

4.7: Acid Base Reactions

Learning Objectives

- To know the characteristic properties of acids and bases

Acid–base reactions are essential in both biochemistry and industrial chemistry. Moreover, many of the substances we encounter in our homes, the supermarket, and the pharmacy are acids or bases. For example, aspirin is an acid (acetylsalicylic acid), and antacids are bases. In fact, every amateur chef who has prepared mayonnaise or squeezed a wedge of lemon to marinate a piece of fish has carried out an acid–base reaction. Before we discuss the characteristics of such reactions, let's first describe some of the properties of acids and bases.

Definitions of Acids and Bases

In [Chapter 4.6](#), we defined **acids** as substances that dissolve in water to produce H^+ ions, whereas **bases** were defined as substances that dissolve in water to produce OH^- ions. In fact, this is only one possible set of definitions. Although the general properties of acids and bases have been known for more than a thousand years, the definitions of *acid* and *base* have changed dramatically as scientists have learned more about them. In ancient times, an acid was any substance that had a sour taste (e.g., vinegar or lemon juice), caused consistent color changes in dyes derived from plants (e.g., turning blue litmus paper red), reacted with certain metals to produce hydrogen gas and a solution of a salt containing a metal cation, and dissolved carbonate salts such as limestone ($CaCO_3$) with the evolution of carbon dioxide. In contrast, a base was any substance that had a bitter taste, felt slippery to the touch, and caused color changes in plant dyes that differed diametrically from the changes caused by acids (e.g., turning red litmus paper blue). Although these definitions were useful, they were entirely descriptive.

Acid
Sour taste
Turns blue litmus red
reacts with some metals to produce H_2
Dissolves carbonate salts, releasing CO_2

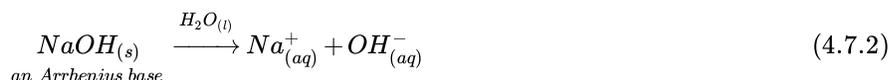


Base
Bitter taste
Turns red litmus blue
Slippery to the touch



The Arrhenius Definition of Acids and Bases

The first person to define acids and bases in detail was the Swedish chemist Svante Arrhenius (1859–1927; Nobel Prize in Chemistry, 1903). According to the *Arrhenius definition*, an acid is a substance like hydrochloric acid that dissolves in water to produce H^+ ions (protons; Equation 4.7.1), and a base is a substance like sodium hydroxide that dissolves in water to produce hydroxide (OH^-) ions (Equation 4.7.2):



According to Arrhenius, the characteristic properties of acids and bases are due exclusively to the presence of H^+ and OH^- ions, respectively, in solution. Although Arrhenius's ideas were widely accepted, his definition of acids and bases had two major limitations:

1. First, because acids and bases were defined in terms of ions obtained from water, the Arrhenius concept applied only to substances in aqueous solution.

2. Second, and more important, the Arrhenius definition predicted that *only* substances that dissolve in water to produce H^+ and OH^- ions should exhibit the properties of acids and bases, respectively. For example, according to the Arrhenius definition, the reaction of ammonia (a base) with gaseous HCl (an acid) to give ammonium chloride (Equation 4.7.3) is not an acid–base reaction because it does not involve H^+ and OH^- :



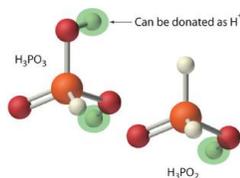
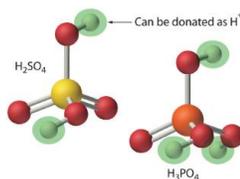
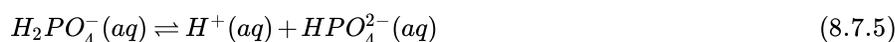
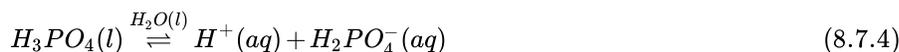
The Brønsted–Lowry Definition of Acids and Bases

Because of the limitations of the Arrhenius definition, a more general definition of acids and bases was needed. One was proposed independently in 1923 by the Danish chemist J. N. Brønsted (1879–1947) and the British chemist T. M. Lowry (1874–1936), who defined acid–base reactions in terms of the transfer of a proton (H^+ ion) from one substance to another.

According to Brønsted and Lowry, an acid (A substance with at least one hydrogen atom that can dissociate to form an anion and an H^+ ion (a proton) in aqueous solution, thereby forming an acidic solution) is any substance that can donate a proton, and a base (a substance that produces one or more hydroxide ions (OH^-) and a cation when dissolved in aqueous solution, thereby forming a basic solution) is any substance that can accept a proton. The Brønsted–Lowry definition of an acid is essentially the same as the Arrhenius definition, except that it is not restricted to aqueous solutions. The Brønsted–Lowry definition of a base, however, is far more general because the hydroxide ion is just one of many substances that can accept a proton. Ammonia, for example, reacts with a proton to form NH_4^+ , so in Equation 4.7.3, NH_3 is a Brønsted–Lowry base and HCl is a Brønsted–Lowry acid. Because of its more general nature, the Brønsted–Lowry definition is used throughout this text unless otherwise specified.

Polyprotic Acids

Acids differ in the number of protons they can donate. For example, monoprotic acids (a compound that is capable of donating one proton per molecule) are compounds that are capable of donating a single proton per molecule. Monoprotic acids include HF, HCl, HBr, HI, HNO_3 , and HNO_2 . All carboxylic acids that contain a single $-CO_2H$ group, such as acetic acid (CH_3CO_2H), are monoprotic acids, dissociating to form RCO_2^- and H^+ (section 4.6). A compound that can donate more than one proton per molecule. can donate more than one proton per molecule. For example, H_2SO_4 can donate two H^+ ions in separate steps, so it is a diprotic acid (a compound that can donate two protons per molecule in separate steps) and H_3PO_4 , which is capable of donating three protons in successive steps, is a triprotic acid (a compound that can donate three protons per molecule in separate steps), (Equation 4.7.4, Equation 4.7.5, and Equation 4.7.6):

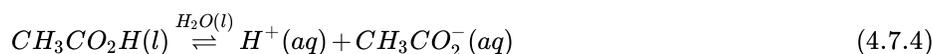


In chemical equations such as these, a double arrow is used to indicate that both the forward and reverse reactions occur simultaneously, so the forward reaction does not go to completion. Instead, the solution contains significant amounts of both reactants and products. Over time, the reaction reaches a state in which the concentration of each species in solution remains constant. The reaction is then said to be in equilibrium (the point at which the rates of the forward and reverse reactions become the same, so that the net composition of the system no longer changes with time).

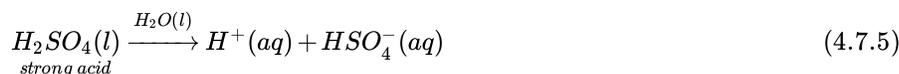
Strengths of Acids and Bases

We will not discuss the strengths of acids and bases quantitatively until next semester. Qualitatively, however, we can state that strong acids (An acid that reacts essentially completely with water) to give H^+ and the corresponding anion. react essentially completely with water to give H^+ and the corresponding anion. Similarly, strong bases (A base that dissociates essentially completely in water) to give OH^- and the corresponding cation) dissociate essentially completely in water to give OH^- and the corresponding cation. Strong acids and strong bases are both strong electrolytes. In contrast, only a fraction of the molecules of weak acids (An acid in which only a fraction of the molecules react with water) to produce H^+ and the corresponding anion. and weak bases (A base in which only a fraction of the molecules react with water to produce OH^- and the corresponding cation) react with water to produce ions, so weak acids and weak bases are also weak electrolytes. Typically less than 5% of a weak electrolyte dissociates into ions in solution, whereas more than 95% is present in undissociated form.

In practice, only a few strong acids are commonly encountered: HCl, HBr, HI, HNO_3 , $HClO_4$, and H_2SO_4 (H_3PO_4 is only moderately strong). The most common strong bases are ionic compounds that contain the hydroxide ion as the anion; three examples are NaOH, KOH, and $Ca(OH)_2$. Common weak acids include HCN, H_2S , HF, oxoacids such as HNO_2 and HClO, and carboxylic acids such as acetic acid. The ionization reaction of acetic acid is as follows:

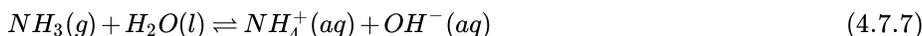


Although acetic acid is *very* soluble in water, almost all of the acetic acid in solution exists in the form of neutral molecules (less than 1% dissociates), as we stated in [section 4.1](#). Sulfuric acid is unusual in that it is a strong acid when it donates its first proton (Equation 4.7.8) but a weak acid when it donates its second proton ([Equation 8.7.9](#)) as indicated by the single and double arrows, respectively:



Consequently, an aqueous solution of sulfuric acid contains $H^+_{(aq)}$ ions and a mixture of $HSO_4^-_{(aq)}$ and $SO_4^{2-}_{(aq)}$ ions, but no H_2SO_4 molecules.

The most common weak base is ammonia, which reacts with water to form small amounts of hydroxide ion:



Most of the ammonia (>99%) is present in the form of $NH_3(g)$. Amines, which are organic analogues of ammonia, are also weak bases, as are ionic compounds that contain anions derived from weak acids (such as S^{2-}).

Table 4.7.1 lists some common strong acids and bases. Acids other than the six common strong acids are almost invariably weak acids. The only common strong bases are the hydroxides of the alkali metals and the heavier alkaline earths (Ca, Sr, and Ba); any other bases you encounter are most likely weak. Remember that *there is no correlation between solubility and whether a substance is a strong or a weak electrolyte!* Many weak acids and bases are extremely soluble in water.

Note the Pattern

There is *no* correlation between the solubility of a substance and whether it is a strong electrolyte, a weak electrolyte, or a nonelectrolyte.

Table 4.7.1 Common Strong Acids and Bases

Strong Acids		Strong Bases	
Hydrogen Halides	Oxoacids	Group 1 Hydroxides	Hydroxides of the Heavier Group 2 Elements
HCl	HNO_3	LiOH	$Ca(OH)_2$
HBr	H_2SO_4	NaOH	$Sr(OH)_2$
HI	$HClO_4$	KOH	$Ba(OH)_2$

Strong Acids		Strong Bases	
Hydrogen Halides	Oxoacids	Group 1 Hydroxides	Hydroxides of the Heavier Group 2 Elements
		RbOH	
		CsOH	

✓ Example 4.7.1

Classify each compound as a strong acid, a weak acid, a strong base, a weak base, or none of these.

1. $\text{CH}_3\text{CH}_2\text{CO}_2\text{H}$
2. CH_3OH
3. $\text{Sr}(\text{OH})_2$
4. $\text{CH}_3\text{CH}_2\text{NH}_2$
5. HBrO_4

Given: compound

Asked for: acid or base strength

Strategy:

A Determine whether the compound is organic or inorganic.

B If inorganic, determine whether the compound is acidic or basic by the presence of dissociable H^+ or OH^- ions, respectively. If organic, identify the compound as a weak base or a weak acid by the presence of an amine or a carboxylic acid group, respectively. Recall that all polyprotic acids except H_2SO_4 are weak acids.

Solution

1. **A** This compound is propionic acid, which is organic. **B** It contains a carboxylic acid group analogous to that in acetic acid, so it must be a weak acid.
2. **A** CH_3OH is methanol, an organic compound that contains the $-\text{OH}$ group. **B** As a covalent compound, it does not dissociate to form the OH^- ion. Because it does not contain a carboxylic acid ($-\text{CO}_2\text{H}$) group, methanol also cannot dissociate to form $\text{H}^+(\text{aq})$ ions. Thus we predict that in aqueous solution methanol is neither an acid nor a base.
3. **A** $\text{Sr}(\text{OH})_2$ is an inorganic compound that contains one Sr^{2+} and two OH^- ions per formula unit. **B** We therefore expect it to be a strong base, similar to $\text{Ca}(\text{OH})_2$.
4. **A** $\text{CH}_3\text{CH}_2\text{NH}_2$ is an amine (ethylamine), an organic compound in which one hydrogen of ammonia has been replaced by an R group. **B** Consequently, we expect it to behave similarly to ammonia (Equation 8.6.7), reacting with water to produce small amounts of the OH^- ion. Ethylamine is therefore a weak base.
5. **A** HBrO_4 is perbromic acid, an inorganic compound. **B** It is not listed in Table 8.3 as one of the common strong acids, but that does not necessarily mean that it is a weak acid. If you examine the periodic table, you can see that Br lies directly below Cl in group 17. We might therefore expect that HBrO_4 is chemically similar to HClO_4 , a strong acid—and, in fact, it is.

? Exercise 4.7.1

Classify each compound as a strong acid, a weak acid, a strong base, a weak base, or none of these.

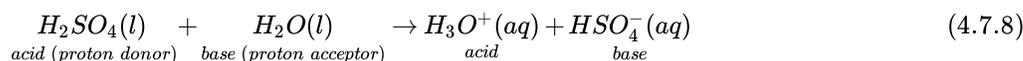
1. $\text{Ba}(\text{OH})_2$
2. HIO_4
3. $\text{CH}_3\text{CH}_2\text{CH}_2\text{CO}_2\text{H}$
4. $(\text{CH}_3)_2\text{NH}$
5. CH_2O

Answer

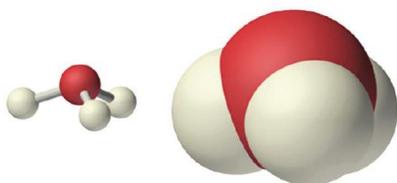
1. strong base
2. strong acid
3. weak acid
4. weak base
5. none of these; formaldehyde is a neutral molecule

The Hydronium Ion

Because isolated protons are very unstable and hence very reactive, an acid never simply “loses” an H^+ ion. Instead, the proton is always transferred to another substance, which acts as a base in the Brønsted–Lowry definition. Thus in every acid–base reaction, one species acts as an acid and one species acts as a base. Occasionally, the same substance performs both roles, as you will see later. When a strong acid dissolves in water, the proton that is released is transferred to a water molecule that acts as a proton acceptor or base, as shown for the dissociation of sulfuric acid:

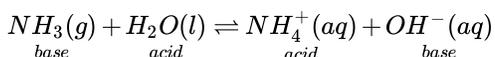


Technically, therefore, it is imprecise to describe the dissociation of a strong acid as producing $H^+_{(aq)}$ ions, as we have been doing. The resulting H_3O^+ ion, called the hydronium ion is a more accurate representation of $H^+_{(aq)}$. For the sake of brevity, however, in discussing acid dissociation reactions, we often show the product as $H^+_{(aq)}$ (as in Equation 4.7.7) with the understanding that the product is actually the $H_3O^+_{(aq)}$ ion.



Hydronium ion,
 H_3O^+

Conversely, bases that do not contain the hydroxide ion accept a proton from water, so small amounts of OH^- are produced, as in the following:



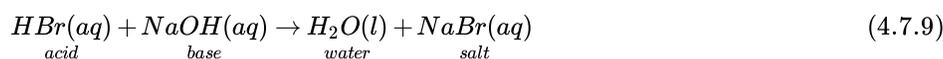
Again, the double arrow indicates that the reaction does not go to completion but rather reaches a state of equilibrium. In this reaction, water acts as an acid by donating a proton to ammonia, and ammonia acts as a base by accepting a proton from water. Thus water can act as either an acid or a base by donating a proton to a base or by accepting a proton from an acid. Substances that can behave as both an acid and a base are said to be amphoteric. When substances can behave as both an acid and a base.

The products of an acid–base reaction are also an acid and a base. In Equation 4.7.11, for example, the products of the reaction are the hydronium ion, here an acid, and the hydrogen sulfate ion, here a weak base. In Equation 4.7.12 the products are NH_4^+ , an acid, and OH^- , a base. The product NH_4^+ is called the conjugate acid. The substance formed when a Brønsted–Lowry base accepts a proton. of the base NH_3 , and the product OH^- is called the conjugate base. The substance formed when a Brønsted–Lowry acid donates a proton. of the acid H_2O . Thus all acid–base reactions actually involve two conjugate acid–base pairs. An acid and a base that differ by only one hydrogen ion. All acid–base reactions involve two conjugate acid–base pairs, the Brønsted–Lowry acid and the base it forms after donating its proton, and the Brønsted–Lowry base and the acid it forms after accepting a proton.; in Equation 4.7.12, they are NH_4^+/NH_3 and H_2O/OH^- .

Neutralization Reactions

A neutralization reaction (a chemical reaction in which an acid and a base react in stoichiometric amounts to produce water and a salt) is one in which an acid and a base react in stoichiometric amounts to produce water and a salt (the general term for any ionic substance that does not have OH^- as the anion or H^+ as the cation), the general term for any ionic substance that does not have OH^- as the anion or H^+ as the cation. If the base is a metal hydroxide, then the general formula for the reaction of an acid with a

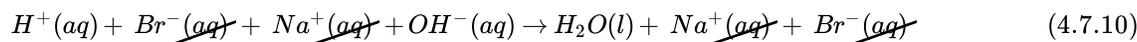
base is described as follows: *Acid plus base yields water plus salt*. For example, the reaction of equimolar amounts of HBr and NaOH to give water and a salt (NaBr) is a neutralization reaction:



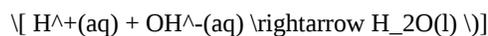
Note the Pattern

Acid plus base yields water plus salt.

If we write the complete ionic equation for the reaction in Equation 4.7.13, we see that $\text{Na}^+_{(\text{aq})}$ and $\text{Br}^-_{(\text{aq})}$ are spectator ions and are not involved in the reaction:

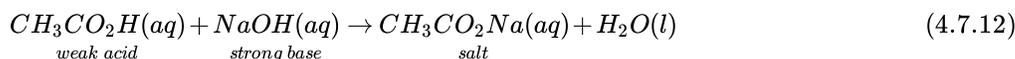


The overall reaction is therefore simply the combination of $\text{H}^+(\text{aq})$ and $\text{OH}^-(\text{aq})$ to produce H_2O , as shown in the net ionic equation:

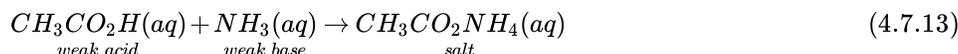


The net ionic equation for the reaction of any strong acid with any strong base is identical to Equation 4.7.15.

The strengths of the acid and the base generally determine whether the reaction goes to completion. The reaction of *any* strong acid with *any* strong base goes essentially to completion, as does the reaction of a strong acid with a weak base, and a weak acid with a strong base. Examples of the last two are as follows:



Sodium acetate is written with the organic component first followed by the cation, as is usual for organic salts. Most reactions of a weak acid with a weak base also go essentially to completion. One example is the reaction of acetic acid with ammonia:



An example of an acid–base reaction that does not go to completion is the reaction of a weak acid or a weak base with water, which is both an extremely weak acid and an extremely weak base. We will discuss these reactions in more detail in Chapter 16

Note the Pattern

Except for the reaction of a weak acid or a weak base with water, acid–base reactions essentially go to completion.

In some cases, the reaction of an acid with an anion derived from a weak acid (such as HS^-) produces a gas (in this case, H_2S). Because the gaseous product escapes from solution in the form of bubbles, the reverse reaction cannot occur. Therefore, these reactions tend to be forced, or driven, to completion. Examples include reactions in which an acid is added to ionic compounds that contain the HCO_3^- , CN^- , or S^{2-} anions, all of which are driven to completion (Figure 4.7.1):





Figure 4.7.1 The Reaction of Dilute Aqueous HCl with a Solution of Na_2CO_3 Note the vigorous formation of gaseous CO_2 .

The reactions in Equation 4.7.21 are responsible for the rotten egg smell that is produced when metal sulfides come in contact with acids.

✓ Example 4.7.2

Calcium propionate is used to inhibit the growth of molds in foods, tobacco, and some medicines. Write a balanced chemical equation for the reaction of aqueous propionic acid ($\text{CH}_3\text{CH}_2\text{CO}_2\text{H}$) with aqueous calcium hydroxide [$\text{Ca}(\text{OH})_2$] to give calcium propionate. Do you expect this reaction to go to completion, making it a feasible method for the preparation of calcium propionate?

Given: reactants and product

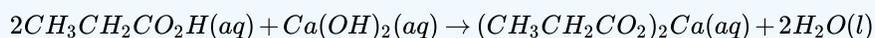
Asked for: balanced chemical equation and whether the reaction will go to completion

Strategy:

Write the balanced chemical equation for the reaction of propionic acid with calcium hydroxide. Based on their acid and base strengths, predict whether the reaction will go to completion.

Solution

Propionic acid is an organic compound that is a weak acid, and calcium hydroxide is an inorganic compound that is a strong base. The balanced chemical equation is as follows:

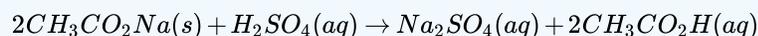


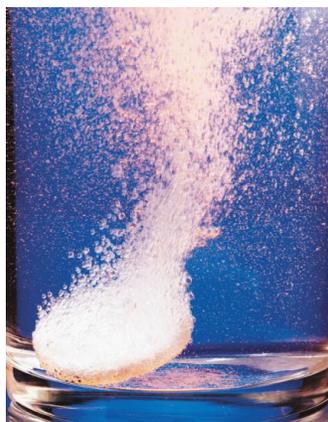
The reaction of a weak acid and a strong base will go to completion, so it is reasonable to prepare calcium propionate by mixing solutions of propionic acid and calcium hydroxide in a 2:1 mole ratio.

? Exercise 4.7.2

Write a balanced chemical equation for the reaction of solid sodium acetate with dilute sulfuric acid to give sodium sulfate.

Answer





Stomach acid. An antacid tablet reacts with 0.1 M HCl (the approximate concentration found in the human stomach).

One of the most familiar and most heavily advertised applications of acid–base chemistry is *antacids*, which are bases that neutralize stomach acid. The human stomach contains an approximately 0.1 M solution of hydrochloric acid that helps digest foods. If the protective lining of the stomach breaks down, this acid can attack the stomach tissue, resulting in the formation of an *ulcer*. Because one factor that is believed to contribute to the formation of stomach ulcers is the production of excess acid in the stomach, many individuals routinely consume large quantities of antacids. The active ingredients in antacids include sodium bicarbonate and potassium bicarbonate (NaHCO₃ and KHCO₃; Alka-Seltzer); a mixture of magnesium hydroxide and aluminum hydroxide [Mg(OH)₂ and Al(OH)₃; Maalox, Mylanta]; calcium carbonate (CaCO₃; Tums); and a complex salt, dihydroxyaluminum sodium carbonate [NaAl(OH)₂CO₃; original Roloids]. Each has certain advantages and disadvantages. For example, Mg(OH)₂ is a powerful laxative (it is the active ingredient in milk of magnesia), whereas Al(OH)₃ causes constipation. When mixed, each tends to counteract the unwanted effects of the other. Although all antacids contain both an anionic base (OH[−], CO₃^{2−}, or HCO₃[−]) and an appropriate cation, they differ substantially in the amount of active ingredient in a given mass of product.

✓ Example 4.7.3

Assume that the stomach of someone suffering from acid indigestion contains 75 mL of 0.20 M HCl. How many Tums tablets are required to neutralize 90% of the stomach acid, if each tablet contains 500 mg of CaCO₃? (Neutralizing all of the stomach acid is not desirable because that would completely shut down digestion.)

Given: volume and molarity of acid and mass of base in an antacid tablet

Asked for: number of tablets required for 90% neutralization

Strategy:

A Write the balanced chemical equation for the reaction and then decide whether the reaction will go to completion.

B Calculate the number of moles of acid present. Multiply the number of moles by the percentage to obtain the quantity of acid that must be neutralized. Using mole ratios, calculate the number of moles of base required to neutralize the acid.

C Calculate the number of moles of base contained in one tablet by dividing the mass of base by the corresponding molar mass. Calculate the number of tablets required by dividing the moles of base by the moles contained in one tablet.

Solution

A We first write the balanced chemical equation for the reaction:



Each carbonate ion can react with 2 mol of H⁺ to produce H₂CO₃, which rapidly decomposes to H₂O and CO₂. Because HCl is a strong acid and CO₃^{2−} is a weak base, the reaction will go to completion.

B Next we need to determine the number of moles of HCl present:

$$75 \text{ mL} \left(\frac{1 \cancel{\text{L}}}{1000 \cancel{\text{mL}}} \right) \left(\frac{0.20 \text{ mol HCl}}{\cancel{\text{L}}} \right) = 0.015 \text{ mol HCl}$$

Because we want to neutralize only 90% of the acid present, we multiply the number of moles of HCl by 0.90:

$$(0.015 \text{ mol HCl})(0.90) = 0.014 \text{ mol HCl}$$

We know from the stoichiometry of the reaction that each mole of CaCO_3 reacts with 2 mol of HCl, so we need

$$\text{moles CaCO}_3 = 0.014 \text{ mol HCl} \left(\frac{1 \text{ mol CaCO}_3}{2 \text{ mol HCl}} \right) = 0.0070 \text{ mol CaCO}_3$$

C Each Tums tablet contains

$$\left(\frac{500 \text{ mg CaCO}_3}{1 \text{ Tums tablet}} \right) \left(\frac{1 \text{ g}}{1000 \text{ mg CaCO}_3} \right) \left(\frac{1 \text{ mol CaCO}_3}{100.1 \text{ g}} \right) = 0.00500 \text{ mol CaCO}_3$$

Thus we need $\frac{0.0070 \text{ mol CaCO}_3}{0.00500 \text{ mol CaCO}_3} = 1.4$ Tums tablets.

? Exercise 4.7.3

Assume that as a result of overeating, a person's stomach contains 300 mL of 0.25 M HCl. How many Rolaids tablets must be consumed to neutralize 95% of the acid, if each tablet contains 400 mg of $\text{NaAl(OH)}_2\text{CO}_3$? The neutralization reaction can be written as follows:



Answer

6.4 tablets

The pH Scale

One of the key factors affecting reactions that occur in dilute solutions of acids and bases is the concentration of H^+ and OH^- ions. The pH scale, a logarithmic scale used to express the hydrogen ion (H^+) concentration of a solution, making it possible to describe acidity or basicity quantitatively, provides a convenient way of expressing the hydrogen ion (H^+) concentration of a solution and enables us to describe acidity or basicity in quantitative terms.

Pure liquid water contains extremely low but measurable concentrations of $\text{H}_3\text{O}^+(aq)$ and $\text{OH}^-(aq)$ ions produced via an *autoionization reaction*, in which water acts simultaneously as an acid and as a base:



The concentration of hydrogen ions in pure water is only 1.0×10^{-7} M at 25°C. Because the autoionization reaction produces both a proton and a hydroxide ion, the OH^- concentration in pure water is also 1.0×10^{-7} M. Pure water is a neutral solution, a solution in which the total positive charge from all the cations is matched by an identical total negative charge from all the anions, in which $[\text{H}^+] = [\text{OH}^-] = 1.0 \times 10^{-7}$ M.

The pH scale describes the hydrogen ion concentration of a solution in a way that avoids the use of exponential notation; pH, the negative base-10 logarithm of the hydrogen ion concentration: $\text{pH} = -\log[\text{H}^+]$ is defined as the negative base-10 logarithm of the hydrogen ion concentration. pH is actually defined as the negative base-10 logarithm of hydrogen ion *activity*. As you will learn in a more advanced course, the activity of a substance in solution is related to its concentration. For dilute solutions such as those we are discussing, the activity and the concentration are approximately the same.

$$\text{pH} = -\log[\text{H}^+] \quad (4.7.19)$$

Conversely,

$$[\text{H}^+] = 10^{-\text{pH}}$$

Because the hydrogen ion concentration is 1.0×10^{-7} M in pure water at 25°C, the pH of pure liquid water (and, by extension, of any neutral solution) is

$$pH = -\log[1.0 \times 10^{-7}] = 7.00 \quad (4.7.20)$$

Adding an acid to pure water increases the hydrogen ion concentration and decreases the hydroxide ion concentration because a neutralization reaction occurs, such as that shown in [Equation 8.7.15](#). Because the negative exponent of $[H^+]$ becomes smaller as $[H^+]$ increases, the pH *decreases* with increasing $[H^+]$. For example, a 1.0 M solution of a strong monoprotic acid such as HCl or HNO_3 has a pH of 0.00:

$$pH = -\log[1.0] = 0.00 \quad (4.7.21)$$

Note the Pattern

pH decreases with increasing $[H^+]$.

Conversely, adding a base to pure water increases the hydroxide ion concentration and decreases the hydrogen ion concentration. Because the autoionization reaction of water does not go to completion, neither does the neutralization reaction. Even a strongly basic solution contains a detectable amount of H^+ ions. For example, a 1.0 M OH^- solution has $[H^+] = 1.0 \times 10^{-14}$ M. The pH of a 1.0 M NaOH solution is therefore

$$pH = -\log[1.0 \times 10^{-14}] = 14.00 \quad (4.7.22)$$

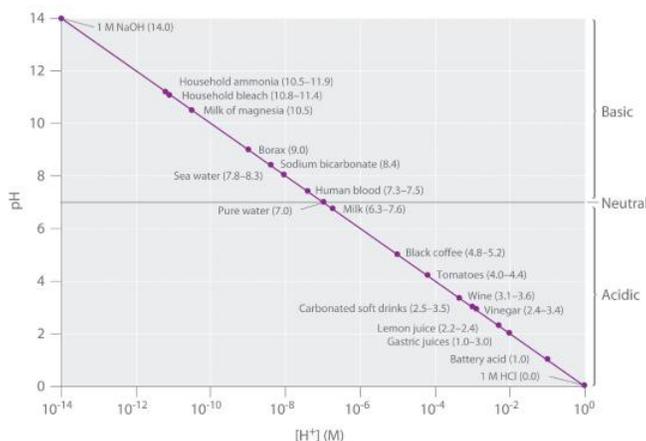
For practical purposes, the pH scale runs from pH = 0 (corresponding to 1 M H^+) to pH 14 (corresponding to 1 M OH^-), although pH values less than 0 or greater than 14 are possible.

We can summarize the relationships between acidity, basicity, and pH as follows:

- If pH = 7.0, the solution is neutral.
- If pH < 7.0, the solution is acidic.
- If pH > 7.0, the solution is basic.

Keep in mind that the pH scale is logarithmic, so a change of 1.0 in the pH of a solution corresponds to a *tenfold* change in the hydrogen ion concentration. The foods and consumer products we encounter daily represent a wide range of pH values, as shown in [Figure 8.7.2](#).

Figure 8.7.2 A Plot of pH versus $[H^+]$ for Some Common Aqueous Solutions



Although many substances exist in a range of pH values (indicated in parentheses), they are plotted using typical values.

✓ Example 4.7.4

1. What is the pH of a 2.1×10^{-2} M aqueous solution of $HClO_4$?
2. The pH of a vinegar sample is 3.80. What is its hydrogen ion concentration?

Given: molarity of acid or pH

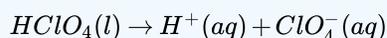
Asked for: pH or $[H^+]$

Strategy:

Using the balanced chemical equation for the acid dissociation reaction and Equation 4.7.24 or 4.7.25, determine $[H^+]$ and convert it to pH or vice versa.

Solution

1. $HClO_4$ (perchloric acid) is a strong acid, so it dissociates completely into H^+ ions and ClO_4^- ions:



The H^+ ion concentration is therefore the same as the perchloric acid concentration. The pH of the perchloric acid solution is thus

$$pH = -\log[H^+] = -\log(2.1 \times 10^{-2}) = 1.68$$

The result makes sense: the H^+ ion concentration is between 10^{-1} M and 10^{-2} M, so the pH must be between 1 and 2.

Note: The assumption that $[H^+]$ is the same as the concentration of the acid is valid for *only* strong acids. Because weak acids do *not* dissociate completely in aqueous solution, a more complex procedure is needed to calculate the pH of their solutions.

2. We are given the pH and asked to calculate the hydrogen ion concentration. From Equation 4.7.24

$$10^{-pH} = [H^+]$$

Thus $[H^+] = 10^{-3.80} = 1.6 \times 10^{-4}$ M.

? Exercise 4.7.4

1. What is the pH of a 3.0×10^{-5} M aqueous solution of HNO_3 ?
2. What is the hydrogen ion concentration of turnip juice, which has a pH of 5.41?

Answer

1. $pH = 4.52$
2. $[H^+] = 3.9 \times 10^{-6}$ M

Tools have been developed that make the measurement of pH simple and convenient (Figure 8.6.3). For example, pH paper consists of strips of paper impregnated with one or more acid–base indicators. An intensely colored organic molecule whose color changes dramatically depending on the pH of the solution., which are intensely colored organic molecules whose colors change dramatically depending on the pH of the solution. Placing a drop of a solution on a strip of pH paper and comparing its color with standards give the solution's approximate pH. A more accurate tool, the pH meter, uses a glass electrode, a device whose voltage depends on the H^+ ion concentration.

Figure 8.6.3 Two Ways of Measuring the pH of a Solution: pH Paper and a pH Meter

Note that both show that the pH is 1.7, but the pH meter gives a more precise value.

Key Equations

definition of pH

Equation 4.7.231: $pH = -\log[H^+]$

Equation 4.7.24: $[H^+] = 10^{-pH}$

Summary

Acid–base reactions require both an acid and a base. In Brønsted–Lowry terms, an **acid** is a substance that can donate a proton (H^+), and a **base** is a substance that can accept a proton. All acid–base reactions contain two acid–base pairs: the reactants and the products. Acids can donate one proton (**monoprotic acids**), two protons (**diprotic acids**), or three protons (**triprotic acids**). Compounds that are capable of donating more than one proton are generally called **polyprotic acids**. Acids also differ in their tendency to donate a proton, a measure of their acid strength. **Strong acids** react completely with water to produce $H_3O^+(aq)$ (the **hydronium ion**), whereas **weak acids** dissociate only partially in water. Conversely, **strong bases** react completely with water to

produce the hydroxide ion, whereas **weak bases** react only partially with water to form hydroxide ions. The reaction of a strong acid with a strong base is a **neutralization reaction**, which produces water plus a **salt**.

The acidity or basicity of an aqueous solution is described quantitatively using the **pH scale**. The **pH** of a solution is the negative logarithm of the H^+ ion concentration and typically ranges from 0 for strongly acidic solutions to 14 for strongly basic ones. Because of the *autoionization reaction* of water, which produces small amounts of hydronium ions and hydroxide ions, a **neutral solution** of water contains 1×10^{-7} M H^+ ions and has a pH of 7.0. An **indicator** is an intensely colored organic substance whose color is pH dependent; it is used to determine the pH of a solution.

Key Takeaway

- An acidic solution and a basic solution react together in a neutralization reaction that also forms a salt.

Conceptual Problems

1. Why was it necessary to expand on the Arrhenius definition of an acid and a base? What specific point does the Brønsted–Lowry definition address?
2. State whether each compound is an acid, a base, or a salt.
 - a. $CaCO_3$
 - b. $NaHCO_3$
 - c. H_2SO_4
 - d. $CaCl_2$
 - e. $Ba(OH)_2$
3. State whether each compound is an acid, a base, or a salt.
 - a. NH_3
 - b. NH_4Cl
 - c. H_2CO_3
 - d. CH_3COOH
 - e. $NaOH$
4. Classify each compound as a strong acid, a weak acid, a strong base, or a weak base in aqueous solution.
 - a. sodium hydroxide
 - b. acetic acid
 - c. magnesium hydroxide
 - d. tartaric acid
 - e. sulfuric acid
 - f. ammonia
 - g. hydroxylamine (NH_2OH)
 - h. hydrocyanic acid
5. Decide whether each compound forms an aqueous solution that is strongly acidic, weakly acidic, strongly basic, or weakly basic.
 - a. propanoic acid
 - b. hydrobromic acid
 - c. methylamine
 - d. lithium hydroxide
 - e. citric acid
 - f. sodium acetate
 - g. ammonium chloride
 - h. barium hydroxide
6. What is the relationship between the strength of an acid and the strength of the conjugate base derived from that acid? Would you expect the $CH_3CO_2^-$ ion to be a strong base or a weak base? Why? Is the hydronium ion a strong acid or a weak acid? Explain your answer.
7. What are the products of an acid–base reaction? Under what circumstances is one of the products a gas?

8. Explain how an aqueous solution that is strongly basic can have a pH, which is a measure of the *acidity* of a solution.

Answer

- 1.
- 2.
- 3.
- 4.
5.
 - a. weakly acidic
 - b. strongly acidic
 - c. weakly basic
 - d. strongly basic
 - e. weakly acidic
 - f. weakly basic
 - g. weakly acidic
 - h. strongly basic
- 6.
- 7.
- 8.

Numerical Problems

Please be sure you are familiar with the topics discussed in *Essential Skills 3 (section 4.11)* before proceeding to the Numerical Problems.

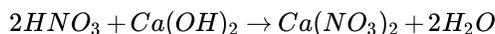
1. Derive an equation to relate the hydrogen ion concentration to the molarity of a solution of a strong monoprotic acid.
2. Derive an equation to relate the hydroxide ion concentration to the molarity of a solution of
 - a. a group I hydroxide.
 - b. a group II hydroxide.
3. Given the following salts, identify the acid and the base in the neutralization reactions and then write the complete ionic equation:
 - a. barium sulfate
 - b. lithium nitrate
 - c. sodium bromide
 - d. calcium perchlorate
4. What is the pH of each solution?
 - a. 5.8×10^{-3} mol of HNO_3 in 257 mL of water
 - b. 0.0079 mol of HI in 750 mL of water
 - c. 0.011 mol of HClO_4 in 500 mL of water
 - d. 0.257 mol of HBr in 5.00 L of water
5. What is the hydrogen ion concentration of each substance in the indicated pH range?
 - a. black coffee (pH 5.10)
 - b. milk (pH 6.30–7.60)
 - c. tomatoes (pH 4.00–4.40)
6. What is the hydrogen ion concentration of each substance in the indicated pH range?
 - a. orange juice (pH 3–4)
 - b. fresh egg white (pH 7.60–7.80)
 - c. lemon juice (pH 2.20–2.40)
7. What is the pH of a solution prepared by diluting 25.00 mL of 0.879 M HCl to a volume of 555 mL?

8. Vinegar is primarily an aqueous solution of acetic acid. Commercial vinegar typically contains 5.0 g of acetic acid in 95.0 g of water. What is the concentration of commercial vinegar? If only 3.1% of the acetic acid dissociates to CH_3CO_2^- and H^+ , what is the pH of the solution? (Assume the density of the solution is 1.00 g/mL.)
9. If a typical household cleanser is 0.50 M in strong base, what volume of 0.998 M strong monoprotic acid is needed to neutralize 50.0 mL of the cleanser?
10. A 25.00 mL sample of a 0.9005 M solution of HCl is diluted to 500.0 mL. What is the molarity of the final solution? How many milliliters of 0.223 M NaOH are needed to neutralize 25.00 mL of this final solution?
11. If 20.0 mL of 0.10 M NaOH are needed to neutralize 15.0 mL of gastric fluid, what is the molarity of HCl in the fluid? (Assume all the acidity is due to the presence of HCl.) What other base might be used instead of NaOH?
12. Malonic acid ($\text{C}_3\text{H}_4\text{O}_4$) is a diprotic acid used in the manufacture of barbiturates. How many grams of malonic acid are in a 25.00 mL sample that requires 32.68 mL of 1.124 M KOH for complete neutralization to occur? Malonic acid is a dicarboxylic acid; propose a structure for malonic acid.
13. Describe how you would prepare 500 mL of a 1.00 M stock solution of HCl from an HCl solution that is 12.11 M. Using your stock solution, how would you prepare 500 mL of a solution that is 0.012 M in HCl?
14. Given a stock solution that is 8.52 M in HBr, describe how you would prepare a 500 mL solution with each concentration.
 - a. 2.50 M
 - b. 4.00×10^{-3} M
 - c. 0.989 M
15. How many moles of solute are contained in each?
 - a. 25.00 mL of 1.86 M NaOH
 - b. 50.00 mL of 0.0898 M HCl
 - c. 13.89 mL of 0.102 M HBr
16. A chemist needed a solution that was approximately 0.5 M in HCl but could measure only 10.00 mL samples into a 50.00 mL volumetric flask. Propose a method for preparing the solution. (Assume that concentrated HCl is 12.0 M.)
17. Write the balanced chemical equation for each reaction.
 - a. perchloric acid with potassium hydroxide
 - b. nitric acid with calcium hydroxide
18. Write the balanced chemical equation for each reaction.
 - a. solid strontium hydroxide with hydrobromic acid
 - b. aqueous sulfuric acid with solid sodium hydroxide
19. A neutralization reaction gives calcium nitrate as one of the two products. Identify the acid and the base in this reaction. What is the second product? If the product had been cesium iodide, what would have been the acid and the base? What is the complete ionic equation for each reaction?

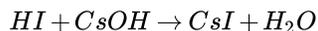
Answers

1. $[\text{H}_3\text{O}^+] = [\text{HA}] \text{ M}$
- 2.
3.
 - a. H_2SO_4 and $\text{Ba}(\text{OH})_2$; $2\text{H}^+ + \text{SO}_4^{2-} + \text{Ba}^{2+} + 2\text{OH}^- \rightarrow 2\text{H}_2\text{O} + \text{Ba}^{2+} + \text{SO}_4^{2-}$
 - b. HNO_3 and LiOH ; $\text{H}^+ + \text{NO}_3^- + \text{Li}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O} + \text{Li}^+ + \text{NO}_3^-$
 - c. HBr and NaOH ; $\text{H}^+ + \text{Br}^- + \text{Na}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O} + \text{Na}^+ + \text{Br}^-$
 - d. HClO_4 and $\text{Ca}(\text{OH})_2$; $2\text{H}^+ + 2\text{ClO}_4^- + \text{Ca}^{2+} + 2\text{OH}^- \rightarrow 2\text{H}_2\text{O} + \text{Ca}^{2+} + 2\text{ClO}_4^-$
- 4.
5.
 - a. $7.9 \times 10^{-6} \text{ M H}^+$
 - b. 5.0×10^{-7} to $2.5 \times 10^{-8} \text{ M H}^+$
 - c. 1.0×10^{-4} to $4.0 \times 10^{-5} \text{ M H}^+$
- 6.

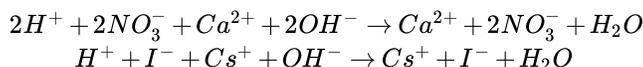
7. pH = 1.402
- 8.
9. 25 mL
- 10.
11. 0.13 M HCl; magnesium carbonate, MgCO_3 , or aluminum hydroxide, $\text{Al}(\text{OH})_3$
- 12.
13. 1.00 M solution: dilute 41.20 mL of the concentrated solution to a final volume of 500 mL. 0.012 M solution: dilute 12.0 mL of the 1.00 M stock solution to a final volume of 500 mL.
- 14.
15. a. 4.65×10^{-2} mol NaOH
 b. 4.49×10^{-3} mol HCl
 c. 1.42×10^{-3} mol HBr
- 16.
17. a. $\text{HClO}_4 + \text{KOH} \rightarrow \text{KClO}_4 + \text{H}_2\text{O}$
 b. $2\text{HNO}_3 + \text{Ca}(\text{OH})_2 \rightarrow \text{Ca}(\text{NO}_3)_2 + 2\text{H}_2\text{O}$
- 18.
19. The acid is nitric acid, and the base is calcium hydroxide. The other product is water.



The acid is hydroiodic acid, and the base is cesium hydroxide. The other product is water.



The complete ionic equations are



Contributors

- Anonymous

Modified by [Joshua Halpern \(Howard University\)](#)

[Video from Colin McKay @ YouTube](#)

This page titled [4.7: Acid Base Reactions](#) is shared under a [CC BY-NC-SA 3.0](#) license and was authored, remixed, and/or curated by [Anonymous](#).