Chapter

8

Acids and Bases

In this chapter, you will be able to

- experimentally determine the empirical properties, including pH, of acids and bases;
- design and conduct an experiment to determine the effect of dilution on pH;
- explain acids and bases, using both Arrhenius and Brønsted-Lowry theories;
- describe and explain the difference between strong and weak acids and bases;
- use the terms: ionization, dissociation, strong acid/base, weak acid/base, hydronium ion, proton transfer, conjugate acid/base, titration, titrant, and endpoint;
- write balanced chemical equations for reactions involving acids and bases;
- develop the skills involved in titration and solve stoichiometry problems using titration evidence;
- describe examples of solutions for which the concentration must be known and exact.

Acid indigestion, commercial antacid remedies for indigestion, pH-balanced shampoos—you don't have to look far in a drugstore to find labels referring to acids or acidity. Many people think that all acids are corrosive, and therefore dangerous, because solutions of acids react with many substances. Yet boric acid is used as an eyewash. Can this be as dangerous as it sounds?

References in the popular media offer no insight into what acids and bases are, or what they do. In fact, such references usually emphasize only one perspective, such as the environmental damage caused by an acid or the cleaning power of a base. As a result, popular ideas are often confusing. An amateur gardener who has just read an article describing the destruction of conifer forests by acid rain may be puzzled by instructions on a package of evergreen fertilizer stating that evergreens are acid-loving plants (**Figure 1**).

This chapter takes a historical approach, presenting evidence and following the development of theories about the substances we call acids and bases. These theories are used to explain and predict the behaviour of acids and bases.

Reflect on your Learning

1. What are some properties of acids?
2. How can you explain these properties of acids?
3. What are some properties of bases?
4. How can you explain these properties of bases?
5. How do your explanations in questions 2 and 4 account for the evidence that acids react with bases?

Try This **Activity**

Consumer Products

Look at home or in a store and read the labels on a variety of cleaning products such as drain, oven, wall, floor, window, and toilet bowl cleaners. Find the lists of ingredients as well as any caution notes.

For each product:

(a) Record the product name and the list of ingredients.
(b) Underline the ingredients on the list that you think are active. Give reasons for your choices.
(c) Classify as many of the active ingredients as you can as acids or bases.
(d) Record any warnings about mixing the product with other substances.
(e) Referring to your list in (d), state which combinations represent mixtures of acids and bases.

Figure 1
All gardeners know that conifers like acidic soil, so why is acid rain so damaging?

8.1 Understanding Acids and Bases

The story of acids and bases is progressive: It is unfolding like a detective story. Our understanding of acids and bases has changed over time as we have extended our concepts to include more and more acids and bases. Early scientists described acids, for example, as compounds that produce hydrogen gas when reacted with an active metal. These scientists realized that acids (at least, some acids) contain hydrogen atoms. To other investigators, acids were substances that contain oxygen and that react with limestone to produce carbon dioxide.

Later, investigators discovered that acids in aqueous solution change blue litmus to red and conduct electricity. These observations did not fit with the earlier definitions of acids. A new explanation was needed.

Acids and bases share some properties with molecular and ionic substances, and have some properties that are unique.

Table 1 shows that pure liquid samples of both ionic compounds and bases conduct electricity. Similarly, aqueous solutions of both ionic compounds and bases conduct electricity. In some way, bases seem to be similar to ionic compounds. What is it about ionic compounds that enables them to conduct electricity? According to Arrhenius, ionic compounds separate into ions when they are liquid or in solution. Can we assume that the same is true of bases? Perhaps they, too, separate into ions. The presence of mobile ions would explain the conductivity.

Table 1: Properties of Pure and Aqueous Substances

Substance	Conductivity			Litmus
	Solid	Liquid	Aqueous	Aqueous
most molecular compounds	no	no	no	no effect
most ionic compounds	no	yes	yes	no effect
acids	no	no	yes	blue to red
bases	no	yes	yes	red to blue

Do all bases release the same ion in aqueous solution? Many bases are ionic compounds that contain a hydroxide, OH^-, which could be released in solution. It seems likely that this ion gives a base its characteristics. After all, sodium chloride, NaCl, is not a base but sodium hydroxide, NaOH, is. Going by the evidence we have at this stage, we can conclude that bases are ionic hydroxides that release mobile hydroxide ions in solution.

$$NaOH_{(s)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$$
$$Ca(OH)_{2(s)} \rightarrow Ca^{2+}_{(aq)} + 2\ OH^-_{(aq)}$$

Table 1 also shows that molecular substances, including acids, do not conduct electricity in their pure states. However, acids (unlike other molecular substances) become conductors when dissolved in water. Arrhenius explained that molecular substances do not conduct electricity because they contain only electrically neutral particles called molecules. Can we conclude, therefore, that acids in their pure forms contain neutral molecules and not ions, but that acids in solution contain ions? This is certainly what the evidence suggests. It seems that acids are somehow different in structure and/or composition from other molecular substances.

DID YOU KNOW ?

Liquid Ionic Compounds

How hot does an ionic compound have to be, before we can test it for conductivity? Because of the strong ionic bonds, it has to be heated to extremely high temperatures (700 –1000°C). When the ionic bonds are broken, the ions are free to move around and the substance can conduct electricity.

By studying the composition of substances that turn blue litmus red in an aqueous solution, scientists found that acids seem to contain hydrogen atoms. This led scientists to write the chemical formula for acids as $HA_{(aq)}$ (**Table 2**). The electrical conductivity of these acidic solutions led to the theory that acids ionize in water to release hydrogen ions, $H^+_{(aq)}$. Acids, then, according to the evidence, are hydrogen-containing compounds that ionize in water to produce hydrogen ions.

Table 2: Properties of Hydrogen Compounds

Empirical formula	Litmus test	Molecular formula
$CH_2O_{3(aq)}$	blue to red	H_2CO_3
$CH_{4(aq)}$	no change	CH_4
$SH_{2(aq)}$	blue to red	H_2S
$PH_{3(aq)}$	no change	PH_3
$NHO_{3(aq)}$	blue to red	HNO_3

$$HCl_{(g)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$$
$$HNO_{3(aq)} \rightarrow H^+_{(aq)} + NO^-_{3(aq)}$$

Arrhenius extended his theory of ions to explain some of the properties of acids and bases. According to Arrhenius, we can write an equation showing that **bases** dissociate into individual positive and negative ions in solution. He proposed that aqueous hydroxide ions were responsible for the properties of basic solutions, such as turning red litmus paper blue. The **dissociation** of bases is similar to that of any other ionic compound, as shown in the following dissociation equation for barium hydroxide.

$$Ba(OH)_{2(s)} \rightarrow Ba^{2+}_{(aq)} + 2\ OH^-_{(aq)}$$

According to the evidence in **Table 1**, acids are electrolytes in solution even though as pure substances, they are molecular compounds. Acids, such as $HCl_{(g)}$ and $H_2SO_{4(l)}$, do not show their acidic properties until they dissolve in water. Since acids in solution are electrolytes, Arrhenius' theory suggests that acid solutions must contain ions. However, the pure solute is molecular; it is made up only of neutral molecules. How, then, can its solution contain ions? Obviously, **acids** do not simply dissolve to form a solution of molecules. According to Arrhenius, after acids dissolve as individual molecules, they then ionize into hydrogen ions and negative ions in solution.

In the case of acids, Arrhenius assumed that the water somehow causes the acid molecules to ionize, but he didn't propose an explanation for this. (We now believe that water molecules help to pull the molecules apart—to ionize the acid.) A typical example of an acid is hydrogen chloride gas dissolving in water to form hydrochloric acid. We can describe this process with an **ionization** equation.

$$HCl_{(g)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$$

So, although HCl is a molecular compound, it appears to behave in solution as if it were ionic. It ionizes into ions, which are capable of conducting electricity in solution. We explain the properties of acids by saying that all acids produce hydrogen ions in solution, and define acids as substances that ionize in water to increase the hydrogen ion concentration.

DID YOU KNOW ?

Formulas of Acids

The 19th century idea was that acids are salts (compounds) of hydrogen. For example, scientists would say that $HCl_{(aq)}$ is the hydrogen salt of $NaCl_{(aq)}$. This led to the practice of writing hydrogen first in the formulas of substances known to form acidic solutions, such as $H_2SO_{4(aq)}$ and $HCl_{(aq)}$.

base: (according to the Arrhenius theory) an ionic hydroxide that dissociates in water to produce hydroxide ions

dissociation: the separation of ions that occurs when an ionic compound dissolves in water

acid: (according to the Arrhenius theory) a compound that ionizes in water to form hydrogen ions

ionization: any process by which a neutral atom or molecule is converted into an ion

The Arrhenius theory was a major advance in understanding chemical substances and solutions. Arrhenius also provided the first comprehensive theory of acids and bases. The empirical and theoretical definitions of acids and bases are summarized in **Table 3**.

Table 3: Acids, Bases, and Neutral Substances

Type of substance	Empirical definition	Theoretical definition
acids	• in solution, turn blue litmus red • are electrolytes • in solution, neutralize bases	• these hydrogen-containing compounds ionize to produce $H^+_{(aq)}$ ions • $H^+_{(aq)}$ ions react with $OH^-_{(aq)}$ ions to produce water
bases	• in solution, turn red litmus blue • are electrolytes • in solution, neutralize acids	• ionic hydroxides dissociate to produce $OH^-_{(aq)}$ ions • $OH^-_{(aq)}$ ions react with $H^+_{(aq)}$ ions to produce water
neutral substances	• in solution, do not affect litmus • some are electrolytes • some are nonelectrolytes	• no $H^+_{(aq)}$ or $OH^-_{(aq)}$ ions are formed • some exist as ions in solution • some exist as molecules in solution

Sample Problem 1

Write dissociation or ionization equations (as appropriate) for the dissolving of the following chemicals in water. Label each equation as either dissociation or ionization.

(a) potassium chloride (a salt substitute)

(b) hydroiodic acid (a strong acid)

Solution

(a) $KCl_{(s)} \rightarrow K^+_{(aq)} + Cl^-_{(aq)}$ (dissociation)

(b) $HI_{(aq)} \rightarrow H^+_{(aq)} + I^-_{(aq)}$ (ionization)

As Sample Problem 1 shows, ionic substances dissociate in water, but acids ionize.

Acids as pure substances are molecular and, as such, may be solids, liquids, or gases at SATP (standard ambient temperature and pressure). When you are writing ionization equations for acids, you may not always know the initial state of matter. If you do know the state, use (s), (l), or (g) subscripts; if you do not know the pure state of the acid, use (aq). This is correct for now, as all the acids you will be using in this course are in aqueous solution. For example,

$$HCl_{(g)} \text{ or } HCl_{(aq)} \rightarrow H^+_{(aq)} + Cl^+_{(aq)} \quad \text{(a gaseous acid)}$$
$$HC_2H_3O_{2(l)} \text{ or } HC_2H_3O_{2(aq)} \rightarrow H^+_{(aq)} + C_2H_3O^-_{2(aq)} \quad \text{(a liquid acid)}$$
$$H_2C_2O_{4(s)} \text{ or } H_2C_2O_{4(aq)} \rightarrow H^+_{(aq)} + HC_2O^-_{4(aq)} \quad \text{(a solid acid)}$$

Recall that chemicals that turn blue litmus red have acid formulas that begin with H.

Strong and Weak Acids

Different acidic solutions have different electrical conductivity. We can see this from laboratory evidence (**Figure 1**). There seem to be two fairly distinctive classes of acids. If we were to test the electrical conductivity of a variety of acids at equal concentration and temperature, we might collect evidence such as that in **Table 4**. If we were then to analyze the evidence, we could classify the acids according to the acid strength, as is shown in the last column.

Figure 1
In solutions of equal concentration, hydrochloric acid is a very good conductor of electricity; acetic acid conducts electricity less well.

Table 4: Electrical Conductivity and Strength of Various Acids

Acid name	Acid formula	Electrical conductivity	Strength
acetic acid	$HC_2H_3O_{2(aq)}$	low	weak
nitrous acid	$HNO_{2(aq)}$	low	weak
carbonic acid	$H_2CO_{3(aq)}$	low	weak
hydrochloric acid	$HCl_{(aq)}$	high	strong
sulfuric acid	$H_2SO_{4(aq)}$	high	strong
nitric acid	$HNO_{3(aq)}$	high	strong

strong acid: (theoretical definition) an acid that ionizes almost completely (>99%) in water to form aqueous hydrogen ions

percentage ionization: the percentage of molecules that form ions in solution

weak acid: (empirical definition) an acid with characteristic properties less than those of a strong acid; (theoretical definition) an acid that ionizes only partially (<50%) in water to form aqueous hydrogen ions, so exists primarily in the form of molecules

We can classify all acids as either strong or weak. Acids whose solutions have high electrical conductivity are called **strong acids**. Their high electrical conductivity is explained by their high **percentage ionization**. Most (more than 99%) of the acid molecules ionize. Sulfuric acid, nitric acid, and hydrochloric acid are examples of strong acids. (These strong acids are among those identified in the table of Concentrated Reagents in Appendix C, page 636.) Most other common acids, such as carbonic acid, are **weak acids**. Their low conductivity is explained by their very low percentage ionization.

You have probably heard that some acids are dangerous. How do you know which ones to treat with particular caution? For safety purposes you need to pay more attention to the strength of the acid than the concentration. A dilute solution of a strong acid can be more dangerous than a concentrated solution of a weak acid. This is because the corrosive property of acids is due to the hydrogen ion. The more hydrogen ions in solution, the more dangerous the solution. As a general rule, show respect for all acids, but especially strong acids.

To communicate the percentage ionization and the strength of an acid undergoing ionization, we can write the percentage ionization over the chemical equation arrow. Strong acids generally ionize >99%, while weak acids generally ionize <50%.

$$HNO_{3(aq)} \xrightarrow{>99\%} H^+_{(aq)} + NO^-_{3(aq)} \qquad \text{(a strong acid)}$$

$$HC_2H_3O_{2(aq)} \xrightarrow{<50\%} H^+_{(aq)} + C_2H_3O^-_{2(aq)} \qquad \text{(a weak acid)}$$

Sample Problem 2

The following acidic solutions were tested (at equal concentration and temperature) for electrical conductivity. Write ionization equations to explain the relative conductivity of each acid.

(a) hydrobromic acid (aqueous hydrogen bromide): high conductivity

(b) hydrofluoric acid (aqueous hydrogen fluoride): low conductivity

Solution

(a) $HBr_{(aq)} \xrightarrow{>99\%} H^+_{(aq)} + Br^-_{(aq)}$

(b) $HF_{(aq)} \xrightarrow{<50\%} H^+_{(aq)} + F^-_{(aq)}$

SUMMARY Acids and Bases

At this point, the only strong acids that you have to know about are hydrochloric acid, nitric acid, and sulfuric acid. You can assume that all other acids are weak, unless you are told otherwise.

You can also assume that all bases are ionic hydroxides, all of which are strong bases.

Strong Acid: $HA_{(s/l/g/aq)} \xrightarrow{>99\%} H^+_{(aq)} + A^-_{(aq)}$ (>99% ionized)

Weak Acid: $HA_{(s/l/g/aq)} \xrightarrow{<50\%} H^+_{(aq)} + A^-_{(aq)}$ (<50% ionized)

Strong Base: $MOH_{(s)} \longrightarrow M^+_{(aq)} + OH^-_{(aq)}$ (100% dissociated as ions)

Practice

Understanding Concepts

1. What evidence is there that ionic compounds exist as ions in their pure state while molecular compounds, including acids, exist as molecules in their pure state?
2. Based upon their chemical formulas, classify the following chemicals as acid, base, or neutral.
   (a) $H_2SO_{3(aq)}$
   (b) $NaOH_{(aq)}$
   (c) $CH_3OH_{(aq)}$
   (d) $HC_3H_5O_{2(aq)}$
   (e) $NaC_2H_3O_{2(aq)}$
   (f) $Ba(OH)_{2(aq)}$
3. Acids are molecular compounds, but they don't behave quite like other molecular compounds. What properties make acids unique?
4. According to the Arrhenius theory, what causes the change in colour of litmus paper in a basic solution? in an acidic solution?
5. Write an empirical and a theoretical definition of an acid.
6. For each of the following compounds, indicate whether they dissociate or ionize in aqueous solution. Write ionic equations to represent the dissociation or the ionization.
   (a) sodium hydroxide (drain cleaner)
   (b) hydrogen acetate (vinegar)
   (c) hydrogen sulfate (battery acid)
   (d) calcium hydroxide (slaked lime)

Applying Inquiry Skills

7. Complete the **Analysis** and **Evaluation** for the following investigation.

Question

Which of the chemicals, numbered 1 to 7, is $KCl_{(s)}$, $Ba(OH)_{2(s)}$, $Zn_{(s)}$, $HC_7H_5O_{2(s)}$, $Ca_3(PO_4)_{2(s)}$, $C_{25}H_{52(s)}$ (paraffin wax), and $C_{12}H_{22}O_{11(s)}$?

Experimental Design

Equal amounts of each chemical are added to equal volumes of water. The chemicals are tested for solubility, and their aqueous solutions are tested for their conductivity and effect on litmus paper.

Evidence

Table 5: Properties of Seven Substances

Chemical	Solubility in water	Conductivity of solution	Effect of solution on litmus paper
1	high	none	no change
2	high	high	no change
3	none	none	no change
4	high	high	red to blue
5	none	none	no change
6	none	none	no change
7	low	low	blue to red

Analysis

(a) Based on the Evidence (**Table 5**), which chemical is which?

Evaluation

(b) Use your knowledge of chemicals to suggest improvements to the Experimental Design.

Making Connections

8. What kinds of compounds can be used in solution to conduct electricity in batteries? Using the Internet, find examples (or illustrations) of batteries that use each kind of compound.

   Follow the links for Nelson Chemistry 11, 8.1.

   GO TO www.science.nelson.com

9. Which of the following acids should be handled particularly carefully? (Use the table of Concentrated Reagents in Appendix C.) Give your reasons.
   (a) hydrochloric acid (concrete etching)
   (b) carbonic acid (carbonated beverages)
   (c) sulfuric acid (car battery acid)
   (d) acetic acid (vinegar)
   (e) nitric acid (copper etching)
   (f) phosphoric acid (rust remover)

10. Pure water is a molecular compound. Can you be electrocuted if you are standing in pure water? Provide your reasoning.

8.2 pH of a Solution

Skin care and hair care products are often advertised as being pH balanced. What does this mean? It sounds like a good thing, but what is pH, and how can it be balanced?

pH is a way of indicating the concentration of hydrogen ions present in a solution. You have just discovered that all acids release H^+ ions when they ionize in water. And you know that we can define concentration as the amount of a substance (in moles) present in a given volume of a solution. Chemists have combined these concepts into a way of communicating the acidity of a solution: a concise code for the concentration of $H^+_{(aq)}$ ions. This code is the pH scale.

The molar concentration of hydrogen ions is extremely important in chemistry. According to Arrhenius's theory, hydrogen ions are responsible for the properties of acids, and the higher the concentration of hydrogen ions, the more acidic a solution will be. Similarly, the higher the concentration of hydroxide ions, the more basic a solution will be. You might not expect a neutral solution or pure water to contain any hydrogen or hydroxide ions at all. However, careful testing yields evidence that even pure water always contains tiny amounts of both hydrogen and hydroxide ions, due to a slight ionization of the water molecules (**Figure 1**). In a sample of pure water, about two of every billion water molecules ionize to form hydrogen and hydroxide ions.

$$H_2O_{(l)} \rightarrow H^+_{(aq)} + OH^-_{(aq)}$$

This gives a hydrogen ion concentration of about 1×10^{-7} mol/L. Most conductivity tests will show no conductivity for pure water (unless the equipment is extremely sensitive). The concentration of hydrogen ions declines when a base is dissolved in the water. The hydroxide ions released by the base react with the hydrogen ions freed by the ionization of water to produce water molecules. The result is a decline in the concentration of hydrogen ions.

Evaluation

(b) Use your knowledge of chemicals to suggest improvements to the Experimental Design.

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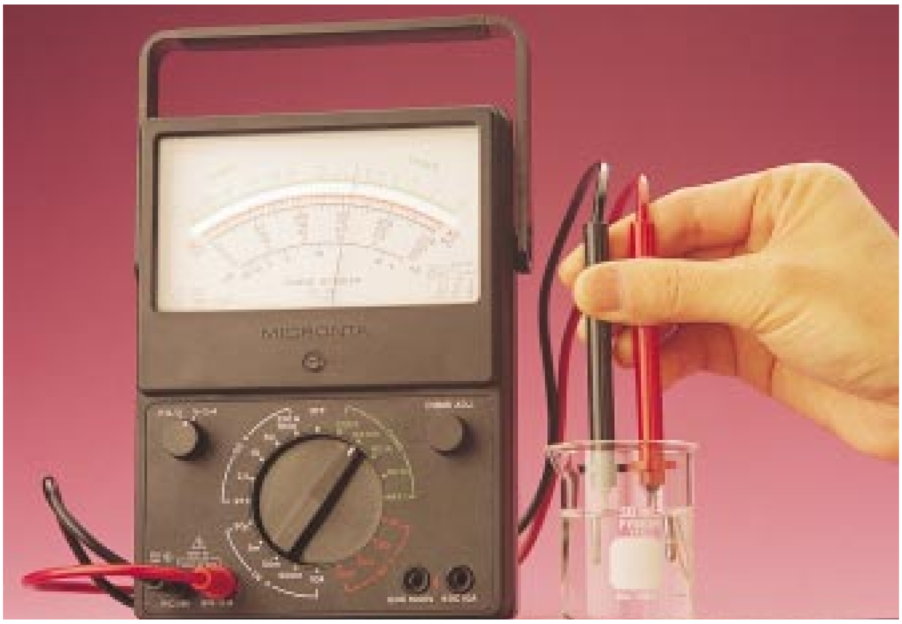

Figure 1
Pure distilled water has a very slight electrical conductivity that is only noticeable when tested with a very sensitive meter.

Aqueous solutions show a phenomenally wide range of hydrogen ion concentrations—from more than 10 mol/L for a concentrated hydrochloric acid solution, to less than 1×10^{-15} mol/L for a concentrated sodium hydroxide solution. Any aqueous solution can be classified as acidic, neutral, or basic using a scale based on the hydrogen ion concentration. Note that the square brackets around the $H^{+}_{(aq)}$ ion indicate "molar concentration."

- In a neutral solution, $[H^{+}_{(aq)}] = 1 \times 10^{-7}$ mol/L.
- In an acidic solution, $[H^{+}_{(aq)}] > 1 \times 10^{-7}$ mol/L.
- In a basic solution, $[H^{+}_{(aq)}] < 1 \times 10^{-7}$ mol/L.

The extremely wide range of hydrogen ion concentrations led to a convenient shorthand method of communicating these concentrations. This method, called pH, was invented in 1909 by Danish chemist Søren Sørenson. The pH of a solution is defined as the negative of the exponent to the base 10 of the hydrogen ion concentration (expressed as moles per litre). This is not quite as complicated as it sounds. For example, a solution with a hydrogen concentration of 10^{-7} mol/L has a pH of 7 (neutral). Similarly, a pH of 2 corresponds to a much higher hydrogen ion concentration of 10^{-2} mol/L (acidic).

We can rearrange this relationship to show that pH is the negative of the power of 10 of the hydrogen ion molar concentration.

$$[H^{+}_{(aq)}] = 10^{-pH}$$

This relationship can be used, without complicated mathematics, to convert between pH and the hydrogen ion concentration. For example, if the hydrogen ion concentration is 1×10^{-5} mol/L, then the pH is 5.0. If the pH is 8.0, then the hydrogen ion concentration is 1×10^{-8} mol/L.

Notice, in **Table 1**, that the certainty (as expressed in the number of significant digits) of the hydrogen ion concentration provides the precision (as expressed in the number of decimal places) of the pH. The number of decimal places for the pH is equal to the number of significant digits in the hydrogen ion concentration. This is because the integer of the pH (e.g., 7) does not count as a significant digit any more than the exponent in 10^{-7} (i.e., 7) does.

DID YOU KNOW ?

The "p" in pH

pH was developed only about 100 years ago, but already the origin of the term has become blurred. Some scientists associate pH with **p**ower of **h**ydrogen H; others with **p**otential of **h**ydrogen. Sørenson was Danish, so perhaps the "p" in pH comes from the Danish word "potenz," meaning "strength," or the French word "potentiel." It is strange that we have so quickly lost the origin of such a familiar term.

Table 1: Sample Conversions Between Hydrogen Ion Concentration and pH

$[H^{+}_{(aq)}]$ (mol/L)	pH
10^{-9}	9
1×10^{-2}	2.0
1.0×10^{-7}	7.00
1.00×10^{-11}	11.000

Sample Problem 1

What is the pH of each of the following solutions?

(a) 1×10^{-2} mol/L hydrogen ion concentration in vinegar

(b) $[H^+_{(aq)}] = 1.0 \times 10^{-12}$ mol/L in household ammonia

Solution

(a) pH = 2.0

(b) pH = 12.00

Sample Problem 2

What is the hydrogen ion concentration for the following solutions?

(a) a carbonated beverage with a pH of 3.0

(b) an antacid solution for which pH = 10.00

Solution

(a) $[H^+_{(aq)}] = 1 \times 10^{-3}$ mol/L

(b) $[H^+_{(aq)}] = 1.0 \times 10^{-10}$ mol/L

pH is specified on the labels of consumer products such as shampoos, in water-quality tests for pools and aquariums, in environmental studies of acid rain, and in laboratory investigations of acids and bases. Since each pH unit corresponds to a factor of 10 in the concentration, a huge $[H^+_{(aq)}]$ range can now be communicated by a simple set of positive numbers (**Figure 2**). In these applications, dilution is often an important consideration. What happens to the pH when a solution is diluted? When the pH changes by one unit (e.g., from 5 to 6), the hydrogen ion concentration has been decreased by a factor of ten (i.e., from 10^{-5} mol/L to 10^{-6} mol/L). When the hydrogen ion concentration changes by a factor of 100 (e.g., from 10^{-2} mol/L to 10^{-4} mol/L), the pH changes by 2 units (i.e., from 2 to 4).

neutralization: a reaction between an acid and a base that results in a pH closer to 7

Another way of changing the pH is by a **neutralization** reaction, in which hydrogen ions react with hydroxide ions to move the pH closer to 7. Removing hydrogen ions from the solution is a more effective method of changing the pH than diluting the solution. For this reason, spills of acids and bases are more often neutralized than diluted. Diluting the solution helps to make the solution less hazardous, but neutralizing the solution is more effective.

$$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$$

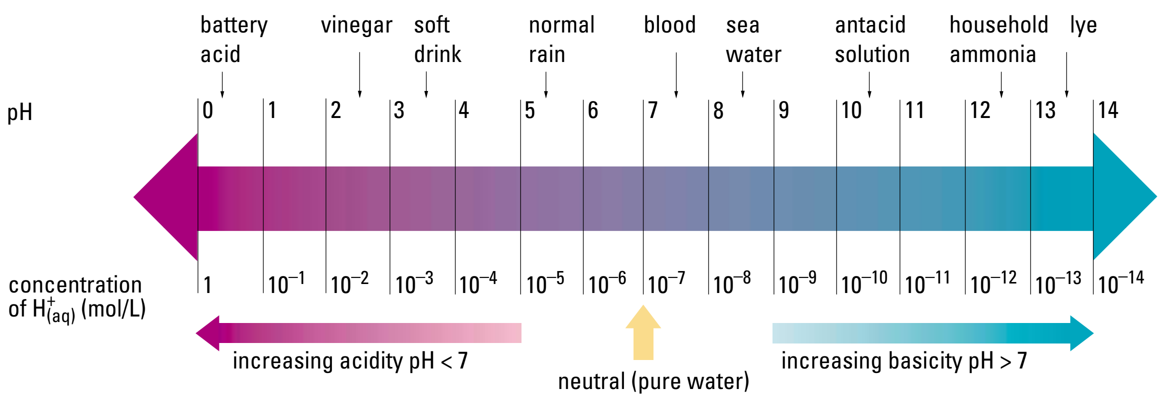

Figure 2
The pH scale can communicate a broad range of hydrogen ion concentrations in a wide variety of substances.

Practice

Understanding Concepts

1. State four examples of products for which the pH may be specified.
2. What is the hydrogen ion concentration in the following household solutions?
   (a) household ammonia: pH = 11.00
   (b) vinegar: pH = 2.0
   (c) soda pop: pH = 4.00
   (d) drain cleaner: pH = 14.00
3. Express the following typical concentrations as pH values.
   (a) grapefruit juice: $[H^+_{(aq)}] = 1.0 \times 10^{-3}$ mol/L
   (b) rainwater: $[H^+_{(aq)}] = 1 \times 10^{-5}$ mol/L
   (c) milk: $[H^+_{(aq)}] = 1 \times 10^{-7}$ mol/L
   (d) liquid soap: $[H^+_{(aq)}] = 1.0 \times 10^{-10}$ mol/L
4. If a water sample test shows a pH of 5, by what factor would the hydrogen ion concentration have to be changed to make the sample neutral? Is this an increase or a decrease in hydrogen ion concentration?
5. Explain why, if the hydrogen ion concentration is 1 mol/L, pH = 0.
6. What amount of hydrogen ions, in moles, is present in 100 L of the following solutions?
   (a) wine: $[H^+_{(aq)}] = 1 \times 10^{-3}$ mol/L
   (b) seawater: pH = 8.00
   (c) stomach acid: $[H^+_{(aq)}] = 10.0$ mmol/L

Reflecting

7. Many chemicals that are potentially toxic or harmful to the environment have maximum allowable concentration levels set by government legislation.
   (a) If the chemical is dangerous, should the limit be zero?
   (b) Is a zero level theoretically possible?
   (c) Is a zero level measurable?
   (d) If a nonzero limit is set, in your opinion, how should this limit be chosen?

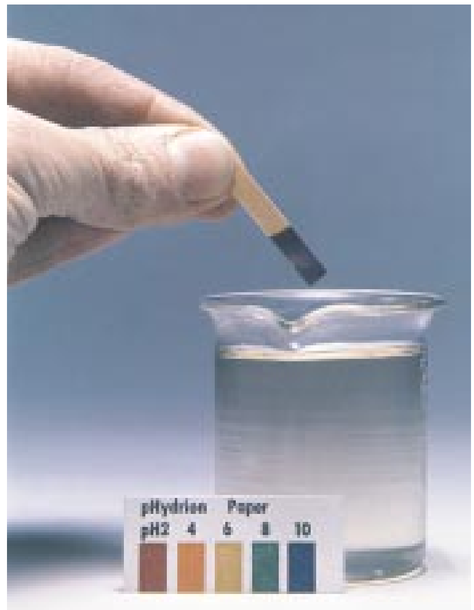

Figure 3
pH paper has a range of possible colours. Each colour corresponds to a particular pH.

Measuring pH of a Solution

You may have already measured the pH of solutions using pH indicators or pH paper (**Figure 3**) to estimate the approximate value. In some situations, such as testing garden soil or aquarium water, this is appropriate. However, in scientific analysis we usually require more precise measurement. For example, when environmental scientists and technicians study the effects of acid rain in waterways, small pH changes in a large body of water can be significant. Precise pH measurements are normally made using a pH meter. All pH meters operate like small electric cells in which the electricity produced by the cell depends on the acidity of the solution. With modern electronics, the tiny electrical signals are detected, converted, and displayed on a screen or dial as the pH. Some pH meters, particularly older ones, are small desktop units (**Figure 4**); some newer meters can be just as precise, yet are as small as a pencil.

Even more recent pH meters consist of a pH probe connected to a computer. The main advantage of this technology is that the pH can be sampled, recorded, and even graphed continuously and automatically. However, if you are just going to take a few pH readings, a system without a computer is perfectly adequate.

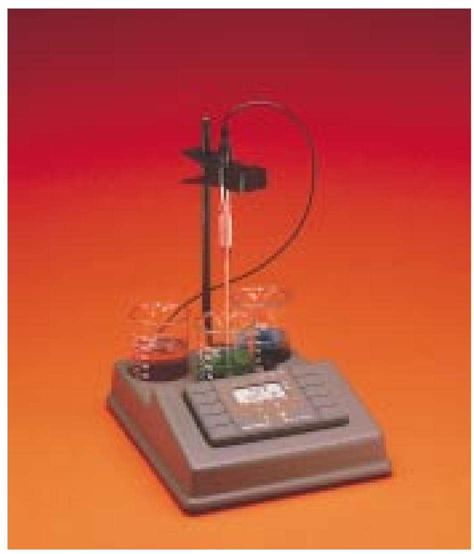

Figure 4
Arnold Beckman invented the pH meter in 1935, 26 years after Søren Sørenson had developed the concept of pH for communicating hydrogen ion concentration.

INQUIRY SKILLS

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

Figure 5
Unfortunately, it is all too common in Canada for industrial chemicals to flow directly into local waterways.

Figure 6
High concentrations of PCBs have been found in the fat of the beluga whales in the St. Lawrence.

Acidic solutions are corrosive. Wear eye protection and a laboratory apron. Keep your hands away from your eyes and wash your hands when finished.

Investigation 8.2.1

Dilution and pH

For many years it has been common practice for some municipalities, industries, businesses, and consumers to dispose of hazardous wastes in lakes and rivers (**Figure 5**). This is often justified by an argument based on the dilution of the wastes by a relatively large volume of water. Some critics refer to this argument as "the solution to pollution is dilution." There are many reasons why this reasoning is not valid. One reason is the environmental effect of bioamplification, in which initially minute quantities of contaminants are concentrated to dangerous quantities higher up the food chain (**Figure 6**). But there may also be other effects that are not immediately obvious. The purpose of this investigation is to study the effect of dilution on the pH of a solution and develop a generalization for any trends observed.

In previous work (see Section 6.5), you practised preparing a solution by a careful dilution of a starting solution. You can use the same skills and equipment to plan and carry out an experiment to answer the following Question. Complete the **Experimental Design**, **Materials**, **Procedure**, and **Analysis** sections of the report.

Question

What effect does the dilution of an acidic solution have on the pH of the solution?

Experimental Design

The pH of a 0.10 mol/L $HCl_{(aq)}$ solution is measured using a pH meter or pH paper. A sample of this solution is precisely diluted and the pH (dependent variable) is measured. This process is repeated several times to obtain measures of pH for a range of concentrations (independent variable). The temperature of the solution is a controlled variable.

(a) Write a detailed Procedure for this investigation, including any safety precautions.
(b) Draw up a table to record your observations.

Materials

(c) Create a list of Materials.

Procedure

1. Obtain your teacher's approval, and then conduct your investigation.

Analysis

(d) Answer the Question. Based on your Evidence, what relationship exists between concentration and pH?

Calculating pH and Hydrogen Ion Concentration

Earlier in this section you saw that a hydrogen ion concentration can be expressed as a simple pH. The relationship between $[H^+_{(aq)}]$ and pH is easy to calculate if the concentration is only a power of ten, such as 10^{-3} mol/L where the pH = 3, or $[H^+_{(aq)}] = 1 \times 10^{-7}$ mol/L where the pH is 7.0. This is clearly illustrated in **Figure 2**, page 370. What if the concentration of hydrogen ions is not so simple, such as 2.7×10^{-3} mol/L? To answer this question you need to know that, in mathematics, the logarithm of a number is the exponent when the number is written in exponential form. For example, ignoring certainty in significant digits,

$$100 = 10^2 \qquad \log_{10}(10^2) = 2$$
$$0.001 = 10^{-3} \qquad \log_{10}(10^{-3}) = -3$$

In general, if $y = 10^x$, then $\log_{10}(y) = x$. Fortunately, all scientific calculators have a function key, labelled "log," that will find the logarithm (exponent) for any number in the display of the calculator (**Figure 7**). Because hydrogen ion concentrations are usually less than 1 mol/L, **pH** is defined as the negative logarithm to avoid having almost all pHs as negative numbers. This is simply a convenience agreed to by all scientists—a convention of communication. On your calculator and in general usage, log is normally understood to mean $\log_{10}$. Therefore, the mathematical definition of pH becomes,

pH: a measure of the acidity of a solution; the negative logarithm, to the base ten, of the hydrogen ion molar concentration

$$\text{pH} = -\log[H^+_{(aq)}]$$

`EXP` `+/–` `log`

Figure 7
You will become familiar with these keys when converting from hydrogen ion concentration to pH.

Values of pH can be calculated from the hydrogen ion concentration, as shown in the following example. The digits preceding the decimal point in a pH value are determined by the digits in the exponent of the hydrogen ion concentration. These digits locate the position of the decimal point in the concentration value and do not indicate the certainty of the value. However, *the number of digits following the decimal point in the pH value is equal to the number of significant digits expressing the certainty of the hydrogen ion concentration.* For example, a hydrogen ion concentration of 2.7×10^{-3} mol/L (note the two significant digits) corresponds to a pH of 2.57 (note the two decimal places). The 2.7 in the 2.7×10^{-3} mol/L value indicates the certainty as two significant digits. The 3 in the 2.7×10^{-3} mol/L value only indicates where the decimal point goes. The .57 in the pH value of 2.57 communicates a certainty of two significant digits. The 2 in the 2.57 pH value does not count as measured digits; it only indicates where the decimal place in the value goes, i.e., the power of 10 of the value. The examples below will help to clarify this for you.

Sample Problem 3

An antacid solution has a hydrogen ion concentration of 4.7×10^{-11} mol/L. What is its pH? (See **Figure 8**.)

`4` `•` `7` `EXP`
`1` `1` `+/–` `log` `+/–`

Figure 8
On many calculators, $-\log(4.7 \times 10^{-11})$ may be entered by pushing the above sequence of keys.

Solution

$$\text{pH} = -\log[H^+_{(aq)}]$$
$$= -\log(4.7 \times 10^{-11})$$
$$\text{pH} = 10.33$$

(Note that the certainty of both values is expressed as two significant digits.)

If pH is measured in an acid–base analysis, you may have to convert a pH reading to the molar concentration of hydrogen ions. This conversion is based on the mathematical concept that a base ten logarithm represents an exponent. Therefore, the pH becomes the exponent.

$$[H^+_{(aq)}] = 10^{-\text{pH}}$$

The method of calculating the hydrogen ion concentration from the pH value is shown in Sample Problem 4.

Sample Problem 4

The pH reading of a solution is 10.33. What is its hydrogen ion concentration? Be sure to indicate your answer with the correct certainty (**Figure 9**).

$$[H^+_{(aq)}] = 10^{-pH}$$
$$= 10^{-10.33} \text{ mol/L}$$
$$[H^+_{(aq)}] = 4.7 \times 10^{-11} \text{ mol/L}$$

(Note that the two decimal places in the pH yield two significant digits in the hydrogen ion concentration. The 10 is not counted in the pH for the same reason that the 11 is not counted in the hydrogen ion concentration: Both the 10 and the 11 only tell you where the decimal place goes—the power of ten.)

1 0 • 3 3

+/− either INV log

or 2nd log

Figure 9
On many calculators, $10^{-10.33}$ may be entered by pushing the sequence of keys shown above.

SUMMARY Hydrogen Ion Concentration and pH

pH is the negative power of ten of the hydrogen ion concentration.

$$pH = -\log[H^+_{(aq)}] \qquad \text{or} \qquad [H^+_{(aq)}] = 10^{-pH}$$

solution:	acidic	neutral	basic
$[H^+_{(aq)}]$:	$>10^{-7}$	10^{-7}	$<10^{-7}$
pH:	<7	7	>7

Note the inverse relationship between $[H^+_{(aq)}]$ and pH. The higher the hydrogen ion molar concentration, the lower the pH.

Practice

Understanding Concepts

8. What are two ways used to measure the pH of a solution?
9. (a) What is the pH of pure water?
   (b) What is the hydrogen ion concentration of pure water?
10. Food scientists and dietitians measure the pH of foods when they devise recipes and special diets.
    (a) Copy and complete **Table 2**.
    (b) Based on pH only, which of the foods should taste the most sour?

Making Connections

11. What are some benefits and risks of using acidic and basic substances in your home? When do you consider the benefits to exceed the risks? When do the risks exceed the benefits? Provide some examples of each.

Answers

10. (a) 2.26; 4×10^{-9} mol/L; 4.6×10^{-4} mol/L; 3.4

Table 2: Acidity of Foods

Food	$[H^+_{aq}]$ (mol/L)	pH
oranges	5.5×10^{-3}	
asparagus		8.4
olives		3.34
blackberries	4×10^{-4}	

Sections 8.1–8.2 Questions

Understanding Concepts

1. A household cleaner has a pH of 12 and some fruit juice has a pH of 3.
   (a) What is the hydrogen ion concentration in each solution?
   (b) Compare the concentration of hydrogen ions in the fruit juice to that of the hydrogen ions in the cleaner. How many times more concentrated is the hydrogen ion in the juice than the cleaner?
2. What is the pH of each of the following water samples?
   (a) tap water: $[H^+_{aq}] = 1 \times 10^{-8}$ mol/L
   (b) pure water: $[H^+_{aq}] = 1 \times 10^{-7}$ mol/L
   (c) normal rainwater: $[H^+_{aq}] = 2.5 \times 10^{-6}$ mol/L
   (d) acid rain: $[H^+_{aq}] = 1.3 \times 10^{-4}$ mol/L
3. Common household vinegar has a pH of 2.4 and some pickling vinegar has a pH of 2.2.
   (a) Which vinegar solution is more acidic?
   (b) Which has a greater hydrogen ion concentration?
   (c) Calculate the hydrogen ion concentration in each solution.
4. A student tested the pH of 0.1 mol/L solutions of hydrochloric acid and acetic acid. The pH of the hydrochloric acid solution was 1.1 and the pH of the acetic acid was 2.9.
   (a) Explain the difference in pH.
   (b) Communicate the difference by writing ionization equations for each of the acids.
   (c) Which of the solutions deserves greater caution when being used? Why?
5. The pH of a cleaning solution was determined using a variety of technologies. Convert the pH into a molar concentration of hydrogen ions, with the correct certainty in the answer. Suggest what technology might have been used in each of these measurements.
   (a) pH = 10
   (b) pH = 9.8
   (c) pH = 9.84
   (d) pH = 9.836
6. Hydrangeas are garden shrubs that may produce blue, purple, or pink flowers. Research has indicated that the colour is dependent on the pH of the soil: blue at pH 5.0–5.5, purple at pH 5.5–6.0, and pink at pH 6.0–6.5. Convert the following expressions of acidity from pH to $[H^+_{(aq)}]$.
   (a) pH = 5.4 (blue)
   (b) pH = 5.72 (purple)

   Convert these concentrations to pH and predict the colour of the flower.
   (c) $[H^+_{(aq)}] = 5 \times 10^{-7}$ mol/L
   (d) $[H^+_{(aq)}] = 7.9 \times 10^{-6}$ mol/L

Applying Inquiry Skills

7. Write an Experimental Design to determine which of six provided acids are strong and which are weak.

(continued)

8. Design and conduct an investigation to test the concept that the pH of a weak acid changes predictably by dilution. Complete every section of the report: Prediction, Experimental Design, Materials, Procedure, Evidence, Analysis, and Evaluation. Before conducting the investigation, have your chemistry teacher authorize your design, materials, and procedure.

   Question
   What will be the increase in the pH of vinegar when a 1.0 mL sample is diluted 100 times?

9. A scientist wants to determine the pH of several toothpastes but must add water to the pastes in order to measure the pH with a pH meter probe. Critique this experimental design by supporting and defending the design and by suggesting an alternative design.

Making Connections

10. Use the Internet or other resources to find the effect of soil pH on plants. Report this information in a short Did You Know? column for a gardening magazine.

    Follow the links for Nelson Chemistry 11, 8.2.

    GO TO www.science.nelson.com

11. Large restaurants give vegetables such as lettuce an acid wash before serving. Why would they follow this practice and which acid are they most likely to use? Provide your reasoning.

Reflecting

12. The term “weak” is sometimes applied to dilute solutions. Why do we have to be careful not to use “weak” in this context, when discussing acids?

Figure 1
Working with solutions requires a variety of specialized glassware.

8.3 Working with Solutions

Everyone deals with solutions on a daily basis. Tap water, soft drinks, air, gasoline, and alloys such as the “gold” used in jewellery are just a few common examples of homogeneous mixtures called solutions. Many foods contain acidic solutions. Cleaning solutions often contain bases. Sometimes we prepare solutions, for example, when we dissolve flavour crystals and sugar in water to make an inexpensive drink, or make a cup of tea, or dilute a concentrated cleaner in water. In most cases we pay some attention to concentration, although it may be measured only roughly, and adjust it to our requirements.

Some people regularly handle solutions as a part of their work. The kinds of solutions vary tremendously, depending on the occupation. However, the preparation and use of solutions is generally more technologically advanced and precise than we would encounter at home (**Figure 1**).

Practice

Applying Inquiry Skills

1. State the name and use of each piece of apparatus in **Figure 1**.

8. Design and conduct an investigation to test the concept that the pH of a weak acid changes predictably by dilution. Complete every section of the report: Prediction, Experimental Design, Materials, Procedure, Evidence, Analysis, and Evaluation. Before conducting the investigation, have your chemistry teacher authorize your design, materials, and procedure.

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Some people regularly handle solutions as a part of their work. The kinds of solutions vary tremendously, depending on the occupation. However, the preparation and use of solutions is generally more technologically advanced and precise than we would encounter at home (**Figure 1**).

Practice

Applying Inquiry Skills

1. State the name and use of each piece of apparatus in **Figure 1**.

CAREER

Career Solutions

The training requirements for careers that involve solutions vary from high school chemistry for a job as a tree-planter to a Ph.D. degree in chemistry for a career in pure research chemistry.

Water-Quality Analyst

A water-quality analyst or technician in a water treatment plant works with aqueous solutions every day. Physical and chemical tests are routinely done to determine the treatment of the raw water and to monitor the quality of the final treated water. Many chemical tests (such as the analysis of dissolved iron ions, calcium ions, and chlorine) require the preparation of other reagent (reactant) solutions to conduct the tests. Both solution preparation and reactions in solution are important parts of the job of a water-quality analyst.

Chemistry Teacher

A chemistry or physical science teacher must have a knowledge of solutions, and be able to handle them safely. In this course and many of your previous science courses, you will have seen and used solutions on many occasions. At most schools, the teachers prepare the solutions that you use, plan the reactions that you do, and sometimes need to be very resourceful in cleaning some stains from glassware by reacting the stain with other chemicals such as acids and bases. Chemistry and other physical science teachers need a good understanding of solutions in order to prepare for lab activities. These teachers are in great demand in schools and universities.

Environmental Chemist

Environmental chemists often specialize in particular aspects of the air, water, or soil. Many of them use solutions as either reactants or samples in chemical analysis. The concentrations of these samples are usually critically important. This career requires a higher degree of chemistry training than technicians and teachers. To be an environmental researcher requires considerable perseverance and optimism as well as an ability to ask questions and design experiments. Some of the research involves understanding the components and processes in the environment and some research may focus on the nature and effects of pollution.

Practice

Making Connections

2. Choose one of the careers discussed that you might be interested in and use the Internet to research this career. What are the specific educational requirements? Does this occupation require certification by some organization? If so, state the organization. What are the job prospects in this area?
Follow the links for Nelson Chemistry 11, 8.3.

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8.4 Acid–Base Theories

Acids and bases can be distinguished by means of a variety of properties (**Table 1**). Some properties of acids and bases are more useful than others to a chemist, especially those that can be used as diagnostic tests, such as the litmus test.

Table 1: Empirical Properties of Acids and Bases

Acids	Bases
sour taste*	taste bitter and feel slippery*
turn blue litmus red	turn red litmus blue
have pH less than 7	have pH greater than 7
neutralize bases	neutralize acids
react with active metals to produce hydrogen gas	
react with carbonates to produce carbon dioxide	

*Note that for reasons of safety it is not appropriate to use taste or touch as diagnostic tests in the laboratory.

Many acids and bases are sold under common or traditional names. As you have learned, concentrated hydrochloric acid is sometimes sold as muriatic acid. Sodium hydroxide, called lye as a pure solid, has a variety of brand names when sold as a concentrated solution for cleaning plugged drains. Generic or "no-name" products often contain the same kind and quantity of active ingredients as brand name products. You can save time, trouble, and money by knowing that, in most cases, the chemical names of compounds used in home products must be given on the label. If you discover that your favourite brand of scale remover is an acetic acid solution, you may be able to substitute vinegar to do the same job less expensively, assuming that the concentrations are similar.

Strong and Weak Acids

Are all acids similar in their reactivity and their pH? Do acidic solutions at the same concentration and temperature possess acidic properties to the same degree? Not surprisingly, the answer to this question is that each acid is unique. The pH of an acid may be only slightly less than 7, or it may be as low as –1. We discussed in Section 8.1 that solutions of strong acids have a much greater conductivity than those of weak acids and the difference can be explained by percentage ionization. As you might suspect, the percentage ionization has an effect on the pH of acid solutions.

We can explain the differences in properties between strong and weak acids using the Arrhenius theory and percentage ionization. For example, hydrogen chloride is a strong acid because it is believed to ionize completely (more than 99%) in water. The high concentration of $H^+_{(aq)}$ ions gives the solution strong acid properties and a low pH.

$$HCl_{(aq)} \xrightarrow{>99\%} H^+_{(aq)} + Cl^-_{(aq)}$$

This means that for each mole of hydrogen chloride dissolved, about one mole of hydrogen ions is produced.

There are relatively few strong acids: Hydrochloric, sulfuric, and nitric acids are the most common.

A weak acid is an acid that ionizes partially in water. Measurements of pH indicate that most weak acids ionize less than 50%. Acetic acid, a common weak acid, is only 1.3% ionized in solution at 25°C and 0.10 mol/L concentration. The relatively low concentration of $H^+_{(aq)}$ ions gives the solution weaker acid properties and a pH closer to 7.

$$HC_2H_3O_{2(aq)} \xrightarrow{1.3\%} H^+_{(aq)} + C_2H_3O^-_{2(aq}$$

For each mole of acetic acid dissolved, only 0.013 mol of $H^+_{(aq)}$ ions is produced.

When we observe chemical reactions involving acids we can see that some acids (such as acetic acid), although they react in the same manner and amount as other acids (such as hydrochloric acid), do not react as quickly. This is why weak acids are generally so much safer to handle, and even to eat or drink, than strong acids. Most of the acids you are likely to encounter are classed as weak acids (**Figure 1**).

The concepts of strong and weak acids were developed to describe, explain, and predict these differences in properties of acids.

Figure 1
Many naturally occurring acids are weak organic (carbon chain) acids. Methanoic (formic) acid is found in the stingers of certain ants, butanoic acid in rancid butter, citric acid in citrus fruits such as lemons and oranges, oxalic acid in tomatoes, and long-chain fatty acids, such as stearic acid, in animal fats.

SUMMARY Properties of Strong and Weak Acids of Equal Concentration

Property	Strong Acid	Weak Acid
pH	$<<7$	<7
Ionization	>99%	<50%
Rate of reaction	fast	slow
Corrosion	fast	slow

Practice

Understanding Concepts

1. Which empirical property listed in **Table 1** is not a diagnostic test used in a chemistry laboratory? Is there a situation where knowledge of this property might be useful?
2. Strong and weak acids can be differentiated by their rates of reaction. Complete and balance the following chemical equations. Predict whether each reaction will be fast or slow.
   (a) $Mg_{(s)} + HCl_{(aq)} \rightarrow$
   (b) $Mg_{(s)} + HC_2H_3O_{2(aq)} \rightarrow$
   (c) $HCl_{(aq)} + CaCO_{3(s)} \rightarrow$
   (d) $HC_2H_3O_{2(aq)} + CaCO_{3(s)} \rightarrow$
3. Which property listed in **Table 1** would be the best to distinguish between strong and weak acids? Justify your choice.
4. What is the theoretical distinction between strong and weak acids?
5. Suppose 100 molecules of a strong acid are dissolved to make a litre of solution and 100 molecules of a weak acid are also dissolved to make a litre of solution in a different container. Assume a 2% ionization for the weak acid.
   (a) What is the concentration of hydrogen ions for each solution, expressed as the number of hydrogen ions per litre?
   (b) How does your answer to (a) explain the difference in properties of strong and weak acids?

Answers

5. (a) 100 H^+/L
   2 H^+/L

Applying Inquiry Skills

6. Bases can also be classified as strong or weak. Predict some differences that you might expect to observe between strong and weak bases. Outline an Experimental Design (including controlled variables) to test your Predictions.
7. Complete the **Experimental Design** and the **Analysis** of the report for the following investigation that determines the relative strength of some acids.

   Question

   What is the order of several common acids in terms of decreasing strength?

   Experimental Design

   (a) Write a description of a design that would produce the Evidence listed in **Table 2**. Include the independent, dependent, and controlled variables in your description.

   Evidence

   Table 2: Acidity of 0.10 mol/L Acids

Acid solution	Formula	pH
hydrochloric acid	$HCl_{(aq)}$	1.00
acetic (ethanoic) acid	$HC_2H_3O_{2(aq)}$	2.89
hydrofluoric acid	$HF_{(aq)}$	2.11
methanoic acid	$HCHO_{2(aq)}$	2.38
nitric acid	$HNO_{3(aq)}$	1.00
hydrocyanic acid	$HCN_{(aq)}$	5.15

   Analysis

   (b) List the acids in decreasing strength.
8. You are given six unlabelled solutions, each containing the same concentration of one of the following six substances: $HCl_{(aq)}$, $HC_2H_3O_{2(aq)}$, $NaCl_{(aq)}$, $C_{12}H_{22}O_{11(aq)}$, $Ba(OH)_{2(aq)}$, and $KOH_{(aq)}$. Your job is to identify each solution. Write an Experimental Design including the specific tests that you would use in your qualitative analysis. You may present your answer as a paragraph, a table of expected evidence, or a flow chart.

Making Connections

9. In the media, especially movies, acids are often portrayed as dangerous, with the ability to "burn through" or "eat away" almost anything. Is this accurate? Justify your answer with personal experience, examples, and explanations. What acids are the most dangerous? How should the media more accurately portray the degree of reactivity of acids?
10. What is acid deposition? What are the typical acids that may be present? Which ones are strong and which are weak? Is it possible to predict which acids have a greater effect in the environment? Why or why not?

    Follow the links for Nelson Chemistry 11, 8.4.

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Reflecting

11. In a 0.01 mol/L solution of an acid, what is the maximum concentration of H^+ ions? What further information would allow you to give a more accurate answer?

The Arrhenius Concept of Acids and Bases

In 1887 Svante Arrhenius created a theory of ions to explain the electrical conductivity of solutions. Arrhenius explained that ionic compounds form solutions that conduct electricity because these compounds dissociate as they dissolve to release an anion and a cation. For example, potassium hydrogen sulfate forms an electrically conductive solution because it dissolves as two ions.

$$KHSO_{4(s)} \rightarrow K^{+}_{(aq)} \text{ or } HSO^{-}_{4(aq)}$$

Scientists had previously agreed that acids were hydrogen compounds. Arrhenius added to this theory by suggesting that acids are hydrogen compounds that ionize to increase the hydrogen ion concentration of a solution. For example, hydrogen chloride gas dissolves in water and ionizes almost completely to increase the hydrogen ion concentration.

$$HCl_{(g)} \rightarrow H^{+}_{(aq)} + Cl^{-}_{(aq)}$$

Arrhenius was also able to explain in a theoretical way why bases have their characteristic properties. He suggested that bases are ionic hydroxides that dissolve in water to increase the hydroxide ion concentration of the solution. For example, potassium hydroxide dissociates in water to increase the hydroxide ion concentration.

$$KOH_{(s)} \rightarrow K^{+}_{(aq)} + OH^{-}_{(aq)}$$

SUMMARY The Arrhenius Theoretical Definitions

acid $\rightarrow H^{+}_{(aq)}$ + anion
base $\rightarrow$ cation + $OH^{-}_{(aq)}$
other ionic compounds $\rightarrow$ ions but no $H^{+}_{(aq)}$ or $OH^{-}_{(aq)}$

Investigation 8.4.1

Testing Arrhenius' Acid–Base Definitions

INQUIRY SKILLS

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

The purpose of this investigation is to test the Arrhenius definitions of acids and bases. As part of this investigation, you are asked to make a Prediction. Do not base your Prediction on experience or a personal hypothesis—use only the Arrhenius theoretical definitions. Assume that the Arrhenius concept restricts dissociation and ionization: Bases dissociate to produce $OH^{-}_{(aq)}$ and a cation; and acids ionize to produce $H^{+}_{(aq)}$ and an anion.

You are expected to design an experiment to classify a number of common substances in solution (see **Materials**) as acidic, basic, or neutral. (Refer to Chapter 2 if you need help writing the formulas for each of the substances.) Complete a report, including the **Prediction**, **Experimental Design**, **Analysis**, and **Evaluation**.

Question

Which of the chemicals tested may be classified as an acid, a base, or neutral?

Prediction

(a) Based on Arrhenius' definitions, predict which of the chemicals in the Materials list will test as an acid, which as a base, and which as neutral.

Experimental Design

(b) Create an Experimental Design. Be sure to identify all variables, including any controls. Your experiment should involve qualitative analysis, incorporating one or more diagnostic tests. You should also note any necessary safety or disposal precautions.

(c) Write up your Procedure. Obtain your teacher's approval before conducting your experiment.

Materials

Hydrochloric acid and sodium hydroxide are irritants. Wear eye protection and a laboratory apron.

lab apron
eye protection
aqueous 0.10 mol/L solutions of:
- hydrogen chloride (a gas in solution)
- hydrogen acetate (vinegar)
- sodium hydroxide (lye, caustic soda)
- calcium hydroxide (slaked lime)
- ammonia (cleaning agent)
- sodium carbonate (washing soda, soda ash)
- sodium hydrogen carbonate (baking soda)
- sodium hydrogen sulfate (toilet bowl cleaner)
- calcium oxide (lime)
- carbon dioxide (carbonated beverage)
- aluminum nitrate (salt solution)
- sodium nitrate (fertilizer)

conductivity apparatus
blue litmus paper
red litmus paper
any other materials necessary for diagnostic tests

Procedure

1. Conduct your experiment.

Analysis

(d) Answer the Question: Which of the substances tested may be classified as an acid, a base, or neutral?

Evaluation

(e) Evaluate the validity of your Experimental Design, your Prediction, and the Arrhenius definition it was based on.

Revision of Arrhenius' Definitions

Arrhenius' definitions cannot always predict whether a substance is an acid or a base. Using Arrhenius' definitions, you would probably correctly predict that $HCl_{(aq)}$ and $HC_2H_3O_{2(aq)}$ are acids; that $NaOH_{(aq)}$ and $Ca(OH)_{2(aq)}$ are bases;

and that $NaNO_{3(aq)}$ is neutral. However, by Arrhenius's definitions, we would predict all of the following compounds to be neutral, but they are not:

- compounds of **hydrogen polyatomic ions** ($NaHCO_{3(aq)}$ and $NaHSO_{4(aq)}$)
- oxides of metals and nonmetals ($CaO_{(aq)}$ and $CO_{2(g)}$)
- compounds that are neither oxides nor hydroxides (e.g., $NH_{3(aq)}$ and $Na_2CO_{3(aq)}$), but yet are bases
- compounds that contain no hydrogen (e.g., $Al(NO_3)_{3(aq)}$), but yet are acids

Clearly, the Arrhenius theoretical definitions of acid and base need to be revised or replaced.

Figure 3
By passing infrared light through solutions of acids, Paul Giguère of the Université Laval, Quebec City, obtained clear evidence for the existence of hydronium ions in solution.

A theoretical concept has two major purposes: to explain current evidence and to predict the results of new experiments. While the first purpose is useful, it is not valued as much as the ability to predict results. Theoretical progress is made when theories not only explain what is known but also allow valid predictions to be made about new situations.

We need to revise Arrhenius' acid–base definitions to explain the exceptions listed above. The new theory involves two key ideas: collisions with water molecules and the nature of the hydrogen ion. Since all substances tested are in aqueous solution, then particles will constantly be colliding with, and may also react with, the water molecules present.

It is highly unlikely that the particle we call an aqueous hydrogen ion, $H^+_{(aq)}$, actually exists in an acidic solution. If such a particle were to come near polar water molecules, it would bond strongly to one or more of the molecules (**Figure 2**), that is, it would be hydrated. There is no evidence for the existence of unhydrated hydrogen ions in aqueous solution. However, the Canadian scientist Paul Giguère has done experiments that provide clear evidence for the existence of hydrated protons (**Figure 3**). The simplest representation of a hydrated proton is $H_3O^+_{(aq)}$, commonly called the **hydronium ion** (**Figure 4**).

We can now explain the formation of acidic solutions by strong acids such as $HCl_{(aq)}$ as a reaction with water, forming hydronium ions (**Figure 5**).

$$HCl_{(g)} + H_2O_{(l)} \xrightarrow{>99\%} H_3O^+_{(aq)} + Cl^-_{(aq)}$$

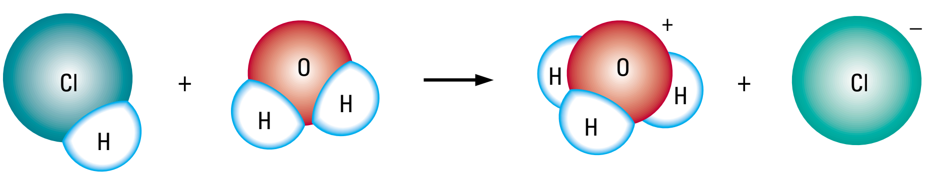

Figure 5
When gaseous hydrogen chloride dissolves in water, the HCl molecules are thought to collide and react with water molecules to form hydronium and chloride ions.

A strong acid, such as HCl, is considered to react completely with water. In other words, the collisions with water molecules are very successful, producing a 100% reaction. What about weak acids? Because they have lesser acidic properties,

hydrogen polyatomic ion: a bi-ion; a polyatomic ion with an available hydrogen (e.g., hydrogen carbonate (bicarbonate) ion, hydrogen sulfite (bisulfite) ion)

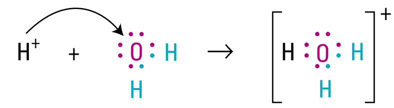

Figure 2
The Lewis (electron dot) diagram for a hydrogen ion has no electrons. A water molecule is believed to have two lone pairs of electrons, as shown in its Lewis diagram. The hydrogen ion (proton) is believed to bond to one of these lone pairs of electrons to produce the H_3O^+ ion.

hydronium ion: a hydrated hydrogen ion (proton), conventionally represented as $H_3O^+_{(aq)}$

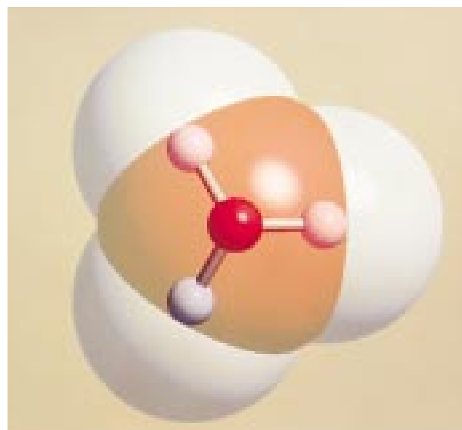

Figure 4
The hydronium ion is represented as a pyramidal structure. The oxygen atom is the apex and the three identical hydrogen atoms form the base of the pyramid.

DID YOU KNOW ?

pH

How does this new theory about acids affect our concept of pH? Fortunately, the mathematical definition of pH works equally well with hydronium ions as it did with hydrogen ions.

there must be fewer hydronium ions produced from the same volume and concentration of solution compared with strong acids. Therefore, the collisions of weak acid molecules with water cannot be very successful. Based on pH measurements, 0.10 mol/L acetic acid, a common weak acid, is only successful in forming hydronium ions in 1.3% of its collisions with water molecules.

$$HC_2H_3O_{2(aq)} + H_2O_{(l)} \xrightarrow{1.3\%} H_3O^+_{(aq)} + C_2H_3O^-_{2(aq)}$$

In general, *acidic solutions form when substances react with water to form hydronium ions.*

The concept of acids reacting with water to produce hydronium ions is a small adjustment in thinking when explaining or predicting the behaviour of typical acids. In most contexts, we can think of acids as either ionizing to produce hydrogen ions or reacting with water to produce hydronium ions.

Strong and Weak Bases

Evidence indicates that there are both strong bases (e.g., sodium hydroxide) and weak bases (e.g., ammonia). For equal concentrations of solutions, strong bases have high electrical conductivity and very high pH (>>7), whereas weak bases have low electrical conductivity and pH closer to 7. We can explain the behaviour of strong bases: They dissociate to increase the hydroxide ion concentration in an aqueous solution. Further evidence indicates that all ionic hydroxides are **strong bases**: 100% of the dissolved ionic hydroxides dissociates to release hydroxide ions.

strong base: (according to the Arrhenius theory and the "reaction-with-water" theory) an ionic hydroxide that dissociates 100% in water to produce hydroxide ions

weak base: (according to the "reaction-with-water" theory) a chemical that reacts less than 50% with water to produce hydroxide ions

What about **weak bases**? How can we explain their properties? The pure compounds (e.g., $NH_{3(g)}$) do not contain hydroxide ions, so they cannot dissociate to release hydroxide ions. Nevertheless, solutions of weak bases appear to contain hydroxide ions in a higher concentration than does pure water. Where do they come from? Clearly, this question cannot be answered by the Arrhenius definition of bases. We need to revise his theory to include a new concept: that weak base molecules or ions react with water to produce hydroxide ions. This remains consistent with the explanation for strong bases and for strong and weak acids. Weak bases do not react 100% with water. Evidence indicates that they commonly react less than 10%. This means that they produce fewer hydroxide ions than a similar amount of a strong base, which accounts for the weaker basic properties of weak bases.

Recall from Section 8.1 that ionic hydroxides produce basic solutions by simple dissociation. We know that ionic hydroxides, such as barium hydroxide, are strong bases.

$$Ba(OH)_{2(s)} \rightarrow Ba^{2+}_{(aq)} + 2\ OH^-_{(aq)} \qquad \text{(a strong base)}$$

Here, there is no need to consider a reaction with water because we know that ionic hydroxides, such as $Ba(OH)_2$, dissociate to produce hydroxide ions. However, there are many common examples of bases that are not ionic hydroxides, such as ammonia (window cleaner) and sodium carbonate (washing soda). Most bases, other than soluble ionic hydroxides, are weak bases. Weak bases may be either ionic or molecular compounds in their pure state.

Ammonia and sodium carbonate each form basic aqueous solutions as demonstrated by a litmus paper test. This equation for ammonia shows the theory to explain the evidence:

$$NH_{3(aq)} + H_2O_{(l)} \xrightarrow{<50\%} OH^-_{(aq)} + NH^+_{4(aq)} \qquad \text{(a weak base)}$$

The presence of hydroxide ions explains the basic solution, and the less than 100% reaction explains the weak base properties. (Note that both atoms and charge are conserved in the balanced equation.)

Sodium carbonate is an ionic compound with high solubility. According to the Arrhenius theory, sodium carbonate dissociates in water to produce aqueous ions of sodium and carbonate.

$$Na_2CO_{3(s)} \rightarrow 2\ Na^+_{(aq)} + CO^{2-}_{3(aq)}$$

The sodium ion cannot be responsible for the basic properties of the solution, because many sodium compounds (e.g., $NaCl_{(aq)}$) form neutral solutions. The basic character of carbonate solutions can be explained as resulting from their reaction with water.

$$CO^{2-}_{3(aq)} + H_2O_{(l)} \xrightarrow{<50\%} OH^-_{(aq)} + HCO^-_{3(aq)}$$

Note again that hydroxide ions explain the basic properties and that atoms and charge are conserved in the balanced equation.

We now have explanations for the production of hydroxide ions by bases: Strong bases (commonly ionic hydroxides) dissociate to produce hydroxide ions; and weak bases react with water to increase the hydroxide ion concentration. This theory is sometimes called the revised Arrhenius theory.

Sample Problem 1

A forensic technician tested the pH of a sodium cyanide solution and found that it had a pH greater than 7. Explain this evidence using chemical equations.

Solution

$$NaCN_{(s)} \rightarrow Na^+_{(aq)} + CN^-_{(aq)}$$

$$CN^-_{(aq)} + H_2O_{(l)} \xrightarrow{<50\%} OH^-_{(aq)} + HCN_{(aq)}$$

The cyanide ion produces hydroxide ions in reaction with water, so is a weak base. As a weak base, we expect the reaction to be less than 50% and the solution to have a pH greater than 7 but not, say, greater than 13.

SUMMARY Strong and Weak Acids and Bases

	Strong acids	Weak acids	Strong bases	Weak bases
empirical	very low pH (<<7)	medium to low pH (<7)	very high pH (>>7)	medium to high pH (>7)
	high conductivity	low conductivity	high conductivity	low conductivity
	fast reaction rate	slow reaction rate	fast reaction rate	slow reaction rate
solute	molecular	molecular and polyatomic ion	ionic hydroxide	molecular and polyatomic ion*
theoretical (Arrhenius)	completely ionized to form $H^+_{(aq)}$	partially ionized to form $H^+_{(aq)}$	completely dissociated into $OH^-_{(aq)}$	—
theoretical (revised Arrhenius)	completely reacted with water to form $H_3O^+_{(aq)}$	partially reacted with water to form $H_3O^+_{(aq)}$	completely dissociated to form $OH^-_{(aq)}$	partially reacted with water to form $OH^-_{(aq)}$

* Except the hydroxide ion, $OH^-_{(aq)}$

Figure 6
Johannes Brønsted created new theoretical definitions for acids and bases based upon proton transfer.

Practice

Understanding Concepts

12. What were some of the early ideas about the chemistry of acids? What evidence eventually showed that these ideas were false?
13. How well does the original Arrhenius theory predict and explain acids and bases?
14. What is the more recent replacement for the idea of a hydrogen ion causing acidic properties? State its name and formula.
15. Write chemical equations to explain the pH of a 0.1 mol/L solution of each substance.
    (a) $HCN_{(aq)}$; pH = 5
    (b) $HNO_{3(aq)}$; pH = 1
    (c) $Na_2SO_{4(aq)}$; pH = 8
    (d) $Sr(OH)_{2(aq)}$; pH = 13
16. In the previous question, identify the strong and weak acids and bases.

The Brønsted–Lowry Concept

Acid and base definitions, revised to include the ideas of the hydronium ion and reaction with water, are more effective in describing, explaining, and predicting the behaviour of acids and bases than are Arrhenius' original definitions. However, chemical research has shown that even these revised definitions are still too restrictive. Reactions of acids and bases do not always involve water. Also, evidence indicates that some entities that form basic solutions (such as $HCO^-_{3(aq)}$) can actually neutralize the solution of a stronger base. A broader concept is needed to describe, explain, and predict these properties of acids and bases.

New theories in science usually result from looking at the evidence in a way that has not occurred to other observers. A new approach to acids and bases was developed in 1923 by Johannes Brønsted (1879–1947) of Denmark (**Figure 6**) and independently by Thomas Lowry (1874–1936) of England. These scientists focused on the role of an acid and a base in a reaction rather than on the acidic or basic properties of their aqueous solutions. An acid, such as aqueous hydrogen chloride, functions in a way opposite to a base, such as aqueous ammonia. According to the Brønsted-Lowry concept, hydrogen chloride donates a proton (H^+) to a water molecule,

H^+

$$HCl_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$$

and ammonia accepts a proton from a water molecule.

H^+

$$NH_{3(aq)} + H_2O_{(l)} \rightarrow OH^-_{(aq)} + NH^+_{4(aq)}$$

Water does not have to be one of the reactants. For example, the hydronium ions present in a hydrochloric acid solution can react directly with dissolved ammonia molecules.

H^+

$$\underset{\text{acid}}{H_3O^+{}_{(aq)}} + \underset{\text{base}}{NH_{3(aq)}} \rightarrow H_2O_{(l)} + NH^+_{4(aq)}$$

We can describe this reaction as NH_3 molecules removing protons from H_3O^+ ions. Hydronium ions act as the **acid**, and ammonia molecules act as the **base**. Water is present as the solvent, but not as a primary reactant. In fact, water does not even have to be present, as evidenced by the reaction of hydrogen chloride and ammonia gases (**Figure 7**).

H^+

$$HCl_{(g)} + NH_{3(g)} \rightarrow NH_4Cl_{(s)}$$

acid base

Figure 7
One hazard of handling concentrated solutions of ammonia and hydrochloric acid is gas fumes. The photograph shows ammonia gas and hydrogen chloride gas escaping from their open bottles, and reacting to form a white cloud of very tiny crystals of $NH_4Cl_{(s)}$.

A substance can be classified as a Brønsted-Lowry acid or base only for a specific reaction. It is not a general property of a substance. This point is important—a substance may gain protons in one reaction, but lose them in another reaction with another substance. (For example, in the reaction of HCl with water shown above, water acts as the Brønsted-Lowry base; whereas, in the reaction of NH_3 with water, water acts as the Brønsted-Lowry acid.) A substance that appears to act as a Brønsted-Lowry acid in some reactions and as a Brønsted-Lowry base in other reactions is called **amphiprotic**. The hydrogen carbonate ion ($HCO^-_{3(aq)}$) in baking soda (**Figure 8**) is amphiprotic, like every other hydrogen polyatomic ion. Hydrogen polyatomic ions, as their name suggests, are polyatomic ions containing hydrogen. Examples of amphiprotic substances include $HCO^-_{3(aq)}$, $H_2O_{(l)}$, $HSO^-_{3(aq)}$, $H_2PO^-_{4(aq)}$, and $HPO^{2-}_{4(aq)}$.

Note that amphiprotic entities can either gain or lose a proton, as shown by the following reactions. First let's see what happens when the bicarbonate ion is added to the solution of a strong acid, which will contain hydronium ions.

acid: (according to the Brønsted-Lowry concept) a proton donor

base: (according to the Brønsted-Lowry concept) a proton acceptor

amphiprotic: a substance capable of acting as an acid or a base in different chemical reactions; an entity that can gain or lose a proton (sometimes called amphoteric)

H^+

$$HCO^-_{3(aq)} + H_3O^+_{(aq)} \rightarrow H_2CO_{3(aq)} + H_2O_{(l)} \qquad \text{(neutralizes a strong acid)}$$

base acid

Now let's look at the reaction of the bicarbonate ion with the solution of a strong base, which will contain hydroxide ions.

H^+

$$HCO^-_{3(aq)} + OH^-_{(aq)} \rightarrow CO^{2-}_{3(aq)} + H_2O_{(l)} \qquad \text{(neutralizes a strong base)}$$

acid base

Figure 8
Baking soda (sodium hydrogen carbonate, $NaHCO_3$) is a common household substance that is useful for many purposes other than baking. You can use it to neutralize both spilled acids and bases, and it can also be used as an extinguisher for small fires.

In both cases, the bicarbonate ion moves the pH of the solutions toward 7.

According to the Brønsted-Lowry concept, acid–base reactions involve the transfer of a proton. Therefore, the products formed in these reactions must differ from the reactants by a proton (H^+). If you look again at the equations above, you can see that this is true.

As another example, when acetic acid reacts with water, an acidic solution (containing hydronium ions) is formed.

H^+

$$HC_2H_3O_{2(aq)} + H_2O_{(l)} \rightarrow C_2H_3O^-_{2(aq)} + H_3O^+_{(aq)}$$

acid base

The acetate ion product is simply what is left after an acetic acid molecule loses its proton. The hydronium ion product is what is formed as a result of a water molecule gaining a proton. Any proton that is lost can, in principle, be regained and any proton that is gained can be lost in some other reaction. Therefore, we can consider the acetate ion to be a potential Brønsted-Lowry base: It could act as a base in another reaction. A product formed as a result of an acid losing a proton is called a **conjugate base.** Similarly, the hydronium ion is a potential Brønsted-Lowry acid: It could act as an acid in another reaction. A product resulting from a base gaining a proton is called a **conjugate acid.**

conjugate base: the base formed by removing a proton (H^+) from an acid

conjugate acid: the acid formed by adding a proton (H^+) to a base

conjugate acid–base pair

$$\underset{\text{acid}}{HC_2H_3O_{2(aq)}} + \underset{\text{base}}{H_2O_{(l)}} \rightarrow \underset{\text{conjugate base}}{C_2H_3O_2{}^-{}_{(aq)}} + \underset{\text{conjugate acid}}{H_3O^+_{(aq)}}$$

conjugate acid–base pair

A pair of substances that differ only by a proton is called a **conjugate acid–base pair** (Table 3).

conjugate acid–base pair: an acid–base pair that differs by one proton (H^+)

Table 3: Some Examples of Conjugate Acid–Base Pairs

Conjugate acid		Conjugate base
$H_2O_{(l)}$	and	$OH^-_{(aq)}$
$H_3O^+_{(aq)}$	and	$H_2O_{(l)}$
$HC_2H_3O_{2(aq)}$	and	$C_2H_3O^-_{2(aq)}$

SUMMARY Definitions of Acids and Bases

Arrhenius Definitions

- An acid ionizes in water to increase the hydrogen ion concentration.
- A base dissociates in water to increase the hydroxide ion concentration.
- A neutralization reaction involves the reaction of a hydrogen ion with a hydroxide ion.

Revised Arrhenius Definitions

- An acid reacts with water to increase the hydronium ion concentration.
- A base reacts with water to increase the hydroxide ion concentration.
- A neutralization reaction involves the reaction of a hydronium ion with a hydroxide ion.

Brønsted–Lowry Definitions

- An acid is a proton donor.
- A base is a proton acceptor.
- An acid–base neutralization reaction involves the transfer of one proton from the strongest acid present to the strongest base present.
- An amphiprotic substance is one that appears to act as a Brønsted-Lowry acid in some reactions and as a Brønsted-Lowry base in other reactions.
- A conjugate acid–base pair consists of two substances that differ only by one proton.

Practice

Understanding Concepts

17. According to the Brønsted-Lowry definitions, how are acids and bases different?
18. Classify each reactant in the following equations as a Brønsted-Lowry acid or base.
    (a) $HF_{(aq)} + SO_{3(aq)}^{2-} \rightarrow F_{(aq)}^{-} + HSO_{3(aq)}^{-}$
    (b) $CO_{3(aq)}^{2-} + HC_2H_3O_{2(aq)} \rightarrow C_2H_3O_{2(aq)}^{-} + HCO_{3(aq)}^{-}$
    (c) $H_3PO_{4(aq)} + OCl_{(aq)}^{-} \rightarrow H_2PO_{4(aq)}^{-} + HOCl_{(aq)}$
19. An aqueous hydrogen sulfate ion acts as the Brønsted-Lowry acid in the neutralization of a solution of hydrogen carbonate ions.
    (a) Write the chemical equation.
    (b) Identify two conjugate acid–base pairs.
20. What restrictions to acid–base reactions do the Brønsted-Lowry definitions remove?

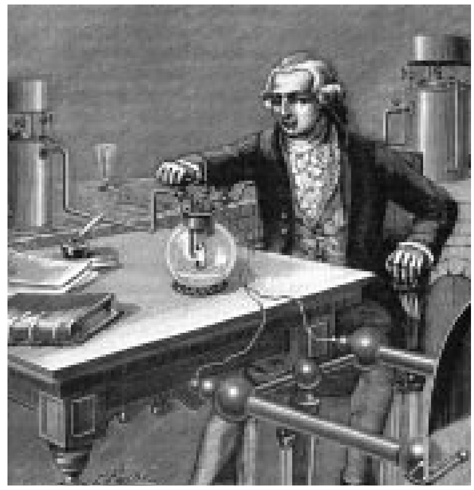

Figure 9
Antoine Lavoisier

Changing Ideas on Acids and Bases

Usually chemists discover the empirical properties of substances long before a theory is developed to describe, explain, and predict their behaviour. For example, several of the distinguishing properties of acids and bases were known by the middle of the 17th century. Additional properties, such as pH and the nature of acid–base reactions, were discovered by the early 20th century.

Let's take an overview of the developing theory of acids and bases. It is a story that took place over several hundred years, thanks to the innovative thoughts and painstaking laboratory investigations of many great scientists.

Antoine Lavoisier (1743–1794) (**Figure 9**) assumed that oxygen was responsible for acid properties and that acids were combinations of oxides and water. For example, sulfuric acid, H_2SO_4, was described as hydrated sulfur trioxide, $SO_3{\cdot}H_2O$. There were immediate problems with this theory because some oxide solutions, such as CaO, are basic, and several acids, such as HCl, are not formed from oxides. This evidence led to the rejection of the oxygen theory, although we still use the generalization that nonmetallic oxides (e.g., SO_3) form acidic solutions.

Figure 10
Sir Humphry Davy

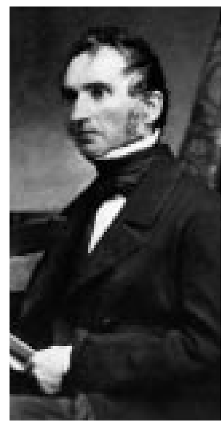

Figure 11
Justus von Liebig

Sir Humphry Davy (1778–1829) (**Figure 10**) advanced a theory that the presence of hydrogen gave a compound acidic properties. Justus von Liebig (1803–1873) (**Figure 11**) later expanded this theory to include the idea that acids are salts (compounds) of hydrogen. This meant that acids could be thought of as ionic compounds in which hydrogen had replaced the metal ion. However, this theory did not explain why many compounds containing hydrogen have neutral properties (e.g., CH_4) or basic properties (e.g., NH_3).

Svante Arrhenius (1859–1927) (**Figure 12**) developed a theory in 1887 that provided the first useful theoretical definition of acids and bases. He described acids as substances that ionize in aqueous solution to form hydrogen ions, and bases as substances that dissociate to form hydroxide ions in solution. This theory explained the process of neutralization by assuming that $H_{(aq)}^{+}$ and $OH_{(aq)}^{-}$ ions combine to form $H_2O_{(l)}$. The various strengths of acids were explained in terms of the degree (percentage) of ionization, but Arrhenius's theory is limited to aqueous solutions and cannot explain the properties of many common substances.

Figure 12
Svante Arrhenius

DID YOU KNOW ?

Prove Me Wrong!

In science, no theory can be proven. Well-established, accepted theories have a substantial quantity of supporting evidence. On the other hand, a theory can be disproven by a single, significant, reproducible observation. In Einstein's words: "No amount of experimentation can ever prove me right; a single experiment can prove me wrong."

Paul Giguère of the Université Laval, Quebec, found evidence that, in a solution, hydrogen ions are bonded to water molecules. The simplest representation of an aqueous hydrogen ion is $H_3O^+_{(aq)}$, commonly known as a hydronium ion. The concept of the hydronium ion was used to revise the Arrhenius definitions of acids and bases. Now acids could be described as substances that react with water to form hydronium ions, and bases could be described as substances that dissociate or react with water to form hydroxide ions. This revised Arrhenius theory is still limited to aqueous solutions but it does provide an explanation for the properties of aqueous solutions of nonmetal oxides, metal oxides, and polyatomic anions.

Johannes Brønsted (1879–1947) of Denmark and Thomas Lowry (1874–1936) of England independently developed a theory that focused on the role of acids and bases in a reaction rather than on the properties of their aqueous solutions. They defined acids as substances that donate protons (H^+) and bases as substances that accept protons in a chemical reaction. In the Brønsted-Lowry concept, a substance can only be defined as an acid or a base for a specific reaction. Ions such as hydrogen carbonate, $HCO^-_{3(aq)}$, or hydrogen sulfite, $HSO^-_{3(aq)}$, can act as acids in one reaction and as bases in another.

Changes in Knowledge

History indicates that it is unwise to assume that any scientific concept is the final word. Whenever scientists assume that they understand a subject, two things usually happen. Conceptual knowledge tends to remain static for a while, because little falsifying evidence exists or because any falsifying evidence is ignored. Then, when enough falsifying evidence accumulates, a revolution in thinking occurs within the scientific community in which the current concept is drastically revised or entirely replaced. The experience of the Swedish chemist Svante Arrhenius gives some insight into the difficulty scientists have in getting new ideas accepted.

While Arrhenius was attending the University of Uppsala near his home, he became intrigued by the question of why some aqueous solutions conduct electricity, but others do not. This problem had puzzled chemists ever since Sir Humphry Davy and Michael Faraday experimented over half a century earlier by passing electric currents through chemical substances.

Faraday believed that an electric current produces particles of electricity, which he called ions, in some solutions. He could not explain what ions were, or why they did not form in aqueous sugar or alcohol solutions.

As a university student, Arrhenius noticed that conducting solutions differed from non-conducting solutions in terms of another important property. The freezing point of any aqueous solution is lower than the freezing point of pure water; the more solute that is dissolved in the water, the more the freezing point is lowered (depressed). Arrhenius found that the freezing point depression of electrolytes in solution was always two or three times lower than that of non-electrolytes, in solutions of the same concentration. He concluded that when a solution such as pure table salt, NaCl, dissolves, it does not separate into NaCl molecules in solution but rather into two types of particles. Since the NaCl solution also conducts electricity, he reasoned that the particles must be electrically charged. In Arrhenius' view, the conductivity and freezing point evidence indicated that pure substances that form electrolytes are composed of charged ions, not neutral atoms. The stage was now set for a scientific controversy. Faraday was an established, respected scientist and his explanation agreed with Dalton's model of indivisible, neutral atoms. Arrhenius was an unknown university student and his theory contradicted Dalton's widely accepted model.

Despite strong supporting evidence, Arrhenius's creative idea was rejected by most of the scientific community, including his teachers. When Arrhenius presented his theory and its supporting evidence as part of his doctoral thesis, the examiners questioned him for four gruelling hours. They grudgingly passed him, but with the lowest possible mark.

For over a decade, only a few people supported Arrhenius' theory. Gradually, more supporting evidence accumulated, including J.J. Thomson's discovery of the electron in 1897. Soon, Arrhenius's theory of ions became widely accepted as the simplest and most logical explanation of the nature of electroytes. In 1903 Arrhenius won the Nobel Prize for the same thesis that had nearly failed him in his PhD examination years earlier.

Arrhenius' struggle to have his ideas accepted is not so unusual. Ideally, scientists are completely open-minded, but they are people, and many people resist change. We are sometimes reluctant to accept new ideas that conflict radically with familiar ones.

DID YOU KNOW ?

Scientific Concepts

"Creating a new theory is not like destroying an old barn and erecting a skyscraper in its place. It is rather like climbing a mountain, gaining new and wider views, discovering unexpected connections between our starting point and its rich environment. But the point from which we started out still exists and can be seen, although it appears smaller and forms a tiny part of our broad view gained by the mastery of the obstacles on our adventurous way up."

Albert Einstein (1879–1955)
German-born American theoretical physicist

Explore an Issue

Role Play: Evaluating New Ideas in Science

You are a member of the PhD examination committee for Svante Arrhenius. Is his experimental work reliable? Were his experiments well designed and carefully repeated? Is his interpretation of the experimental results valid? Is his reasoning logical and based on the evidence?

(a) Read the short summary of Arrhenius' work above, and research more detailed information in other references.

(b) Choose a role and prepare to question Arrhenius, played by your teacher, on his PhD thesis. You might consider the following roles:
- a senior professor at the university who firmly believes in Dalton's theory that atoms are indivisible, neutral particles;
- a scientist who frequently corresponded with Michael Faraday during his long, distinguished career;
- the professor who supervised Arrhenius' research and who frequently discussed the experimental results with him;
- a scientist who is dissatisfied with the current theories of electricity; or
- a young scientist who wants to know how Arrhenius' ideas would explain the acid–base properties of solutions.

DECISION-MAKING SKILLS

- ○ Define the Issue
- ● Identify Alternatives
- ● Research
- ● Analyze the Issue
- ● Defend a Decision
- ○ Evaluate

Sections 8.3–8.4 Questions

Understanding Concepts

1. Distinguish between a strong and weak acid using the concept of reaction with water.
2. What class of substances are strong bases? Explain their properties.
3. What are the properties of a weak base? Explain these properties.
4. Write appropriate chemical equations to explain the acidic or basic properties of each of the following substances added to water.
   (a) hydrogen bromide (acidic)
   (b) potassium hydroxide (basic)
   (c) benzoic acid, $HC_7H_5O_{2(aq)}$ (acidic)
   (d) sodium sulfide (basic)
5. Theories in science develop over a period of time. Illustrate this development by writing theoretical definitions of an acid, using the following concepts. Begin your answer with, "According to [name of concept], acids are substances that…"
   (a) the Arrhenius concept
   (b) the revised Arrhenius concept
   (c) the Brønsted-Lowry concept
6. Repeat question 5, defining bases. Refer to both strong and weak bases in your answer.
7. According to the Brønsted-Lowry concept, what happens in an acid–base reaction?
8. Use the Brønsted-Lowry definitions to identify each of the reactants in the following equations as acids or bases.
   (a) $HCO^-_{3(aq)} + S^{2-}_{(aq)} \rightarrow HS^-_{(aq)} + CO^{2-}_{3(aq)}$
   (b) $H_2CO_{3(aq)} + OH^-_{(aq)} \rightarrow HCO^-_{3(aq)} + H_2O_{(l)}$
9. Complete the following chemical equations to predict the acid–base reaction products.
   (a) $HSO^-_{4(aq)} + PO^{3-}_{4(aq)} \rightarrow$
   (b) $H_3O^+_{(aq)} + HPO^{2-}_{4(aq)} \rightarrow$
10. Some ions can form more than one conjugate acid–base pair. List the two conjugate acid–base pairs involving a hydrogen carbonate ion.
11. Identify the two acid–base conjugate pairs in each of the following reactions.
    (a) $H_3O^+_{(aq)} + HSO^-_{3(aq)} \rightarrow H_2O_{(l)} + H_2SO_{3(aq)}$
    (b) $OH^-_{(aq)} + HSO^-_{3(aq)} \rightarrow H_2O_{(l)} + SO^{2-}_{3(aq)}$

Applying Inquiry Skills

12. Baking soda is a common chemical but its chemical properties are difficult for chemists to explain and predict. Baking soda is amphiprotic and forms a basic solution. List some of the chemical properties of baking soda and indicate why some of these properties are difficult to explain and predict.
    Follow the links for Nelson Chemistry 11, 8.4.

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Making Connections

13. Common kitchen-variety baking soda has so many uses that it has entire books written about it. Use references to gather a list of uses for baking soda. Identify the uses that involve acid–base reactions.

8.5 Acid–Base Reactions

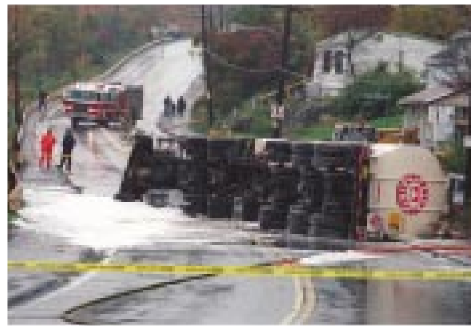

Figure 1
The first thing the emergency crew must do is identify the spilled chemical.

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:

$$HCl_{(aq)} + KOH_{(aq)} \rightarrow H_2O_{(l)} + KCl_{(aq)}$$

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

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

or $H_3O^+_{(aq)} + OH^-_{(aq)} \rightarrow 2\ H_2O_{(l)}$ (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:

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

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.

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?

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.

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

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**.

standard solution: a solution of precisely and accurately known concentration

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.

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.

$2\ HCl_{(aq)}$	+	$Na_2CO_{3(aq)}$	→	$H_2CO_{3(aq)}$	+	$2\ NaCl_{(aq)}$
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.

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

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

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

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

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

$$= 1.12\text{ mmol}$$

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

$$= 0.562\text{ mmol}$$

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

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

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.

Wear eye protection and a lab apron.

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

Materials

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

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.

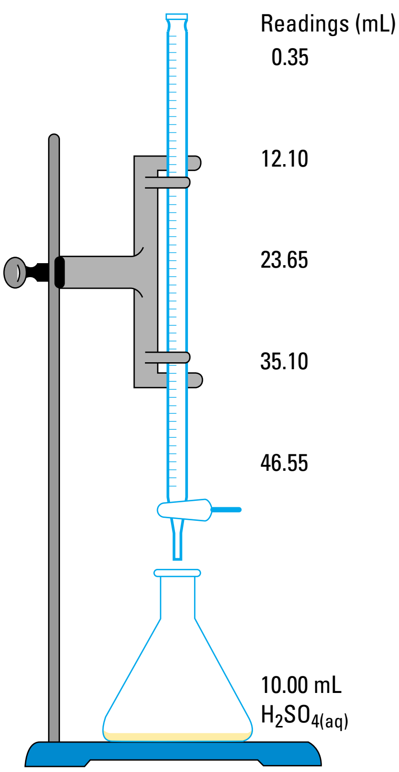

Figure 3
Sodium hydroxide titrant is added to samples of sulfuric acid in successive trials.

Section 8.5 Questions

Understanding Concepts

1. Write specific balanced chemical equations to illustrate each of the three types of characteristic reactions of acids.
2. An antacid tablet contains 0.912 g aluminum hydroxide to neutralize excess stomach acid. What volume of 0.10 mol/L stomach acid (assume $HCl_{(aq)}$) can one tablet neutralize?
3. Slaked lime, $Ca(OH)_{2(s)}$, can be used to neutralize the water in lakes that have been “killed” by acid rain. Ecologists hope that the original plants and animals will become reestablished. If the concentration of sulfuric acid in an acid lake is 1.2 x 10^{-3} mol/L, and 1.0 t of slaked lime is added to the lake, then what is the volume of the lake that could be neutralized?

Applying Inquiry Skills

4. Write two different Experimental Designs, involving different kinds of chemical reactions, to answer this Question: “What is the concentration of sodium hydroxide in an unknown solution?” List the Materials that you intend to use in your designs.
5. Laboratory technicians perform quantitative chemical analyses to determine the concentration of oxalic acid in, for example, foods and blood. Using the typical reactions of acids, and the technique of titration, write an Experimental Design to answer the Question: What is the concentration of oxalic acid, $H_2C_2O_{4(aq)}$, in a provided solution? Include a list of Materials and any necessary safety and disposal precautions. Then write out a detailed Procedure. With your teacher’s approval, carry out your Procedure, record your Evidence, and perform an Analysis. Your Evaluation should critique your Design, Materials, Procedure, skills, and therefore, your Evidence.

DID YOU KNOW ?

Oxalic Acid

Oxalic acid is a solid in its pure form and is relatively toxic to humans and animals. In nature, it is found in a number of green-leaved plants, including rhubarb (especially the leaves). Daily consumption of oxalic acid in an ordinary diet averages around 150 mg; the lethal dose is around 1500 mg.

Chapter 8 Summary

Key Expectations

Throughout this chapter, you have had the opportunity to do the following:

- Demonstrate an understanding of the Arrhenius and Brønsted-Lowry theories of acids and bases. (8.1, 8.4, 8.5)
- Explain qualitatively the difference between strong and weak acids and bases. (8.1, 8.2, 8.4)
- Demonstrate an understanding of the operational definition of pH. (8.2)
- Design and conduct an experiment to determine the effect of dilution on pH. (8.2)
- Identify and describe science- and technology-based careers that work with acidic and basic solutions. (8.3)
- Describe examples of solutions for which the concentration must be known and exact. (8.3, 8.5)
- Use the terms: ionization, dissociation, strong acid/base, weak acid/base, hydronium ion, proton transfer, conjugate acid/base, titration, titrant, and endpoint. (8.4, 8.5)
- Write balanced chemical equations for reactions involving acids and bases. (8.4, 8.5)
- Use titration procedure to determine the concentration of an acid or base in solution. (8.5)

Key Terms

acid (2)	neutralization (2)
amphiprotic	percentage ionization
base (2)	pH
conjugate acid	standard solution
conjugate acid–base pair	strong acid
conjugate base	strong base
dissociation	titrant
endpoint	titration
hydrogen polyatomic ion	weak acid
hydronium ion	weak base
ionization	

Make a Summary

In this chapter, you have studied acids and bases. To summarize your learning, sketch the equipment for a titration.

- Use, for example, 0.12 mol/L hydrochloric acid as the titrant to determine the concentration of sodium bicarbonate in "gripe water," an antacid for babies.
- Label the equipment and use key expectations and key terms to accompany your sketch.
- Try to display all of your learning from this chapter in and around this sketch.
- Use the chemicals, their concentrations, and their chemical formulas and chemical reactions as much as possible in your summary.

Reflect on your Learning

Revisit your answers to the Reflect on Your Learning questions at the begining of this chapter.

- How has your thinking changed?
- What new questions do you have?

Chapter 8 Review

Understanding Concepts

1. Provide an empirical definition for
   (a) an acid
   (b) a base
2. Many familiar household substances are acids or bases. Write IUPAC names for the following compounds and classify them as acids or bases.
   (a) lye, $NaOH_{(s)}$
   (b) vinegar, $HC_2H_3O_{2(aq)}$
   (c) milk of magnesia, $Mg(OH)_{2(s)}$
   (d) muriatic acid, $HCl_{(aq)}$
   (e) slaked lime, $Ca(OH)_{2(s)}$
   (f) window cleaner, $NH_{3(aq)}$
3. For each of the substances listed in the previous question, write a Prediction for whether the substance would form an acidic or a basic solution; and a chemical equation to explain your Prediction.
4. Solutions of hydrochloric acid and acetic acid are prepared, both with the same volume and concentration.
   (a) Compare the concentrations of hydrogen ions in the two solutions.
   (b) Explain the difference in hydrogen ion concentrations in the two solutions.
   (c) Compare the volumes of the sodium hydroxide solution required to neutralize each solution.
5. A student prepared 0.10 mol/L solutions of acetic acid, ammonia, hydrochloric acid, sodium chloride, and sodium hydroxide. Rank the solutions in order of increasing pH.
6. Calculate the pH of the following solutions.
   (a) lemon juice with $[H^+_{(aq)}] = 7.5 \times 10^{-3}$ mol/L
   (b) 2.5×10^{-3} mol/L nitric acid
7. Calculate the hydrogen ion concentration in each of the following solutions.
   (a) cleaning solution with a pH = 11.56
   (b) fruit juice with a pH = 3.50
8. List three empirical properties that may be used to rank acids in order of strength.
9. Aqueous solutions of nitric acid ($HNO_{3(aq)}$, a strong acid) and nitrous acid ($HNO_{2(aq)}$, a weak acid) of the same concentration are prepared.
   (a) How do their pH values compare?
   (b) Explain your answer, using chemical equations and the Brønsted-Lowry concept.
10. One sample of rainwater has a pH = 5, while another has a pH = 6. How do the hydrogen ion concentrations in the two samples compare?
11. Formal concepts of acids have existed since the 18th century. State the main idea and the limitations of each of the following: the oxygen concept; the hydrogen concept; Arrhenius's concept.
12. Use the Brønsted-Lowry concept to identify the reactants as acids or bases in the following reactions. In your answers, indicate the conjugate acid–base pairs.
   (a) $HSO^-_{4(aq)} + HCO^-_{3(aq)} \rightarrow SO^{2-}_{4(aq)} + H_2CO_{3(aq)}$
   (b) $HPO^{2-}_{4(aq)} + HSO^-_{4(aq)} \rightarrow H_2PO^-_{3(aq)} + SO^{2-}_{4(aq)}$
   (c) $H_2PO^-_{4(aq)} + H_2BO^-_{3(aq)} \rightarrow HPO^{2-}_{4(aq)} + H_3BO_{3(aq)}$
   (d) $HS^-_{(aq)} + HCO^-_{3(aq)} \rightarrow CO^{2-}_{3(aq)} + H_2S_{(aq)}$
   (e) $HSO^-_{3(aq)} + NH_{3(aq)} \rightarrow SO^{2-}_{3(aq)} + NH^+_{4(aq)}$
13. Some sources of drinking water contain low concentrations of nitrogen compounds. Because many nitrogen compounds are harmful to human health, especially that of infants, Health Canada has established maximum allowable levels for nitrogen compounds in drinking water. Write IUPAC names for the following nitrogen-containing entities and classify them as potential Brønsted-Lowry acids and/or bases.
   (a) $NH_{3(aq)}$
   (b) $NH^+_{4(aq)}$
   (c) $NO^-_{2(aq)}$
   (d) $NO^-_{3(aq)}$
14. The household cleaner TSP contains sodium phosphate as the active ingredient. In solution the phosphate ion reacts with water as shown in the following equation.

$$PO^{3-}_{4(aq)} + H_2O_{(aq)} \rightarrow HPO^{2-}_{4(aq)} + OH^-_{(aq)}$$

   (a) Identify two conjugate acid/base pairs in the above equation.
   (b) Predict whether a TSP solution would be acidic, basic, or neutral, and explain your prediction.
15. One component of acid rain is sulfurous acid. It forms when sulfur dioxide in the atmosphere dissolves in rainwater. The neutralization of sulfurous acid with a strong base proceeds in a two-step reaction. Complete the reaction equations.
   (a) $H_2SO_{3(aq)} + OH^-_{(aq)} \rightarrow$
   (b) $HSO^-_{3(aq)} + OH^-_{(aq)} \rightarrow$
16. Baking soda, $NaHCO_{3(s)}$, is a versatile substance as it reacts with both strong acids and strong bases.
   (a) Write a balanced chemical equation and a net ionic equation for the reaction of baking soda with hydrochloric acid.
   (b) Write a balanced chemical equation and a net ionic equation for the reaction of baking soda with sodium hydroxide.

(c) Classify each of the entities in the net ionic equations as Brønsted-Lowry acids or bases. Identify the two conjugate acid–base pairs in each reaction.
(d) What do the reactions in (a) and (b) indicate about the nature of the hydrogen carbonate ion? What term is used for this characteristic?

17. List the requirements of a chemical reaction used in a titration analysis to determine the concentration of a solute.
18. Why is it necessary to start a titration with at least one standard solution?
19. Define the following terms:
    (a) titration
    (b) titrant
    (c) endpoint
20. The chemical reactions that define acids are illustrated by the following reactions for sulfuric acid accidentally spilled from a car battery during an upset and a cleanup. Write chemical equations for each combination and indicate what it is about each reaction that helps define sulfuric acid as an acid.
    (a) Sulfuric acid reacts with zinc on the galvanized fenders of the car.
    (b) Sulfuric acid reacts with washing soda used to neutralize the acid.
    (c) Sulfuric acid reacts with slaked lime used to neutralize the acid.

Applying Inquiry Skills

21. What specific volumetric equipment is required to
    (a) contain the solution in the final steps of preparing a standard solution?
    (b) deliver precisely 7.8 mL of a solution in a dilution procedure?
    (c) deliver precisely 10.00 mL of a sample to be analyzed in a titration?
22. Design an experiment to test the generalization that diluting an acidic and a basic solution by a factor of 10 changes the pH by 1. Provide a list of Materials necessary to carry out your Experimental Design.
23. Complete the **Analysis** and **Evaluation** sections of the following investigation report.

Question

Which of the 0.1 mol/L solutions labelled 1, 2, 3, 4, and 5 is $KNO_{3(aq)}$, $NaHCO_{3(aq)}$, $HCl_{(aq)}$, $H_2SO_{3(aq)}$, and $NaOH_{(aq)}$?

Experimental Design

The solutions are all at the same concentration and temperature. A sample of each solution is tested to show its effect on litmus, and tested for conductivity and pH. Each solution is tested in an identical way.

Evidence

Table 1: Properties of the Unknown Solutions

Unknown	Litmus paper	Conductivity	pH
1	red to blue	high	10
2	no change	high	7
3	blue to red	low	2
4	blue to red	high	1
5	red to blue	high	13

Analysis

(a) Using the Evidence in **Table 1**, identify the unknown solutions.

Evaluation

(b) Critique the Experimental Design.

24. A chemistry student was given the task of determining the concentration of a hydrochloric acid solution so it can be used as a standard solution. Complete the **Evidence** and **Analysis** sections of the following investigation report.

Question

What is the molar concentration of the hydrochloric acid solution?

Experimental Design

Samples of a standard sodium carbonate solution are titrated with the unknown solution of hydrochloric acid. The colour change of a methyl orange indicator from yellow to orange is used as the endpoint. (Methyl orange is yellow in a basic solution and red in an acidic solution.)

Evidence

(a) Copy and complete **Table 2**.

Table 2: Titration of 10.00 mL of 0.120 mol/L $Na_2CO_{3(aq)}$ with $HCl_{(aq)}$

Trial	1	2	3	4
final buret reading (mL)	17.9	35.0	22.9	40.1
initial buret reading (mL)	0.3	17.9	5.9	22.9
volume of $HCl_{(aq)}$ added				
colour at endpoint	red	orange	orange	orange

Analysis

(b) Calculate the molar concentration of the hydrochloric acid solution.

25. Phosphoric acid is the active ingredient in many commercial rust removers. A technician in a product testing lab uses titration to determine the concentration of phosphoric acid in a bottle of commercial rust remover. Complete the **Evidence**, **Analysis**, and **Evaluation** sections for the investigation report.

Question

What is the molar concentration of the phosphoric acid in a rust-removing solution?

Prediction

According to the label on the bottle of commercial rust remover, the concentration of phosphoric acid in the product is 42.5% W/V, or 7.30 mol/L.

Experimental Design

The rust-removing solution is diluted by a factor of 10, and then 10.00-mL samples of the diluted product are titrated with a standardized 1.25 mol/L $NaOH_{(aq)}$ solution to the endpoint of the following reaction. Phenolphthalein is used as the indicator.

$$2\ NaOH_{(aq)} + H_3PO_{4(aq)} \rightarrow Na_2HPO_{4(aq)} + 2\ H_2O_{(l)}$$

Evidence

(a) Copy and complete **Table 3**.

Table 3: Titration of 10.00 mL of $H_3PO_{4(aq)}$ with 1.25 mol/L $NaOH_{(aq)}$

Trial	1	2	3	4
final buret reading (mL)	13.3	25.0	36.8	48.4
initial buret reading (mL)	0.4	13.3	25.0	36.8
volume of $NaOH_{(aq)}$ added				
colour at endpoint	deep red	pale pink	pale pink	pale pink

Analysis

(b) Calculate the molar concentration of the phosphoric acid in the solution.

Evaluation

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

Making Connections

26. Acids are frequently portrayed in movies as being extremely dangerous. For example, in *Dante's Peak* the aluminum boat is quickly eaten away by the acidic lake water. Find a clip from a movie that involves acids, and critique the portrayal of its chemistry.
27. Laboratory safety rules require students to wear eye protection when handling acids such as hydrochloric acid and sulfuric acid, yet dilute boric acid, $H_3BO_{3(aq)}$, is sold in drugstores as a soothing eye wash. Explain, including balanced chemical equations, why many acids are harmful to your eyes, but this solution is not.
28. Hardness of water is directly related to the quantity of calcium and magnesium ions present. At home, the scale that appears inside kettles, in pots, in coffee makers, and on showerheads after prolonged usage is evidence of hard water. Assuming the scale is entirely carbonates, design a process to remove it. Identify common Materials to be used, write appropriate chemical equations, and outline a Procedure, including safety and disposal. If you have this problem at home, obtain permission from a parent or guardian to remove the scale from some household items.
29. Measure the pH of tap water in your community, home, or school. If you live in a community with municipal drinking water treatment, use the Internet to find out how the pH level of your drinking water is regulated.

    Follow the links for Nelson Chemistry 11, Chapter 8 Review.

    GO TO www.science.nelson.com

30. All gardeners know that conifers like acidic soil, so why is acid rain so damaging? Find out why acid is good for trees in one case, but not in others.

    Follow the links for Nelson Chemistry 11, Chapter 8 Review.

    GO TO www.science.nelson.com

Exploring

31. The Brønsted-Lowry concept is not the last acid–base theory. Research the Lewis theory of acids and bases. What are its advantages and disadvantages, compared with the Brønsted-Lowry concept?

    Follow the links for Nelson Chemistry 11, Chapter 8 Review.

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Figure 1

Analysis of ASA

Acetylsalicylic acid, widely known as ASA or Aspirin (**Figure 1**), is the world's most commonly used pain-relieving drug—over ten thousand tonnes are manufactured every year in North America alone. ASA, $HC_9H_7O_{4(s)}$, is an organic (carbon chain) acid like acetic acid, $HC_2H_3O_{2(l)}$, and reacts with strong bases such as sodium hydroxide in the same way. You will be provided with a standard solution of $NaOH_{(aq)}$—one for which the concentration is known with a fair degree of certainty.

The main purpose of this task is to perform an investigation to accurately determine the ASA content of an over-the-counter pain reliever, and to compare the result with the quantity stated on the product label. There are several complicating factors that arise from the nature of ASA tablets:

- All tablet medications contain other "inert" ingredients besides the "medicinal" ones—and some of these other ingredients are not soluble.
- ASA is a large molecule with only one OH group to provide hydrogen bonding, so it is not very soluble in water. *The CRC Handbook* lists this compound as soluble in alcohols, ethers, and chloroform.
- Some tablets of this type have special coatings or are "buffered," meaning that other compounds are added that can and do interfere with ASA's reaction as an acid.

The use of a solvent other than water to dissolve your ASA tablet presents some more complications: There are three alcohols commonly available to the public. They are methanol, $CH_3OH_{(l)}$, ethanol, $C_2H_5OH_{(l)}$, and isopropanol, $CH_3CHOHCH_{3(l)}$. Methanol is sold as fondue fuel and gas line or windshield washer antifreeze, often called methyl hydrate. It is quite toxic if ingested, so it is not usually named as an alcohol for public sale. The idea is not to tempt or confuse consumers into thinking that it is drinkable. The other two alcohols are sold as "rubbing" alcohols in pharmacies. Note that while ethanol is not toxic in small quantities, ethanol used in anything other than liquors (alcoholic beverages) has other additives to make it toxic and undrinkable.

Investigation

Titration Analysis of an ASA Tablet

For this task you will use a standardized solution of sodium hydroxide to perform a titration to test the manufacturer's claim for the mass of ASA in a commercial tablet. The equation for the reaction of sodium hydroxide and ASA is

$$HC_9H_7O_{4(al)} + NaOH_{(aq)} \rightarrow NaC_9H_7O_{4(aq)} + H_2O_{(l)}$$

Question

What is the quantity of "active ingredient" in a commercial ASA tablet?

Prediction

(a) Write a prediction, based on the manufacturer's claim on the label of the bottle of ASA tablets.

Experimental Design

(b) Describe the design of your experiment.
(c) Explain why it would not be appropriate to use ether or chloroform as the solvent for your ASA tablet titration.
(d) Write a complete Procedure, including the precautions you will take to avoid safety hazards, and have it approved by your teacher.

Materials

ASA tablets
(e) Complete the Materials list.

Analysis

(f) Use stoichiometric calculations to determine the mass of ASA in a tablet.

Evaluation

(g) Evaluate the evidence gathered by evaluating the Experimental Design, the Materials, the Procedure, and the skills of experimenters. Indicate sources of error and suggest ways to improve the investigation.
(h) Calculate the percentage difference between the predicted and experimental mass.
(i) Evaluate the Prediction and the manufacturer's claim.

Synthesis

(j) Why are some ASA tablets coated or buffered?
(k) What other ingredients are included in medication tablets, and why?

Assessment

Your completed task will be assessed according to the following criteria:

Process

- Develop an appropriate Experimental Design.
- Choose and safely use appropriate Materials.
- Carry out the approved investigation.
- Record observations with appropriate precision.
- Analyze the results.
- Evaluate the Evidence.

Product

- Prepare a suitable lab report, including a discussion of the ingredients in this typical medication.
- Demonstrate an understanding of the relevant scientific concepts.
- Use terms, symbols, equations, and SI metric units correctly.

Unit 3 Review

Understanding Concepts

1. Using examples, explain why solutions are important to the study of chemistry.
2. A water testing lab uses samples of different types of water for reference. Indicate which of the following samples are solutions. Give reasons for your answers.
   (a) cloudy water
   (b) hard water
   (c) pure water
   (d) soft water
   (e) water with high iron content
3. What classes of compounds are
   (a) electrolytes?
   (b) nonelectrolytes?
4. The label on a bottle of a popular soft drink indicates that it contains the following solutes. Write the chemical formulas for these solutes, and classify them as ionic, molecular, or acid. (You may find the list of Common Chemicals in Appendix C useful.)
   (a) glucose
   (b) citric acid
   (c) sodium citrate
   (d) sodium benzoate
   (e) carbon dioxide
5. Modern solution theory is based on experimental work done by the Swedish chemist Svante Arrhenius.
   (a) What two properties of solutions did Arrhenius study to develop his theory of dissociation of electrolytes?
   (b) According to Arrhenius' theory, which ions are responsible for the acidic and the basic properties of a solution?
   (c) Use Arrhenius' theory to explain why acids and bases are always electrolytes, but compounds that form neutral solutions can be electrolytes or nonelectrolytes.
6. What happens when scientists find a theory to be inadequate? Use the Arrhenius theory of acids as an example.
7. The theoretical definition of a base changed from the Arrhenius concept, to the revised Arrhenius concept, to the Brønsted-Lowry concept. Describe these changes.
8. Most of the common household chemicals used in cooking, cleaning, and gardening are aqueous solutions. Copy and complete **Table 1**, classifying the active compound in each of the household products as acid, base, or neutral, and then as electrolyte or nonelectrolyte. Also provide the IUPAC name of each compound.

Table 1: Properties of Household Chemicals

Household chemical	Chemical formula	Acid, base, or neutral	Electrolyte or nonelectrolyte	IUPAC name
(a) syrup	$C_{12}H_{22}O_{11(aq)}$			
(b) kettle scale remover	$HCl_{(aq)}$			
(c) windshield washer fluid	$CH_3OH_{(aq)}$			
(d) oven cleaner	$NaOH_{(aq)}$			
(e) plant fertilizer	$KNO_{3(aq)}$			

9. A chemistry student conducts qualitative analysis tests on some common household products. Write net ionic equations for the predicted reactions.
   (a) Silver nitrate solution is added to a dilute solution of table salt, sodium chloride.
   (b) Aqueous sodium hydroxide is added to a solution of bluestone, copper(II) sulfate.
   (c) Calcium chloride solution is added to a solution of washing soda, sodium carbonate.
   (d) Aqueous lead(II) nitrate is added to a solution of potassium chloride lawn fertilizer.
   (e) Silver nitrate solution is added to a dilute solution of road salt, calcium chloride.
10. Give one example in which a high concentration of a solute is beneficial and one example in which it is harmful.
11. Ethane, $C_2H_{6(g)}$, from natural gas is used to synthesize hundreds of different compounds. Predict the solubility in water of each of the following ethane derivatives as high or low, based on molecular polarity.
   (a) $C_2H_3Cl_{(l)}$
   (b) $CH_3CHO_{(l)}$
   (c) $C_2H_5OH_{(l)}$
   (d) $C_2H_{4(g)}$
   (e) $HC_2H_3O_{2(l)}$
   (f) $C_2H_{2(g)}$
   (g) $C_2H_5NH_{3(l)}$
12. Predict from theory which compound in each pair is more soluble in water. Provide a theoretical explanation for your choice.
   (a) $(CH_3)_2CO_{(l)}$, nail polish remover, or $C_3H_7OH_{(l)}$, rubbing alcohol
   (b) $C_2H_4(OH)_{2(l)}$, radiator antifreeze, or $C_2H_4Cl_{2(l)}$, industrial solvent
   (c) two different industrial solvents, $CCl_{4(l)}$ or $CHCl_{3(l)}$

13. According to the label on a package of tea, a typical cup of green tea contains 20 mg of caffeine compared with 100 mg in a typical cup of black tea. Assuming a typical cup holds 225 mL, calculate the caffeine concentration for each beverage.
14. A 25.0-mL sample of saturated potassium chlorate solution at 22°C is evaporated to dryness, forming 2.16 g of crystals. What is the concentration of saturated potassium chlorate in g/100 mL at 22°C?
15. Water from Lake Ontario contains 162 ppm of dissolved minerals. If 2.5 L of this water is boiled to dryness in a kettle, what mass of minerals would remain?
16. Water is added to a 40.0-mL sample of 2.50 mol/L aqueous sodium hydroxide solution until the volume becomes 5.00 L. Calculate the concentration of the final solution.
17. One brand of bottled water contains 150 mg of calcium in a 2.00-L bottle. Calculate the concentration of calcium in
    (a) parts per million
    (b) moles per litre
18. A bottle of household vinegar is labelled 5% acetic acid (by volume). What minimum volume of vinegar contains 60 mL of acetic acid?
19. Copper(II) sulfate pentahydrate, $CuSO_4 \cdot 5H_2O_{(s)}$, is a fungicide used to treat fenceposts (before they are put in the ground) to delay the rotting process. The maximum concentration of copper(II) sulfate pentahydrate at 0°C is 14.3 g/100 mL of solution.
    (a) Calculate the molar solubility of copper(II) sulfate pentahydrate at 0°C.
    (b) Find the maximum mass of copper(II) sulfate pentahydrate that can be dissolved to make 4.54 L of solution.
20. Precipitation reactions are often used in "chemistry magic shows" to produce sudden colour changes. Write net ionic equations to represent the following reactions.
    (a) Solutions of silver nitrate and potassium iodide are mixed to produce a pale yellow precipitate.
    (b) Solutions of copper(II) nitrate and potassium hydroxide produce a blue precipitate when mixed.
    (c) Lead(II) nitrate and sodium sulfide solutions are mixed to produce a black precipitate.
    (d) Solutions of sodium carbonate and calcium chloride are mixed to produce a white precipitate.
    (e) When solutions of lead(II) nitrate and sodium iodide are mixed, a bright yellow precipitate is formed.
21. Dilution of aqueous solutions is an essential laboratory skill. Use the values for concentrated reagents in Appendix C to answer the following questions.
    (a) A student needs to prepare 500 mL of 1.00 mol/L $HCl_{(aq)}$. Calculate the volume of concentrated reagent required.
    (b) A lab technician finds that she has only 250 mL of concentrated phosphoric acid in the storage cabinet. What is the maximum volume of 2.00 mol/L $H_3PO_{4(aq)}$ that she will be able to prepare?
    (c) A student dilutes 25 mL of saturated sulfurous acid to a final volume of 750 mL. What is the concentration of sulfurous acid in the diluted solution?
    (d) A lab technician needs to prepare 4.54 L of 2.50 mol/L ammonia solution. What volume of concentrated ammonia will he require?
22. On what factors does the hydrogen ion concentration depend in any solution of a weak acid?
23. Copy and complete **Table 2**.

Table 2: pH and Hydrogen Ion Concentration of Some Common Beverages

Beverage	pH	$[H^+_{(aq)}]$ (mol/L)
antacid	10.46	
apple juice	3.14	
beer		3.12×10^{-5}
cider		7.18×10^{-4}
tap water	7.86	

24. The following chemicals are sometimes used in water treatment plants. Write dissociation equations for the solutes in each of the following solutions.
    (a) $Al_2(SO_4)_{3(aq)}$
    (b) $Ca(OH)_{2(aq)}$
    (c) $NaOCl_{(aq)}$
25. The pesticide malathion has a solubility of 145 ppm at room temperature. Calculate the volume of room-temperature water needed to completely dissolve 1.00 g of malathion.
26. Canadians use an average of 200 L of water per person per day. The sodium carbonate concentration in a sample of well water is tested and found to be 225 ppm. Determine the mass of sodium carbonate in 200 L of well water.
27. Arsenic commonly occurs in the ores of gold, lead, copper, and nickel. The unintended release of compounds containing arsenic into ground water is a concern in areas where these metals are mined and smelted.

Because of its toxicity, Health Canada has set the maximum acceptable concentration (MAC) of arsenic in drinking water at 25 ppb.

(a) If a community's water supply contained the MAC of arsenic, what mass of arsenic would a person consume if he/she drank an average of 1.5 L of this water per day for one year (365 days)?

(b) A 25.0-mL sample of water taken from a stream near an abandoned gold mine is found to contain 1.2 μg of arsenic. Does this exceed the Health Canada MAC for arsenic?

(c) Arsenic in the compound sodium arsenate, $Na_3AsO_{4(aq)}$, can be removed from drinking water by treatment with iron(III) chloride solution, to precipitate the toxic anion. Write a net ionic equation for the reaction of iron(III) chloride solution with the aqueous sodium arsenate.

28. A 250-mL sample of polluted, acidic pond water is titrated with 0.0085 mol/L $NaOH_{(aq)}$. If 9.3 mL of the base was required to reach the endpoint, what was the molar concentration of acid (assume $HCl_{(aq)}$) in the pond water? (Assume the acid in the pond water is dilute $HCl_{(aq)}$.)

29. When 0.10 mol/L solutions of potassium benzoate and sodium hydrogen sulfate are mixed, a precipitate of white crystalline flakes is formed. The net ionic equation for the reaction is

$$C_7H_5O_{2(aq)}^- + HSO_{4(aq)}^- \rightarrow HC_7H_5O_{2(s)} + SO_{4(aq)}^{2-}$$

(a) Identify each reactant as a Brønsted-Lowry acid or base.

(b) Provide a theoretical explanation for the observation that the $C_7H_5O_2^-$ ion is more soluble in water than the $HC_7H_5O_2$ molecule.

30. Does a Brønsted-Lowry acid in solution always form an acidic solution? Provide an example to explain your answer.

Applying Inquiry Skills

31. Design an experiment to identify five unlabelled solutions of equal concentration that are known to contain the following compounds: road salt, $CaCl_{2(aq)}$; table sugar, $C_{12}H_{22}O_{11(aq)}$; vinegar, $HC_2H_3O_{2(aq)}$; oven cleaner, $NaOH_{(aq)}$; and battery acid, $H_2SO_{4(aq)}$. Your materials can include litmus paper and a conductivity meter.

32. A solution is believed to contain chloride ions and sulfide ions. Assuming that the solution does not contain any interfering ions, design an experiment to test for the presence of these two ions.

33. An unknown solution conducts electricity, turns red litmus paper blue, and forms a precipitate when sodium sulfate solution is added. What is one possible chemical formula for the solute present in the original solution?

34. A lab technician wants to prepare 250 mL of 0.05 mol/L nitric acid by diluting a stock solution.

(a) Calculate the volume of 15.4 mol/L nitric acid required to prepare this solution.

(b) Write a complete Procedure for preparing this solution.

35. Complete the Analysis of the following investigation report, and write an Evaluation of the Experimental Design.

Question

Which of the 0.1 mol/L solutions labelled 1, 2, 3, 4, 5, 6, and 7 is $KCl_{(aq)}$, $CuBr_{2(aq)}$, $HCl_{(aq)}$, $HC_2H_3O_{2(aq)}$, $NH_{3(aq)}$, $CH_3OH_{(aq)}$, and $NaOH_{(aq)}$?

Experimental Design

The solutions are prepared so that they all have the same concentration and temperature. A sample of each solution is observed to determine its colour, and is then tested for pH, conductivity, and effect on blue and red litmus papers.

Evidence

Table 3: Properties of the Unknown Solutions

Unknown	Colour	pH	Conductivity	Litmus paper
1	colourless	13	high	red to blue
2	blue	7	high	no change
3	colourless	7	low	no change
4	colourless	1	high	blue to red
5	colourless	7	high	no change
6	colourless	11	high	red to blue
7	colourless	3	high	blue to red

Analysis

(a) Using the Evidence in **Table 3**, identify each of the unknown solutions.

Evaluation

(b) Was the Experimental Design effective?

Making Connections

36. Several solutions, such as CLR, are sold in retail stores as household rust and lime (hard-water scale) removers. The labels all indicate that the active ingredient is glycolic acid, but do not provide the formula for this com-

pound. A student wonders if another acid could be substituted for the commercial solution, and checks out a library reference. The student finds that the percentage ionization of a 0.10 mol/L solution of glycolic acid is 3.9%. In Chapter 8, the percentage ionization of a 0.10 mol/L solution of acetic acid (the acid in vinegar) is given as 1.3%.

(a) Use any library or Internet resource to find the correct chemical formula for glycolic acid (also known by its IUPAC name: hydroxyacetic acid).
(b) Is glycolic acid a strong acid or a weak acid?
(c) Would a solution of glycolic acid react more or less rapidly with rust than a vinegar solution of the same concentration? Explain your reasoning.
(d) Would a solution of glycolic acid react with more or less iron(III) hydroxide (rust) than a vinegar solution of the same concentration, or with the same amount? Explain your reasoning.
(e) Identify and briefly discuss several factors that might affect the student's decision on whether to substitute vinegar for a commercial glycolic acid solution for rust and hard-water scale removal tasks around the home.

37. A shopper finds a 700-mL spray bottle of "bleach cleaner" with a label claiming that the product has a "fresh" scent, contains a powerful cleaner, and that it kills 99.9% of germs. The fine print states that it contains 1.84 % W/V sodium hypochlorite, $NaOCl_{(aq)}$, bleach when packed. The "powerful" cleaner is not identified, so it is probably a non-controlled product like liquid detergent. Many cleaning solutions have perfumes added, so that likely explains the "fresh" scent. A 3.6-L jug of regular household bleach with another brand name, found on the same shelf, is labelled 5.25% W/V sodium hypochlorite, when packed.
(a) When bleach bottles are opened, an easily identified odour of chlorine tells you that some chlorine has escaped. Explain what this observation tells you about the tendency of aqueous hypochlorite ions to react, and why the bottle labels must use the phrase, "when packed," when listing the minimum concentration value of sodium hypochlorite.
(b) What minimum mass of $NaOCl_{(s)}$ was required to make the bleach solution contained in the large 3.6-L jug?
(c) If your main purpose is to kill microorganisms, you could just dilute some regular bleach and wipe your cutting board, countertop or stove surface. How many times more NaOCl solute, by mass, is in the large jug than is in the spray bottle?
(d) Find some bleach solution containers at home or in the supermarket. Try to find large and small ones with the same brand name, such as Clorox or Javex. Record the volumes, concentration, and prices; compare the two by calculating the cost per gram of NaOCl in each container.
(e) The shopper notes that both bottles are labelled as corrosive, and marked with the Hazardous Household Product Symbol (HHPS) in a triangular frame. What is the symbol used, and what does the triangular frame mean?
(f) The labels also say "Dangerous Gas Formed When Mixed with Acid." Write the net ionic equation for hydrogen ions reacting with hypochlorite ions, $OCl^-_{(aq)}$, and chloride ions, to form chlorine gas and water. Explain why this reaction as written cannot be considered a simple Brønsted-Lowry acid–base reaction.

Exploring

38. On a large jug of household bleach, the words CAUTION and ATTENTION, and a HHPS corrosive logo are in large print, in black, on a white background. On a smaller spray bottle of a more dilute solution, the same words and logo are used, but in small print, in light green, on a slightly darker green background. Both containers also state "Dangerous Gas Formed When Mixed with Acid"—which is a *very* important caution—so that people do not try to clean their sinks and toilets by mixing acids (such as rust-removing cleaners) with bleach.
(a) Use the Internet to determine what the legal limit is for exposure to chlorine gas in air, and what effect breathing chlorine has on humans.
(b) Use the Internet to determine what the Canadian legal standards are (if any) for hazard warnings on product labels. Suggest some changes in the standards that would improve safety.

Follow the links for Nelson Chemisty 11, Unit 3 Review.

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Unit 3 Review

Understanding Concepts

1. Using examples, explain why solutions are important to the study of chemistry.
2. A water testing lab uses samples of different types of water for reference. Indicate which of the following samples are solutions. Give reasons for your answers.
   (a) cloudy water
   (b) hard water
   (c) pure water
   (d) soft water
   (e) water with high iron content
3. What classes of compounds are
   (a) electrolytes?
   (b) nonelectrolytes?
4. The label on a bottle of a popular soft drink indicates that it contains the following solutes. Write the chemical formulas for these solutes, and classify them as ionic, molecular, or acid. (You may find the list of Common Chemicals in Appendix C useful.)
   (a) glucose
   (b) citric acid
   (c) sodium citrate
   (d) sodium benzoate
   (e) carbon dioxide
5. Modern solution theory is based on experimental work done by the Swedish chemist Svante Arrhenius.
   (a) What two properties of solutions did Arrhenius study to develop his theory of dissociation of electrolytes?
   (b) According to Arrhenius' theory, which ions are responsible for the acidic and the basic properties of a solution?
   (c) Use Arrhenius' theory to explain why acids and bases are always electrolytes, but compounds that form neutral solutions can be electrolytes or nonelectrolytes.
6. What happens when scientists find a theory to be inadequate? Use the Arrhenius theory of acids as an example.
7. The theoretical definition of a base changed from the Arrhenius concept, to the revised Arrhenius concept, to the Brønsted-Lowry concept. Describe these changes.
8. Most of the common household chemicals used in cooking, cleaning, and gardening are aqueous solutions. Copy and complete **Table 1**, classifying the active compound in each of the household products as acid, base, or neutral, and then as electrolyte or nonelectrolyte. Also provide the IUPAC name of each compound.

Table 1: Properties of Household Chemicals

Household chemical	Chemical formula	Acid, base, or neutral	Electrolyte or nonelectrolyte	IUPAC name
(a) syrup	$C_{12}H_{22}O_{11(aq)}$			
(b) kettle scale remover	$HCl_{(aq)}$			
(c) windshield washer fluid	$CH_3OH_{(aq)}$			
(d) oven cleaner	$NaOH_{(aq)}$			
(e) plant fertilizer	$KNO_{3(aq)}$			

9. A chemistry student conducts qualitative analysis tests on some common household products. Write net ionic equations for the predicted reactions.
   (a) Silver nitrate solution is added to a dilute solution of table salt, sodium chloride.
   (b) Aqueous sodium hydroxide is added to a solution of bluestone, copper(II) sulfate.
   (c) Calcium chloride solution is added to a solution of washing soda, sodium carbonate.
   (d) Aqueous lead(II) nitrate is added to a solution of potassium chloride lawn fertilizer.
   (e) Silver nitrate solution is added to a dilute solution of road salt, calcium chloride.
10. Give one example in which a high concentration of a solute is beneficial and one example in which it is harmful.
11. Ethane, $C_2H_{6(g)}$, from natural gas is used to synthesize hundreds of different compounds. Predict the solubility in water of each of the following ethane derivatives as high or low, based on molecular polarity.
   (a) $C_2H_3Cl_{(l)}$
   (b) $CH_3CHO_{(l)}$
   (c) $C_2H_5OH_{(l)}$
   (d) $C_2H_{4(g)}$
   (e) $HC_2H_3O_{2(l)}$
   (f) $C_2H_{2(g)}$
   (g) $C_2H_5NH_{3(l)}$
12. Predict from theory which compound in each pair is more soluble in water. Provide a theoretical explanation for your choice.
   (a) $(CH_3)_2CO_{(l)}$, nail polish remover, or $C_3H_7OH_{(l)}$, rubbing alcohol
   (b) $C_2H_4(OH)_{2(l)}$, radiator antifreeze, or $C_2H_4Cl_{2(l)}$, industrial solvent
   (c) two different industrial solvents, $CCl_{4(l)}$ or $CHCl_{3(l)}$

13. According to the label on a package of tea, a typical cup of green tea contains 20 mg of caffeine compared with 100 mg in a typical cup of black tea. Assuming a typical cup holds 225 mL, calculate the caffeine concentration for each beverage.
14. A 25.0-mL sample of saturated potassium chlorate solution at 22°C is evaporated to dryness, forming 2.16 g of crystals. What is the concentration of saturated potassium chlorate in g/100 mL at 22°C?
15. Water from Lake Ontario contains 162 ppm of dissolved minerals. If 2.5 L of this water is boiled to dryness in a kettle, what mass of minerals would remain?
16. Water is added to a 40.0-mL sample of 2.50 mol/L aqueous sodium hydroxide solution until the volume becomes 5.00 L. Calculate the concentration of the final solution.
17. One brand of bottled water contains 150 mg of calcium in a 2.00-L bottle. Calculate the concentration of calcium in
    (a) parts per million
    (b) moles per litre
18. A bottle of household vinegar is labelled 5% acetic acid (by volume). What minimum volume of vinegar contains 60 mL of acetic acid?
19. Copper(II) sulfate pentahydrate, $CuSO_4 \cdot 5H_2O_{(s)}$, is a fungicide used to treat fenceposts (before they are put in the ground) to delay the rotting process. The maximum concentration of copper(II) sulfate pentahydrate at 0°C is 14.3 g/100 mL of solution.
    (a) Calculate the molar solubility of copper(II) sulfate pentahydrate at 0°C.
    (b) Find the maximum mass of copper(II) sulfate pentahydrate that can be dissolved to make 4.54 L of solution.
20. Precipitation reactions are often used in "chemistry magic shows" to produce sudden colour changes. Write net ionic equations to represent the following reactions.
    (a) Solutions of silver nitrate and potassium iodide are mixed to produce a pale yellow precipitate.
    (b) Solutions of copper(II) nitrate and potassium hydroxide produce a blue precipitate when mixed.
    (c) Lead(II) nitrate and sodium sulfide solutions are mixed to produce a black precipitate.
    (d) Solutions of sodium carbonate and calcium chloride are mixed to produce a white precipitate.
    (e) When solutions of lead(II) nitrate and sodium iodide are mixed, a bright yellow precipitate is formed.
21. Dilution of aqueous solutions is an essential laboratory skill. Use the values for concentrated reagents in Appendix C to answer the following questions.
    (a) A student needs to prepare 500 mL of 1.00 mol/L $HCl_{(aq)}$. Calculate the volume of concentrated reagent required.
    (b) A lab technician finds that she has only 250 mL of concentrated phosphoric acid in the storage cabinet. What is the maximum volume of 2.00 mol/L $H_3PO_{4(aq)}$ that she will be able to prepare?
    (c) A student dilutes 25 mL of saturated sulfurous acid to a final volume of 750 mL. What is the concentration of sulfurous acid in the diluted solution?
    (d) A lab technician needs to prepare 4.54 L of 2.50 mol/L ammonia solution. What volume of concentrated ammonia will he require?
22. On what factors does the hydrogen ion concentration depend in any solution of a weak acid?
23. Copy and complete **Table 2**.

Table 2: pH and Hydrogen Ion Concentration of Some Common Beverages

Beverage	pH	$[H^+_{(aq)}]$ (mol/L)
antacid	10.46	
apple juice	3.14	
beer		3.12×10^{-5}
cider		7.18×10^{-4}
tap water	7.86	

24. The following chemicals are sometimes used in water treatment plants. Write dissociation equations for the solutes in each of the following solutions.
    (a) $Al_2(SO_4)_{3(aq)}$
    (b) $Ca(OH)_{2(aq)}$
    (c) $NaOCl_{(aq)}$
25. The pesticide malathion has a solubility of 145 ppm at room temperature. Calculate the volume of room-temperature water needed to completely dissolve 1.00 g of malathion.
26. Canadians use an average of 200 L of water per person per day. The sodium carbonate concentration in a sample of well water is tested and found to be 225 ppm. Determine the mass of sodium carbonate in 200 L of well water.
27. Arsenic commonly occurs in the ores of gold, lead, copper, and nickel. The unintended release of compounds containing arsenic into ground water is a concern in areas where these metals are mined and smelted.

Because of its toxicity, Health Canada has set the maximum acceptable concentration (MAC) of arsenic in drinking water at 25 ppb.

(a) If a community's water supply contained the MAC of arsenic, what mass of arsenic would a person consume if he/she drank an average of 1.5 L of this water per day for one year (365 days)?

(b) A 25.0-mL sample of water taken from a stream near an abandoned gold mine is found to contain 1.2 μg of arsenic. Does this exceed the Health Canada MAC for arsenic?

(c) Arsenic in the compound sodium arsenate, $Na_3AsO_{4(aq)}$, can be removed from drinking water by treatment with iron(III) chloride solution, to precipitate the toxic anion. Write a net ionic equation for the reaction of iron(III) chloride solution with the aqueous sodium arsenate.

28. A 250-mL sample of polluted, acidic pond water is titrated with 0.0085 mol/L $NaOH_{(aq)}$. If 9.3 mL of the base was required to reach the endpoint, what was the molar concentration of acid (assume $HCl_{(aq)}$) in the pond water? (Assume the acid in the pond water is dilute $HCl_{(aq)}$.)

29. When 0.10 mol/L solutions of potassium benzoate and sodium hydrogen sulfate are mixed, a precipitate of white crystalline flakes is formed. The net ionic equation for the reaction is

$$C_7H_5O^-_{2(aq)} + HSO^-_{4(aq)} \rightarrow HC_7H_5O_{2(s)} + SO^{2-}_{4(aq)}$$

(a) Identify each reactant as a Brønsted-Lowry acid or base.

(b) Provide a theoretical explanation for the observation that the $C_7H_5O_2^-$ ion is more soluble in water than the $HC_7H_5O_2$ molecule.

30. Does a Brønsted-Lowry acid in solution always form an acidic solution? Provide an example to explain your answer.

Applying Inquiry Skills

31. Design an experiment to identify five unlabelled solutions of equal concentration that are known to contain the following compounds: road salt, $CaCl_{2(aq)}$; table sugar, $C_{12}H_{22}O_{11(aq)}$; vinegar, $HC_2H_3O_{2(aq)}$; oven cleaner, $NaOH_{(aq)}$; and battery acid, $H_2SO_{4(aq)}$. Your materials can include litmus paper and a conductivity meter.

32. A solution is believed to contain chloride ions and sulfide ions. Assuming that the solution does not contain any interfering ions, design an experiment to test for the presence of these two ions.

33. An unknown solution conducts electricity, turns red litmus paper blue, and forms a precipitate when sodium sulfate solution is added. What is one possible chemical formula for the solute present in the original solution?

34. A lab technician wants to prepare 250 mL of 0.50 mol/L nitric acid by diluting a stock solution.

(a) Calculate the volume of 15.4 mol/L nitric acid required to prepare this solution.

(b) Write a complete Procedure for preparing this solution.

35. Complete the Analysis of the following investigation report, and write an Evaluation of the Experimental Design.

Question

Which of the 0.1 mol/L solutions labelled 1, 2, 3, 4, 5, 6, and 7 is $KCl_{(aq)}$, $CuBr_{2(aq)}$, $HCl_{(aq)}$, $HC_2H_3O_{2(aq)}$, $NH_{3(aq)}$, $CH_3OH_{(aq)}$, and $NaOH_{(aq)}$?

Experimental Design

The solutions are prepared so that they all have the same concentration and temperature. A sample of each solution is observed to determine its colour, and is then tested for pH, conductivity, and effect on blue and red litmus papers.

Evidence

Table 3: Properties of the Unknown Solutions

Unknown	Colour	pH	Conductivity	Litmus paper
1	colourless	13	high	red to blue
2	blue	7	high	no change
3	colourless	7	low	no change
4	colourless	1	high	blue to red
5	colourless	7	high	no change
6	colourless	11	high	red to blue
7	colourless	3	high	blue to red

Analysis

(a) Using the Evidence in **Table 3**, identify each of the unknown solutions.

Evaluation

(b) Was the Experimental Design effective?

Making Connections

36. Several solutions, such as CLR, are sold in retail stores as household rust and lime (hard-water scale) removers. The labels all indicate that the active ingredient is glycolic acid, but do not provide the formula for this com-

pound. A student wonders if another acid could be substituted for the commercial solution, and checks out a library reference. The student finds that the percentage ionization of a 0.10 mol/L solution of glycolic acid is 3.9%. In Chapter 8, the percentage ionization of a 0.10 mol/L solution of acetic acid (the acid in vinegar) is given as 1.3%.

(a) Use any library or Internet resource to find the correct chemical formula for glycolic acid (also known by its IUPAC name: hydroxyacetic acid).
(b) Is glycolic acid a strong acid or a weak acid?
(c) Would a solution of glycolic acid react more or less rapidly with rust than a vinegar solution of the same concentration? Explain your reasoning.
(d) Would a solution of glycolic acid react with more or less iron(III) hydroxide (rust) than a vinegar solution of the same concentration, or with the same amount? Explain your reasoning.
(e) Identify and briefly discuss several factors that might affect the student's decision on whether to substitute vinegar for a commercial glycolic acid solution for rust and hard-water scale removal tasks around the home.

37. A shopper finds a 700-mL spray bottle of "bleach cleaner" with a label claiming that the product has a "fresh" scent, contains a powerful cleaner, and that it kills 99.9% of germs. The fine print states that it contains 1.84 % W/V sodium hypochlorite, $NaOCl_{(aq)}$, bleach when packed. The "powerful" cleaner is not identified, so it is probably a non-controlled product like liquid detergent. Many cleaning solutions have perfumes added, so that likely explains the "fresh" scent. A 3.6-L jug of regular household bleach with another brand name, found on the same shelf, is labelled 5.25% W/V sodium hypochlorite, when packed.

(a) When bleach bottles are opened, an easily identified odour of chlorine tells you that some chlorine has escaped. Explain what this observation tells you about the tendency of aqueous hypochlorite ions to react, and why the bottle labels must use the phrase, "when packed," when listing the minimum concentration value of sodium hypochlorite.
(b) What minimum mass of $NaOCl_{(s)}$ was required to make the bleach solution contained in the large 3.6-L jug?
(c) If your main purpose is to kill microorganisms, you could just dilute some regular bleach and wipe your cutting board, countertop or stove surface. How many times more NaOCl solute, by mass, is in the large jug than is in the spray bottle?
(d) Find some bleach solution containers at home or in the supermarket. Try to find large and small ones with the same brand name, such as Clorox or Javex. Record the volumes, concentration, and prices; compare the two by calculating the cost per gram of NaOCl in each container.
(e) The shopper notes that both bottles are labelled as corrosive, and marked with the Hazardous Household Product Symbol (HHPS) in a triangular frame. What is the symbol used, and what does the triangular frame mean?
(f) The labels also say "Dangerous Gas Formed When Mixed with Acid." Write the net ionic equation for hydrogen ions reacting with hypochlorite ions, $OCl^-_{(aq)}$, and chloride ions, to form chlorine gas and water. Explain why this reaction as written cannot be considered a simple Brønsted-Lowry acid–base reaction.

Exploring

38. On a large jug of household bleach, the words CAUTION and ATTENTION, and a HHPS corrosive logo are in large print, in black, on a white background. On a smaller spray bottle of a more dilute solution, the same words and logo are used, but in small print, in light green, on a slightly darker green background. Both containers also state "Dangerous Gas Formed When Mixed with Acid"—which is a *very* important caution—so that people do not try to clean their sinks and toilets by mixing acids (such as rust-removing cleaners) with bleach.

(a) Use the Internet to determine what the legal limit is for exposure to chlorine gas in air, and what effect breathing chlorine has on humans.
(b) Use the Internet to determine what the Canadian legal standards are (if any) for hazard warnings on product labels. Suggest some changes in the standards that would improve safety.

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