unrounded value in your calculator as usual.) The rest of the stoichiometry procedure follows the usual steps.

$2\ HCl_{(aq)}$	+	$Na_2CO_{3(aq)}$	$\rightarrow$	$H_2CO_{3(aq)}$	+	$2\ NaCl_{(aq)}$
12.8 mL		10.00 mL				
C		0.150 mol/L				

Remember that we are using only 10.00 mL of the Na_2CO_3 solution for each trial, not 100 mL.

$$n_{Na_2CO_3} = 10.00\ \text{m}\cancel{\text{L}} \times \frac{0.150\ \text{mol}}{1\ \cancel{\text{L}}}$$

$$= 1.50\ \text{mmol}$$

$$n_{HCl} = 1.50\ \text{mmol} \times \frac{2}{1}$$

$$= 3.00\ \text{mmol}$$

$$C_{HCl} = \frac{3.00\ \cancel{\text{m}}\text{mol}}{12.8\ \cancel{\text{m}}\text{L}}$$

$$C_{HCl} = 0.234\ \text{mol/L}$$

Alternatively, we could combine these steps into one calculation as shown below.

$$C_{HCl} = 10.00\ \cancel{\text{mL}\ Na_2CO_3} \times \frac{0.150\ \cancel{\text{mol}\ Na_2CO_3}}{1\ \text{L}\ \cancel{Na_2CO_3}} \times \frac{2\ \text{mol HCl}}{1\ \cancel{\text{mol}\ Na_2CO_3}} \times \frac{1}{12.8\ \cancel{\text{mL}}}$$

$$C_{HCl} = 0.234\ \text{mol/L}$$

Since the sample of muriatic acid had been diluted by a factor of 10, the original concentration of hydrochloric acid must be 10 times greater, or 2.35 mol/L.

Sample Problem 1

An acid rain sample containing sulfurous acid was analyzed in a laboratory using a titration with a standard solution of sodium hydroxide. Use the evidence given in **Table 3** to determine the concentration of the sulfurous acid.

Table 3: Titration of 25.0 mL of $H_2SO_{3(aq)}$ with 0.105 mol/L $NaOH_{(aq)}$

Trial	1	2	3
final buret reading (mL)	11.1	21.7	32.4
initial buret reading (mL)	0.3	11.1	21.7
volume of $NaOH_{(aq)}$ added	10.8	10.6	10.7

Solution

$$\begin{array}{lllllllll} H_2SO_{3(aq)} & + & 2\ NaOH_{(aq)} & \rightarrow & Na_2SO_{3(aq)} & + & 2\ H_2O_{(l)} \\ 25.0\ \text{mL} & & 10.7\ \text{mL} & & & & \\ C & & 0.105\ \text{mol/L} & & & & \end{array}$$

$$n_{NaOH} = 10.7\ \text{m}\cancel{\text{L}} \times \frac{0.105\ \text{mol}}{1\ \cancel{\text{L}}}$$

$$= 1.12\ \text{mmol}$$

$$n_{H_2SO_3} = 1.12\ \text{mmol} \times \frac{1}{2}$$

$$= 0.562\ \text{mmol}$$

$$C_{H_2SO_3} = \frac{0.562\ \cancel{\text{m}}\text{mol}}{25.0\ \cancel{\text{m}}\text{L}}$$

$$C_{H_2SO_3} = 0.0225\ \text{mol/L}$$

or

$$C_{H_2SO_3} = 10.7\ \cancel{\text{mL}}\ \cancel{\text{NaOH}} \times \frac{0.105\ \cancel{\text{mol}}\ \ \cancel{\text{NaOH}}}{1\ \text{L}\ \cancel{\text{NaOH}}} \times \frac{1\ \text{mol}\ H_2SO_3}{2\ \cancel{\text{mol}}\ \ \cancel{\text{NaOH}}} \times \frac{1}{25.0\ \cancel{\text{mL}}}$$

$$C_{H_2SO_3} = 0.0225\ \text{mol/L}$$

The concentration of the sulfurous acid in the sample is 0.0225 mol/L or 22.5 mmol/L.

INQUIRY SKILLS

- ○ Questioning
- ○ Hypothesizing
- ○ Predicting
- ○ Planning
- ● Conducting
- ● Recording
- ● Analyzing
- ● Evaluating
- ● Communicating

Investigation 8.5.1

Titration Analysis of Vinegar

Consumer products are required by law to have the minimum quantity of the active ingredient listed on the product label. Companies that produce chemical products usually employ analytical chemists and technicians to monitor the final product in a process known as quality control. Government consumer affairs departments also use chemists and technicians to check products, particularly in response to consumer complaints.

In this investigation, you will be the quality control chemist. You have received a report that a local high-school cafeteria has been serving watered-down vinegar to the students. Your purpose is to test the acetic acid concentration of the vinegar to discover whether it has been diluted (i.e., below the 5.0% W/V acetic acid indicated on the purchased container). Complete the **Analysis** and **Evaluation** sections of the report.

Question

What is the molar concentration of acetic acid in a sample of vinegar?

Prediction

The manufacturer claims on the label that the vinegar contains 5.0% acetic acid, which translates into a 0.87 mol/L concentration of acetic acid. The concentration of acetic acid in the vinegar sample should be the same.

Experimental Design

A sample of vinegar from the school cafeteria is diluted by a factor of 10 to make a 100.0-mL solution. The diluted solution is titrated with a standard sodium hydroxide solution using phenolphthalein as the indicator.

Materials

Wear eye protection and a lab apron.

At these dilutions, the chemicals are fairly safe and can be disposed of down the drain.

lab apron
eye protection
$NaOH_{(aq)}$
vinegar
phenolphthalein
wash bottle of pure water
two 100-mL or 150-mL beakers
250-mL beaker.
100-mL volumetric flask with stopper
50-mL buret
10-mL volumetric pipet
pipet bulb
ring stand
buret clamp
stirring rod
small funnel
two 250-mL Erlenmeyer flasks
meniscus finder

Procedure

1. Obtain about 30 mL of vinegar in a clean, dry 100-mL beaker.
2. Pipet one 10.00-mL portion into a clean 100-mL volumetric flask and dilute to the mark.
3. Stopper and invert several times to mix thoroughly.
4. Obtain about 70 mL of $NaOH_{(aq)}$ in a clean, dry, labelled 100-mL beaker.
5. Set up the buret with $NaOH_{(aq)}$, following the accepted procedure for rinsing and clearing the air bubble.
6. Pipet a 10.00-mL sample of diluted vinegar into a clean Erlenmeyer flask.
7. Add 1 or 2 drops of phenolphthalein indicator.
8. Record the initial buret reading to the nearest 0.1 mL.
9. Titrate the sample with $NaOH_{(aq)}$ until a single drop produces a permanent change from colourless to faint pink.
10. Record the final buret reading to the nearest 0.1 mL.
11. Repeat steps 6 to 10 until three consistent results are obtained.

Analysis

(a) Answer the Question: What is the molar concentration of acetic acid in a sample of vinegar?

Evaluation

(b) Evaluate your evidence: How confident are you that your techniques and measurements resulted in good evidence?

(c) Evaluate the Prediction: Assuming the manufacturer's claim is accurate, is someone in the cafeteria diluting the vinegar? Include an accuracy calculation (percentage difference) in your evaluation.

Practice

Understanding Concepts

4. Briefly describe three types of characteristic reactions of acids.
5. What are the four reaction requirements in order to use a reaction in a titration in a chemical analysis?
6. What are the two reactants in a titration, and what equipment is used to contain them?
7. What is a standard solution?
8. Why are several trials usually done in a titration?

Applying Inquiry Skills

9. Analysis shows that 9.44 mL of 0.0506 mol/L $KOH_{(aq)}$ is needed for the titration of 10.00 mL of a water sample taken from an acidic lake. Determine the molar concentration of acid in the lake water, assuming that the acid is sulfuric acid.

Answer

9. 0.0239 mol/L or 23.9 mmol/L

Answers

10. (b) 1.08 mol/L

11. (b) 2.66 mol/L

12. 0.278 mol/L

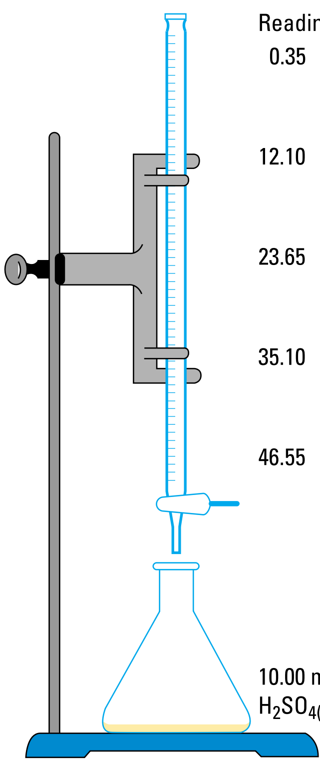

Figure 3
Sodium hydroxide titrant is added to samples of sulfuric acid in successive trials.

10. Solutions of oxalic acid, $H_2C_2O_{4(aq)}$, have many applications. Like $H_2SO_{4(aq)}$, oxalic acid reacts in a 2:1 mole ratio with sodium hydroxide. Complete the **Evidence**, **Analysis**, and **Evaluation** sections of the following investigation report.

Question
What is the concentration of oxalic acid in a rust-removing solution?

Prediction
The oxalic acid solution is labelled as 10% W/V, or 1.11 mol/L.

Experimental Design
The original oxalic acid solution (rust remover) is diluted by a factor of 100, that is, 10.00 mL to 1000 mL. The concentration of dilute oxalic acid solution is determined by titration with a sodium hydroxide solution.

Evidence
(a) Copy and complete **Table 4.**

Table 4: Volume of 0.0161 mol/L Sodium Hydroxide Required to Neutralize 10.00 mL of Diluted Oxalic Acid

Trial	**1**	**2**	**3**	**4**
Final buret reading (mL)	14.3	27.8	41.1	13.8
Initial buret reading (mL)	0.2	14.3	27.8	0.4
Volume of $NaOH_{(aq)}$ used (mL)				

Analysis
(b) Using the Evidence in **Table 5**, calculate the concentration of oxalic acid in the rust remover.

Evaluation
(c) Evaluate the Prediction: Is the manufacturer's label accurate?

11. Complete the **Evidence** and **Analysis** for the following titration.

Question
What is the molar concentration of the hydrochloric acid in a solution of kettle-scale remover?

Experimental Design
The hydrochloric acid in a solution of kettle-scale remover is titrated with a standardized solution of barium hydroxide. The colour change of bromothymol blue indicator (from blue to green) is the endpoint.

Evidence
(a) Copy and complete **Table 5.**

Table 5: Titration of 10.00-mL Samples of $HCl_{(aq)}$ with 0.974 mol/L $Ba(OH)_{2(aq)}$

Trial	**1**	**2**	**3**	**4**
final buret reading (mL)	15.6	29.3	43.0	14.8
initial buret reading (mL)	0.6	15.6	29.3	1.2
volume of $Ba(OH)_{2(aq)}$ added (mL)				
colour at endpoint	blue	green	green	green

Analysis
(b) Using the Evidence in **Table 5**, calculate the concentration of the hydrochloric acid in the kettle-scale remover.

12. Samples of sulfuric acid were titrated with 0.484 mol/L sodium hydroxide. The evidence is shown in **Figure 3**. Calculate the concentration of the sulfuric acid solution.