Strong and Weak Acids and Bases

10.2

Re-examine Figure 10.1 on page xxx. Look at the photographs of hydrochloric acid and acetic acid. The conductivity tester is testing the same concentrations of both acids. As you can see, the bulb glows brightly in the hydrochloric acid. The bulb glows dimly in the acetic acid. How can these different results be explained?

Section Preview/ Specific Expectations

In this section, you will

- **explain**, in terms of the degree to which they dissociate, the difference between strong and weak acids and bases
- **distinguish** between binary acids and oxoacids
- **define** pH, and experimentally determine the effect on pH of diluting an acidic solution
- **communicate** your understanding of the following terms: *strong acid, weak acid, strong base, weak base, binary acid, oxoacid, pH*

Strong Acids and Weak Acids

You know that ions are present in an aqueous solution of an acid. These ions result from the dissociation of the acid. An acid that dissociates completely into ions in water is called a **strong acid.** For example, hydrochloric acid is a strong acid. *All* the molecules of hydrochloric acid in an aqueous solution dissociate into H^+ and Cl^- ions. The H^+ ions, as you know, bond with surrounding water molecules to form hydronium ions, H_3O^+. (See Figure 10.6.) *The concentration of hydronium ions in a dilute solution of a strong acid is equal to the concentration of the acid.* Thus, a 1.0 mol/L solution of hydrochloric acid contains 1.0 mol/L of hydronium ions. Table 10.4 lists the strong acids.

Table 10.4 Strong Acids

hydrochloric acid, HCl
hydrobromic acid, HBr
hydroiodic acid, HI
nitric acid, HNO_3
sulfuric acid, H_2SO_4
perchloric acid, $HClO_4$

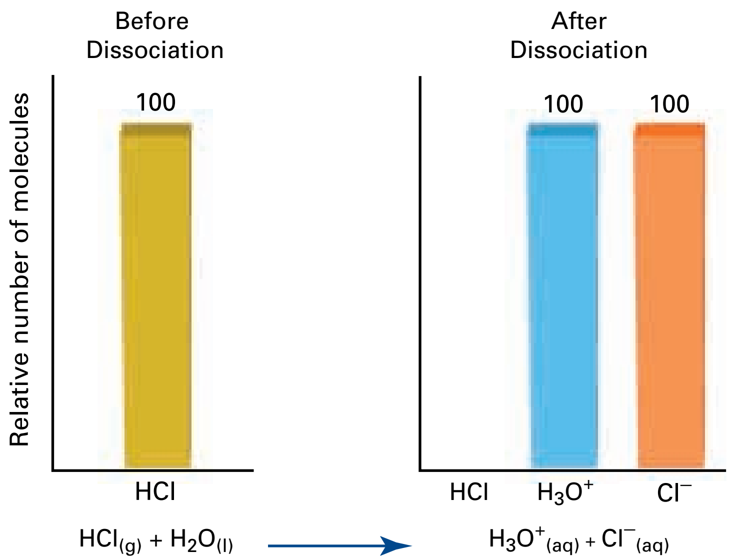


Figure 10.6 **When hydrogen chloride molecules enter an aqueous solution, 100% of the hydrogen chloride molecules dissociate. As a result, the solution contains the same percent of H^+ ions (in the form of H_3O^+) and Cl^- ions: 100%.**

A **weak acid** is an acid that dissociates very slightly in a water solution. Thus, only a small percentage of the acid molecules break apart into ions. Most of the acid molecules remain intact. For example, acetic acid is a weak acid. On average, only about 1% (one in a hundred) of the acetic acid molecules dissociate at any given moment in a 0.1 mol/L solution. (The number of acid molecules that dissociate depends on the concentration and temperature of the solution.) In fact, *the concentration of hydronium ions in a solution of a weak acid is always less than the concentration of the dissolved acid.* (See Figure 10.7.)

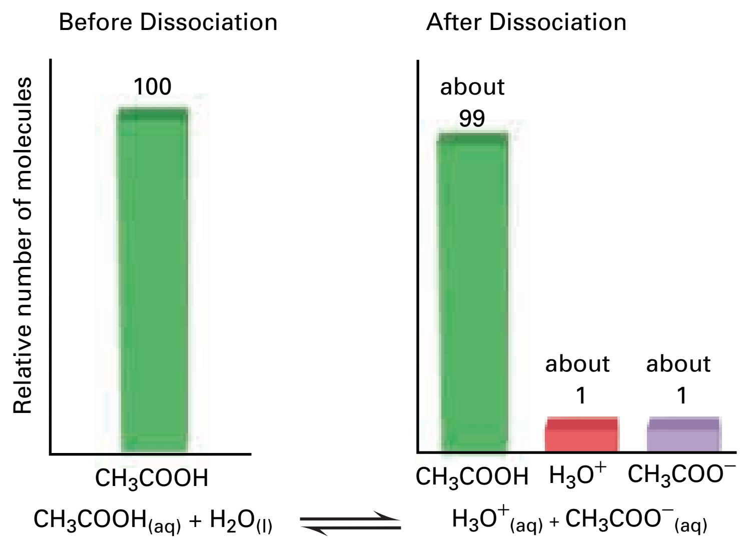


Figure 10.7 **When acetic acid molecules enter an aqueous solution, only about 1% of them dissociate. Thus, the number of acetic acid molecules in solution is far greater than the number of hydronium ions and acetate ions.**

Notice the arrow that is used in the equation in Figure 10.7. It points in both directions, indicating that the reaction is *reversible.* In other words, the products of the reaction also react to produce the original reactants. In this reaction, molecules of acetic acid dissociate just as quickly and as often as the dissociated ions re-associate to produce acetic acid molecules. (Figure 10.8 will help you visualize what happens.)

Most acids are weak acids. Whenever you see a reversible chemical equation involving an acid, you can safely assume that the acid is weak.

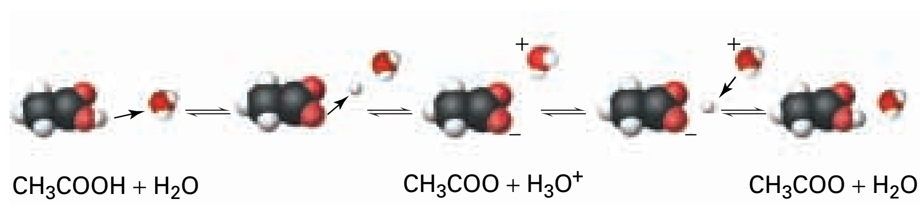


Figure 10.8 **When acetic acid dissolves in water, acetic acid molecules dissociate and re-associate at the same time and at the same rate.**

A few acids contain only a single hydrogen ion that can dissociate. These acids are called *monoprotic acids.* (The prefix mono- means "one." The root -protic refers to "proton.") Hydrochloric acid, hydrobromic acid, and hydroiodic acid are strong monoprotic acids. Hydrofluoric acid, HF, is weak monoprotic acid.

Many acids contain two or more hydrogen ions that can dissociate. For example, sulfuric acid, $H_2SO_{4(aq)}$, has two hydrogen ions that can dissociate. As you know from Table 10.4, sulfuric acid is a strong acid. This is true only for its first dissociation, however.

$$H_2SO_{4(aq)} \rightarrow H^+{}_{(aq)} + HSO_4{}^-{}_{(aq)}$$

The resulting aqueous hydrogen sulfate ion, $HSO^-{}_4$, is a weak acid. It dissociates to form the sulfate ion in the following reversible reaction:

$$HSO_4{}^-{}_{(aq)} \rightleftharpoons H^+{}_{(aq)} + SO_4{}^{2-}{}_{(aq)}$$

Thus, acids that contain two hydrogen ions dissociate to form two anions. These acids are sometimes called *diprotic acids.* (The prefix di-, as you know, means "two.") The acid that is formed by the first dissociation is stronger than the acid that is formed by the second dissociation.

Acids that contain three hydrogen ions are called *triprotic acids.* Phosphoric acid, $H_3PO_{4(aq)}$, is a triprotic acid. It gives rise to three anions, as follows:

$$H_3PO_{4(aq)} + H_2O_{(\ell)} \rightleftharpoons H^+_{(aq)} + H_2PO_4{}^-_{(aq)}$$
$$H_2PO_4{}^-_{(aq)} + H_2O_{(\ell)} \rightleftharpoons H^+_{(aq)} + HPO_4{}^{2-}_{(aq)}$$
$$HPO_4{}^{2-}_{(aq)} + H_2O_{(\ell)} \rightleftharpoons H^+_{(aq)} + PO_4{}^{3-}_{(aq)}$$

Here again, the acid that is formed by the first dissociation is stronger than the acid that is formed by the second dissociation. This acid is stronger than the acid that is formed by the third dissociation. Keep in mind, however, that all three of these acids are weak because only a small proportion of them dissociates.

Strong Bases and Weak Bases

Like a strong acid, a **strong base** dissociates completely into ions in water. All oxides and hydroxides of the alkali metals—Group 1 (IA)—are strong bases. The oxides and hydroxides of the alkaline earth metals—Group 2 (IIA)—below beryllium are also strong bases.

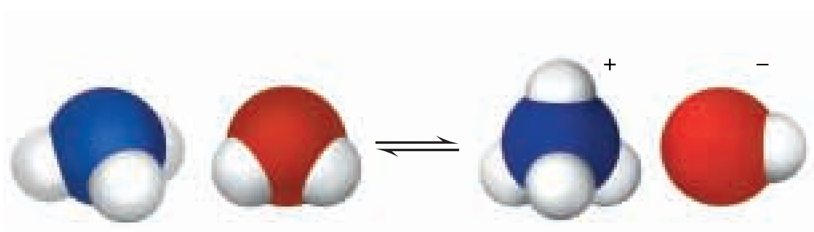
Recall that the concentration of hydronium ions in a dilute solution of a strong acid is equal to the concentration of the acid. Similarly, the concentration of hydroxide ions in a dilute solution of a strong base is equal to the concentration of the base. For example, a 1.0 mol/L solution of sodium hydroxide (a strong base) contains 1.0 mol/L of hydroxide ions.

Table 10.5 lists some common strong bases. Barium hydroxide, $Ba(OH)_2$, and strontium hydroxide, $Sr(OH)_2$, are strong bases that are soluble in water. Magnesium oxide, MgO, and magnesium hydroxide, $Mg(OH)_2$, are also strong bases, but they are considered to be insoluble. A small amount of these compounds does dissolve in water, however. Virtually all of this small amount dissociates completely.

Most bases are weak. A **weak base** dissociates very slightly in a water solution. The most common weak base is aqueous ammonia. In a 0.1 mol/L solution, only about 1% of the ammonia molecules react with water to form hydroxide ions. This reversible reaction is represented in Figure 10.9.

Table 10.5 Common Strong Bases

sodium hydroxide, NaOH
potassium hydroxide, KOH
calcium hydroxide, $Ca(OH)_2$
strontium hydroxide, $Sr(OH)_2$
barium hydroxide, $Ba(OH)_2$



$$NH_3 + H_2O \rightleftharpoons NH_4{}^+ + OH^-$$

Figure 10.9 **Ammonia does not contain hydroxide ions, so it is not an Arrhenius base. As you can see, however, an ammonia molecule can remove a proton from water, leaving a hydroxide ion behind. Thus, ammonia is a Brønsted-Lowry weak base.**

CHECKPOINT

The terms "strong" and "concentrated" have very different meanings when describing solutions of acids, bases, or salts. Similarly, the terms "weak" and "dilute" have very different meanings. In your own words, summarize the difference between "strong" and "concentrated." Give examples that illustrate this difference. Then summarize the difference between "weak" and "dilute," again giving examples.

mind STRETCH

The conjugate base of a strong acid is always a weak base. Conversely, the conjugate base of a weak acid is always a strong base. Explain this inverse relationship.

Naming Acids and Their Anions

There are two main kinds of acids: binary acids and oxoacids. A **binary acid** is composed of two elements: hydrogen and a non-metal. Two examples of binary acids are hydrofluoric acid and hydrochloric acid. All binary acids have the general formula $HX_{(aq)}$. The H represents one or more hydrogen atoms. The X represents the non-metal. As you can see in Table 10.6, the names of binary acids are made up of the following parts:

- the prefix hydro-
- a root that is formed from the name of the non-metal
- the suffix -ic
- the word "acid" at the end

Table 10.6 Examples of Naming Binary Acids

Binary acid	Prefix	Non-metal root	Suffix
hydrofluoric acid, $HF_{(aq)}$	hydro-	-fluor-	-ic
hydrochloric acid, $HCl_{(aq)}$	hydro-	-chlor-	-ic
hydrosulfuric acid, H_2S	hydro-	-sulfur-	-ic

As you know, anions are formed when binary acids dissociate. The names of these anions end in the suffix -ide. For example, hydroflouric acid forms the anion fluoride, F^-. Hydrochloric acid forms the anion chloride, Cl^-.

An **oxoacid** is an acid formed from a polyatomic ion that contains oxygen, hydrogen, and another element. (Oxoacids are called oxyacids in some chemistry textbooks). In Chapter 3, you learned the names of common polyatomic ions and their valences (oxidation numbers). The names of oxoacids are similar to the names of their polyatomic oxoanions. Only the suffix is different. Study the three rules and examples for naming oxoacids below. Then try the Practice Problems that follow.

1. For anions that end in -ate, the suffix of the acid is -ic. For example, the acid of the chlorate anion ClO_3^-, is chloric acid, $HClO_3$.
2. For anions that end in -ite, the suffix of the acid is -ous. For example, the acid of the chlorite anion, ClO_2^-, is chlorous acid, $HClO_2$.
3. The prefixes hypo- and per- remain as part of the acid name. For example, the acid of the perchlorate anion, ClO_4^-, is perchloric acid, $HClO_4$. The acid of the hypochlorite anion, ClO^-, is hypochlorous acid, HClO.

Practice Problems

4. (a) Write the chemical formula for hydrobromic acid. Then write the name and formula for the anion that it forms.
   (b) Hydrosulfuric acid, H_2S, forms two anions. Name them and write their formulas.
5. Write the chemical formulas for the following acids. Then name and write the formulas for the oxoanions that form from each acid. Refer to Chapter 3, Table 3.5, Names and Valences of Some Common Polyatomic Ions, as necessary.
   (a) nitric acid
   (b) nitrous acid
   (c) hyponitrous acid
   (d) phosphoric acid
   (e) phosphorous acid
   (f) periodic acid

Describing Acid and Base Strength Quantitatively: pH

You are probably familiar with the term "pH" from a variety of sources. Advertisers talk about the "pH balance" of products such as soaps, shampoos, and skin creams. People who own aquariums and swimming pools must monitor the pH of the water. (See Figure 10.10.) Gardeners and farmers use simple tests to determine the pH of the soil. They know that plants and food crops grow best within a narrow range of pH. Similarly, the pH of your blood must remain within narrow limits for you to stay healthy.

Figure 10.10 **Maintaining a safe environment for an aquarium or a swimming pool requires measuring the pH of the water and knowing how to adjust it.**

pH is clearly related to health, and to the proper functioning of products and systems. (Notice that the "p" is always lower case, even at the start of a sentence.) What exactly is pH? How is it measured? To answer these questions, consider a familiar substance: water.

The Power of Hydrogen in Water

As you know, all aqueous solutions contain ions. Even pure water contains a few ions that are produced by the dissociation of water molecules. **Remember:** The double arrow in the equation shows that the reaction is reversible. The ions recombine to form water molecules.

$$H_2O_{(\ell)} + H_2O_{(\ell)} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$$

On average, at 25°C, only about two water molecules in a billion are dissociated at any given moment. As you know, it is the ions in solution that conduct electricity. If there are virtually no ions, no electricity is conducted. This is why pure water is such a poor conductor. Chemists have determined that the concentration of hydronium ions in neutral water, at 25°C, is only 1.0×10^{-7} mol/L. The dissociation of water also produces the same, very small number of hydroxide ions. Therefore, the concentration of hydroxide ions is also 1.0×10^{-7} mol/L.

Chemists sometimes use square brackets around a chemical formula. This shorthand notation means "the concentration of" the chemical inside the brackets. For example, $[H_3O^+]$ is read as "the concentration of hydronium ions." Thus, the concentration of hydronium ions and hydroxide ions in neutral water can be written as

the concentration of hydronium ions

$$[H_3O^+] = [OH^-] = 1.0 \times 10^{-7}\ \text{mol/L}$$

the concentration of hydroxide ions

Compared with neutral water, acidic solutions contain a higher concentration of hydronium ions. Basic solutions contain a lower concentration of hydronium ions. Therefore, the dissociation of water provides another way of thinking about acids and bases. *An acid is any compound that increases [H_3O^+] when it is dissolved in water. A base is any compound that increases [OH^-] when it is dissolved in water.* (See Figure 10.11.)

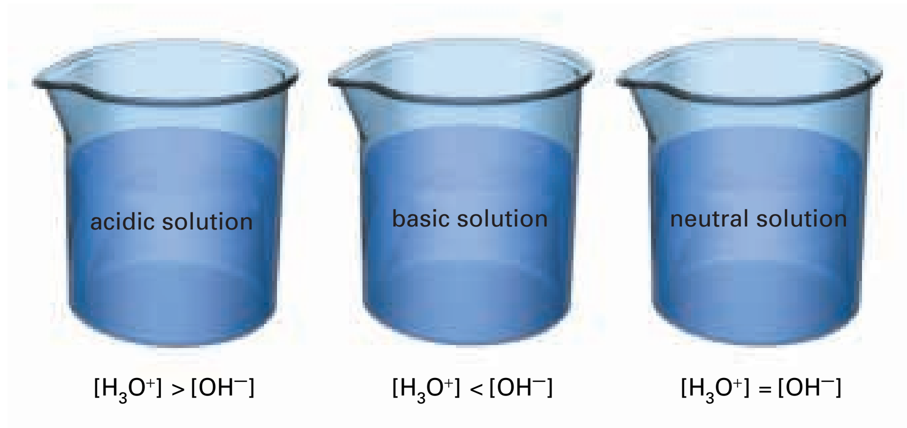


Figure 10.11 The relationship between the concentrations of hydronium ions and hydroxide ions in a solution determines whether the solution is acidic, basic, or neutral.

The pH Scale: Measuring by Powers of Ten

The concentration of hydronium ions ranges from about 10 mol/L for a concentrated strong acid to about 10^{-15} mol/L for a concentrated strong base. This wide range of concentrations, and the negative powers of 10, are not very convenient to work with. In 1909, a Danish biochemist, Søren Sørensen, suggested a method for converting concentrations to positive numbers. His method involved using the numerical system of logarithms.

The logarithm of a number is the power to which you must raise 10 to equal that number. For example, the logarithm of 10 is 1, because $10^1 = 10$. The logarithm of 100 is 2, because $10^2 = 100$. (See Appendix E for more information about exponents and logarithms.)

Sørensen defined **pH** as –log [H^+]. Since Sørensen did not know about hydronium ions, his definition of pH is based on Arrhenius' hydrogen ion. Many chemistry references reinterpret the H so that it refers to the Brønsted-Lowry hydronium ion, H_3O^+, instead. This textbook adopts the hydronium ion usage. Thus, the definition for pH becomes $pH = -\log [H_3O^+]$. Recall, though, that chemists use [H^+] as a shorthand notation for [H_3O^+]. As a result, both equations give the same product.

As you can see in Figure 10.12, the "p" in pH stands for the word "power." The power referred to is exponential power: the power of 10. The "H" stands for the concentration of hydrogen ions (or H_3O^+ ions), measured in mol/L.

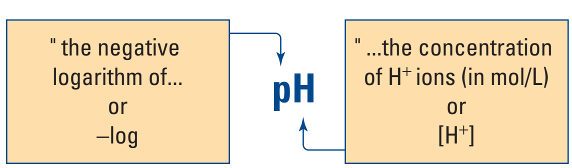


Figure 10.12 The concept of pH makes working with very small values, such as 0.000 000 000 000 01, much easier.

The concept of pH allows hydronium (or hydrogen) ion concentrations to be expressed as positive numbers, rather than negative exponents. For example, recall that $[H_3O^+]$ of neutral water at 25°C is 1.0×10^{-7}.

$$\begin{aligned} \therefore \text{pH} &= -\log [H_3O^+] \\ &= -\log (1.0 \times 10^{-7}) \\ &= -(-7.00) \\ &= 7.00 \end{aligned}$$

$[H_3O^+]$ in acidic solutions is greater than $[H_3O^+]$ in neutral water. For example, if $[H_3O^+]$ in an acid is 1.0×10^{-4} mol/L, this is 1000 times greater than $[H_3O^+]$ in neutral water. Use Table 10.7 to make sure that you understand why. The pH of the acid is 4.00. All acidic solutions have a pH that is less than 7.

Table 10.7 Understanding pH

Range of acidity and basicity	$[H_3O^+]$ (mol/L)	Exponential notation (mol/L)	log	pH ($-\log [H_3O^+]$)
strong acid	1	1×10^{0}	0	0
	0.1	1×10^{-1}	−1	1
	0.01	1×10^{-2}	−2	2
	0.001	1×10^{-3}	−3	3
	0.000 1	1×10^{-4}	−4	4
	0.000 01	1×10^{-5}	−5	5
	0.000 001	1×10^{-6}	−6	6
neutral $[H^+] = [OH^-]$ $= 1.0 \times 10^{-7}$	0.000 000 1	1×10^{-7}	−7	7
	0.000 000 01	1×10^{-8}	−8	8
	0.000 000 001	1×10^{-9}	−9	9
	0.000 000 000 1	1×10^{-10}	−10	10
	0.000 000 000 01	1×10^{-11}	−11	11
	0.000 000 000 001	1×10^{-12}	−12	12
	0.000 000 000 000 1	1×10^{-13}	−13	13
strong base	0.000 000 000 000 01	1×10^{-14}	−14	14

$[H_3O^+]$ in basic solutions is less than $[H_3O^+]$ in pure water. For example, if $[H_3O^+]$ in a base is 1.0×10^{-11} mol/L, this is 10 000 times less than $[H_3O^+]$ in neutral water. The pH of the base is 11.00. All basic solutions have a pH that is greater than 7.

The relationship among pH, $[H_3O^+]$, and the strength of acids and bases is summarized in the Concept Connection on the next page. Use the following Sample Problem and Practice Problems to assess your understanding of this relationship. Then, in Investigation 10-A, you will look for a pattern involving the pH of a strong acid, a weak acid, and dilutions of both.

Science LINK

Using logarithms is a convenient way to count a wide range of values by powers of 10. Chemists are not the only scientists who use such logarithms, however. Audiologists (scientists who study human hearing) use logarithms, too. Research the decibel scale to find out how it works. Present your findings in the medium of your choice.

Math LINK

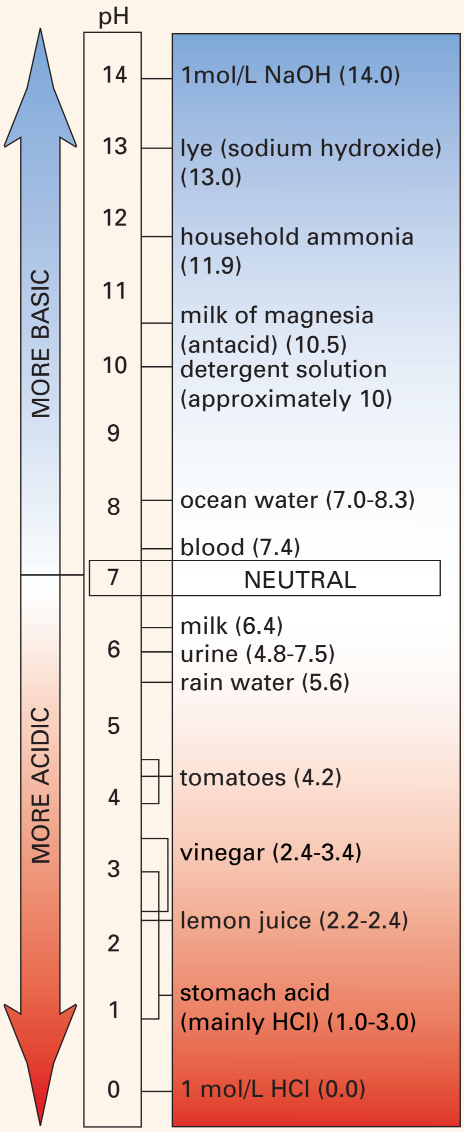
How do you determine the number of significant digits in a pH? You count only the digits to the right of the decimal point. For example, suppose that the concentration of hydronium ions in a sample of orange juice is 2.5×10^{-4} mol/L. This number has two significant digits: the 2 and the 5. The power of 10 only tells us where to place the decimal: 0.000 25. The pH of the sample is $-\log (2.5 \times 10^{-4}) = 3.602\ 059$. The digit to the left of the decimal (the 3) is derived from the power of 10. Therefore, it is not considered to be a significant digit. Only the two digits to the right of the decimal are significant. Thus, the pH value is rounded off to 3.60.

Identify the significant digits in each pH value below.

1. The pH of drain cleaner is 13.1.
2. The pH of milk is 6.4.
3. The pH of vinegar is 2.85.
4. The pH of lemon juice is 2.310.

Concept Organizer — pH, $[H_3O^+]$, and the Strength of Acids and Bases

Type of solution	$[H_3O^+]$ (mol/L)	Concentration of hydronium and hydroxide ions	pH at 25°C
acidic solution	greater than 1×10^{-7}	$[H_3O^+] > [OH^-]$	< 7.00
neutral solution	1×10^{-7}	$[H_3O^+] = [OH^-]$	7.00
basic solution	less than 1×10^{-7}	$[H_3O^+] < [OH^-]$	> 7.00



The pH values of many common solutions fall within a range from 0 to 14, as shown on this pH scale. The table above the pH scale relates the positive pH values to their hydronium ion concentrations and their logarithms.

Sample Problem

Calculating the pH of a Solution

Problem

Calculate the pH of a solution with $[H_3O^+] = 3.8 \times 10^{-3}$ mol/L.

What Is Required?

You need to calculate the pH, given $[H_3O^+]$.

What Is Given?

You know that $[H_3O^+]$ is 3.8×10^{-3} mol/L.

Plan Your Strategy

Use the equation $pH = -\log [H_3O^+]$ to solve for the unknown.

Act on Your Strategy

$$\begin{aligned} pH &= -\log (3.8 \times 10^{-3}) \\ &= 2.42 \end{aligned}$$

Check Your Solution

$[H_3O^+]$ is greater than 1.0×10^{-7} mol/L. Therefore, the pH should be less than 7.00. The solution is acidic, as you would expect.

Practice Problems

6. Calculate the pH of each solution, given the hydronium ion concentration.
- **(a)** $[H_3O^+] = 0.0027$ mol/L
- **(b)** $[H_3O^+] = 7.28 \times 10^{-8}$ mol/L
- **(c)** $[H_3O^+] = 9.7 \times 10^{-5}$ mol/L
- **(d)** $[H_3O^+] = 8.27 \times 10^{-12}$

7. $[H_3O^+]$ in a cola drink is about 5.0×10^{-3} mol/L. Calculate the pH of the drink. State whether the drink is acidic or basic.

8. A glass of orange juice has $[H_3O^+]$ of 2.9×10^{-4} mol/L. Calculate the pH of the juice. State whether the result is acidic or basic.

9. (a) $[H_3O^+]$ in a dilute solution of nitric acid, HNO_3, is 6.3×10^{-3} mol/L. Calculate the pH of the solution.
- **(b)** $[H_3O^+]$ of a solution of sodium hydroxide is 6.59×10^{10} mol/L. Calculate the pH of the solution.

PROBLEM TIP

Appendix E, "Math and Chemistry", explains how you can do these calculations with a calculator.

Investigation 10-A

SKILL FOCUS

Predicting

Performing and recording

Analyzing and interpreting

The Effect of Dilution on the pH of an Acid

In this investigation, you will compare the effects of diluting a strong acid and a weak acid.

In Part 1, you will measure the pH of a strong acid. Then you will perform a series of ten-fold dilutions. That is, each solution will be one-tenth as dilute as the previous solution. You will measure and compare the pH after each dilution.

In Part 2, you will measure the pH of a weak acid with the same initial concentration as the strong acid. Then you will perform a series of ten-fold dilutions with the weak acid. Again, you will measure and compare the pH after each dilution.

Problem

How does the pH of dilutions of a strong acid compare with the pH of dilutions of a weak acid?

Prediction

Predict each pH, and explain your reasoning.

(a) the pH of 0.10 mol/L hydrochloric acid

(b) the pH of the hydrochloric acid after one ten-fold dilution

(c) the pH of the hydrochloric acid after each of six more ten-fold dilutions

(d) the pH of 0.10 mol/L acetic acid, compared with the pH of 0.10 mol/L hydrochloric acid

(e) the pH of the acetic acid after one ten-fold dilution

Safety Precautions

Hydrochloric acid is corrosive. Wash any spills on skin or clothing with plenty of cool water. Inform your teacher immediately.

Materials

100 mL graduated cylinder
100 mL beaker
2 beakers (250 mL)
universal indicator paper and glass rod
pH meter
0.10 mol/L hydrochloric acid (for Part 1)
0.10 mol/L acetic acid (for Part 2)
distilled water

Procedure

Part 1 The pH of Solutions of a Strong Acid

1. Copy the table below into your notebook. Record the pH you predicted for each dilution.
2. Pour about 40 mL of 0.10 mol/L hydrochloric acid into a clean, dry 100 mL beaker. Use the end of a glass rod to transfer a drop of solution to a piece of universal pH paper into the acid. Compare the colour against the colour chart to determine the pH. Record the pH. Then measure and record the pH of the acid using a pH meter. Rinse the electrode with distilled water afterward.

Data Table for Part 1

$[HCl_{(aq)}]$ mol/L	Predicted pH	pH measured with universal indicator	pH measured with pH meter
1×10^{-1}			
1×10^{-2}			
1×10^{-3}			
1×10^{-4}			
1×10^{-5}			
1×10^{-6}			
1×10^{-7}			
1×10^{-8}			

3. Measure 90 mL of distilled water in a 100 mL graduated cylinder. Add 10 mL of the acid from step 2. The resulting 100 mL of solution is one-tenth as concentrated as the acid from step 2. Pour the dilute solution into a clean, dry 250 mL beaker. Use universal pH paper and a pH meter to measure the pH. Record your results.
4. Repeat step 3. Pour the new dilute solution into a second clean, dry beaker. Dispose of the more concentrated acid solution as directed by your teacher. Rinse and dry the beaker so you can use it for the next dilution.
5. Make further dilutions and pH measurements until the hydrochloric acid solution is 1.0×10^{-8} mol/L

Part 2 The pH of Solutions of a Weak Acid

1. Design a table to record your predictions and measurements for 0.10 mol/L and 0.010 mol/L concentrations of acetic acid.
2. Use the same procedure that you used in Part 1 to measure and record the pH of a 0.10 mol/L sample of acetic acid. Then dilute the solution to 0.010 mol/L. Measure the pH again.

Analysis

1. Which do you think gave the more accurate pH: the universal indicator paper or the pH meter? Explain.
2. For the strong acid, compare the pH values you predicted with the measurements you made. How can you explain any differences for the first few dilutions?
3. What was the pH of the solution that had a concentration of 1.0×10^{-8} mol/L? Explain the pH you obtained.
4. Compare the pH of 0.10 mol/L acetic acid with the pH of 0.10 mol/L hydrochloric acid. Why do you think the pH values are different, even though the concentrations of the acids were the same?
5. What effect does a ten-fold dilution of a strong acid (hydrochloric acid) have on the pH of the acid? What effect does the same dilution of a weak acid (acetic acid) have on its pH? Compare the effects for a strong acid and a weak acid. Account for any differences.

Conclusion

6. Use evidence from your investigation to support the conclusion that a weak acid ionizes less than a strong acid of identical concentration.
7. Why is the method for calculating the pH of a strong acid (if it is not too dilute) not appropriate for a weak acid?

Applications

8. Nicotinic acid is a B vitamin. The pH of a 0.050 mol/L solution of this acid is measured to be 3.08. Is it a strong acid or a weak acid? Explain. What would be the pH of a solution of nitric acid having the same concentration?
9. Would you expect to be able to predict the pH of a weak base, given its concentration? Explain. Design an experiment you could perform to check your answer.

Chemistry Bulletin

Science Technology Society Environment

The Chemistry of Oven Cleaning

Oven cleaning is not a job that most people enjoy. Removing baked-on grease from inside an oven requires serious scrubbing. Any chemical oven cleaners that help to make the job easier are usually welcome. Like all chemicals, however, the most effective oven cleaners require attention to safety.

Cleaners that contain strong bases are the most effective for dissolving grease and grime. Bases are effective because they produce soaps when they react with the fatty acids in grease. When a strong base (such as sodium hydroxide, $NaOH$, or potassium hydroxide, KOH) is used on a dirty oven, the fat molecules that make up the grease are split into smaller molecules. Anions from the base then bond with some of these molecules to form soap.

One end of a soap molecule is non-polar (uncharged), so it is soluble in dirt and grease, which are also non-polar. The other end of a soap molecule is polar (charged), so it is soluble in water. Because of its two different properties, soap acts like a "bridge" between the grease and the water. Soap enables grease to dissolve in water and be washed away, thus allowing the cleaner to remove the grease from the oven surface.

Cleaners that contain sodium hydroxide and potassium hydroxide are very effective. They are also caustic and potentially very dangerous. For example, sodium hydroxide, in the concentrations that are used in oven cleaners, can irritate the skin and cause blindness if it gets in the eyes. As well, it is damaging to paints and fabrics.

There are alternatives to sodium hydroxide and other strong base cleaners. One alternative involves using ammonia, NH_3, which is a weak base. If a bowl of dilute ammonia solution is placed in an oven and left for several hours, most of the grease and grime can be wiped off.

Ammonia does not completely ionize in water. Only a small portion dissociates. Although an ammonia solution is less caustic than sodium hydroxide, it can be toxic if inhaled directly. As well, ammonia vapours can cause eye, lung, and skin irritations. At higher concentrations, ammonia can be extremely toxic.

Baking soda is a non-toxic alternative, but it is much less effective. Therefore, it requires even more scrubbing. An abrasive paste can be made by mixing baking soda and water. The basic properties of baking soda also have a small effect on grease and grime if it is applied to the oven and left for several hours.

Making Connections

1. Survey the cleaners in your home or school. Which cleaners contain bases and which contain acids? What cleaning jobs can an acid cleaner perform well? How do most acid cleaners work?
2. Some companies claim to make environmentally sensitive cleaners. Investigate these cleaners. What chemicals do they contain? See if you can infer how they work. You might like to design a controlled experiment to test the effectiveness of several oven cleaners. **CAUTION** Obtain permission from your teacher before performing such an experiment.

Section Wrap-up

In this section, you considered the relationship among the strength of acids and bases, the concentration of hydronium and hydroxide ions, and pH. Much of the time, you examined acids and bases acting independently of each other. However, acids and bases often interact. In fact, acid-base reactions have many important applications in the home, as well as in the laboratory. In section 10.3, you will investigate acid-base reactions.

Section Review

1 K/U Distinguish, in terms of degree of dissociation, between a strong acid and a weak acid, and a strong base and a weak base.

2 K/U Give one example of the following:

(a) a weak acid

(b) a strong acid

(c) a strong base

(d) a weak base

3 K/U Formic acid, HCOOH, is responsible for the painful bites of fire ants. Is formic acid strong or weak? Explain.

4 K/U $KMnO_4$ is an intense purple-coloured solid that can be made into a solution to kill bacteria. What is the name of this compound? Give the name and the formula of the acid that forms when $KMnO_4$ combines with water.

5 K/U State the name or the formula for the acid that forms from each of the following anions:

(a) hydrogen sulfate

(b) F^-

(c) HS^-

(d) bromite

6 K/U Explain the meaning of pH, both in terms of hydrogen ions and hydronium ions.

7 K/U Arrange the following foods in order of increasing acidity: beets, pH = 5.0; camembert cheese, pH = 7.4; egg white, pH = 8.0; sauerkraut, pH = 3.5; yogurt, pH = 4.5.

8 I Calculate the pH of each body fluid, given the concentration of hydronium ions.

(a) tears, $[H_3O^+] = 4.0 \times 10^{-8}$ mol/L

(b) stomach acid, $[H_3O^+] = 4.0 \times 10^{-2}$ mol/L

9 I Calculate the pH of the solution that is formed by diluting 50 mL of 0.025 mol/L hydrochloric acid to a final volume of 1.0 L

10 C What is $[H_3O^+]$ in a solution with pH = 0? Why do chemists not usually use pH to describe $[H_3O^+]$ when the pH value would be a negative number?

Unit Issue Prep

As you investigate the contamination of Prince Edward Island's soils with sodium arsenite from pesticides, investigate the link between the pH and solubility. For example, water polluted with sodium arsenite may be treated with lime (calcium oxide), CaO. What is the purpose of this treatment?