FAU FLORIDA ATLANTIC UNIVERSITY

Lab Manual 9: Acids and Bases

General Chemistry for Health Sciences Lab

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Introduction

Acids and bases are not only important in chemistry, they are probably the most common types of solutions we encounter in our everyday life. Although we learn acids and bases in chemistry class as harmful and toxic if inhaled, the fact is acids and bases play an important role in our life. For example, many foods and drinks we consume, some medicines we take, and home remedies we use contain acids and bases.

Physiological function of the human body depends on acid-base concentration. The body's pH balance is the level of acids and bases in our blood at which our body works best. If our lungs or kidneys are not working properly, our blood's pH level can become imbalanced and this can cause medical conditions, such as acidosis and alkalosis. Another example of pH balance and physiological function is the pH of our stomach acids. Pepsin is a necessary enzyme for the digestion of protein in the stomach and can only work at very low pH. So, it is important our stomach maintains an acidic environment for digestion.

pH is used to measure the amount of hydrogen ions in the solution that shows the strength of acid and stands for potential hydrogen. The pH scale ranges from 0 to 14. A pH of 0 is considered highly acidic, while a pH of 14 is very basic. A pH of 7 is neutral. The average pH in the human blood is around 7.35-7.45. A small change in this range can affect a cell's function. Everything we do in our daily life might affect pH range, such as eating food and breathing air.

Three theories that describe acid and bases are Arrhenius, Brønsted -Lowry, and Lewis Theory. Theory and background section covers all these theories, but we will focus on learning about Arrhenius and Brønsted -Lowry theory.

Goal of Lab 9: Acids and Bases

In this lab, we will learn about three different theories that describe acid and bases. We will become familiar with the pH scale, strength of acids, calculations of pH of strong acids and bases, and will discover H^+ and OH^- concentrations in solution by using pH values. We will also learn acid-base reactions, such as neutralization reactions, which have many practical applications. Ionization reaction of pure water, called autoionization of water, will also be covered in this lab.

Theory and Background

Acid-base reactions have been studied for quite some time. In 1680, Robert Boyle reported traits of acid solutions that included their ability to dissolve many substances, to change the colors of certain natural dyes, and to lose these traits after coming in contact with alkali (base) solutions. In the eighteenth century, it was recognized that acids have a sour taste, react with limestone to liberate a gaseous substance (now known to be CO_2), and interact with alkalis to form neutral substances. In 1815, Humphry Davy contributed greatly to the development of the modern acid-base concept by demonstrating that hydrogen is the essential constituent of acids. Around that same time, Joseph Louis Gay-Lussac concluded that acids are substances that can neutralize bases and that these two classes of substances can be defined only in terms of each other. The significance of hydrogen was reemphasized in 1884 when Svante Arrhenius defined an acid as a compound that dissolves in water to yield hydrogen cations (now recognized to be hydronium ions) and a base as a compound that dissolves in water to yield hydroxide anions.

Brønsted-Lowry Acids and Bases

Johannes Brønsted and Thomas Lowry proposed a more general description in 1923 in which acids and bases were defined in terms of the transfer of hydrogen ions, H^+. (Note that these hydrogen ions are often referred to simply as *protons*, since that subatomic particle is the only component of cations derived from the most abundant hydrogen isotope, 1H.) A compound that donates a proton to another compound is called a **Brønsted-Lowry acid**, and a compound that accepts a proton is called a **Brønsted-Lowry base**. An acid-base reaction is, thus, the transfer of a proton from a donor (acid) to an acceptor (base).

The concept of *conjugate pairs* is useful in describing Brønsted-Lowry acid-base reactions (and other reversible reactions, as well). When an acid donates H^+, the species that remains is called the **conjugate base** of the acid because it reacts as a proton acceptor in the reverse reaction. Likewise, when a base accepts H^+, it is converted to its **conjugate acid**. The reaction between water and ammonia illustrates this idea. In the forward direction, water acts as an acid by donating a proton to ammonia and subsequently becoming a hydroxide ion, OH^-, the conjugate base of water. The ammonia acts as a base in accepting this proton, becoming an ammonium ion, NH_4^+, the conjugate acid of ammonia. In the reverse direction, a hydroxide ion acts as a base in accepting a proton from ammonium ion, which acts as an acid. Figure 9.1 illustrates conversion between conjugate pairs.

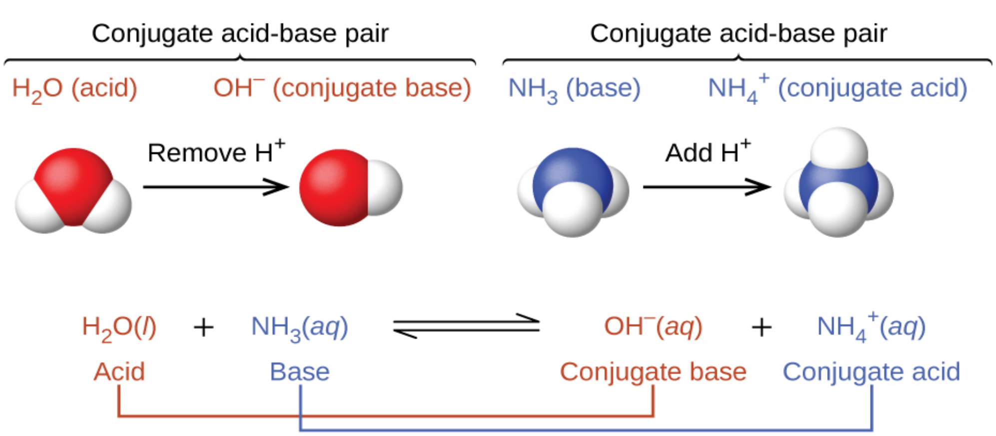


Figure 9.1 Conversion of conjugate acid-base pairs. [credit: *Chemistry 2e.* Section 14.1. OpenStax. CC BY.]

The reaction between a Brønsted-Lowry acid and water is called **acid ionization**. For example in Figure 9.2, when hydrogen fluoride dissolves in water and ionizes, protons are transferred from hydrogen fluoride molecules to water molecules, yielding hydronium ions and fluoride ions:

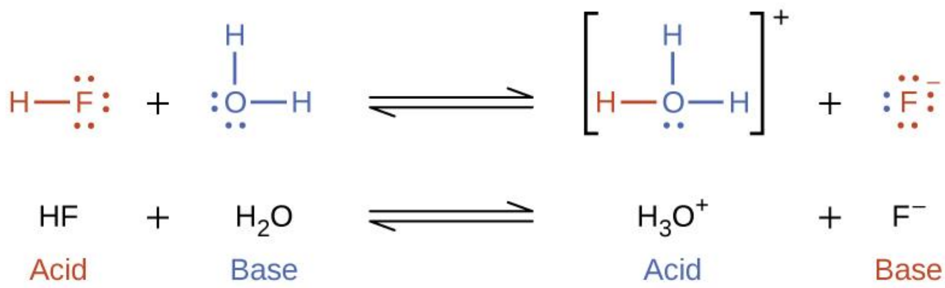


Figure 9.2 Diagram of acid ionization yielding hydronium ions and fluoride ions. [credit: *Chemistry 2e.* Section 14.1. OpenStax. CC BY.]

Base ionization of a species occurs when it accepts protons from water molecules. Shown in Figure 9.3, pyridine molecules, C_5NH_5, undergo base ionization when dissolved in water, yielding hydroxide and pyridinium ions:

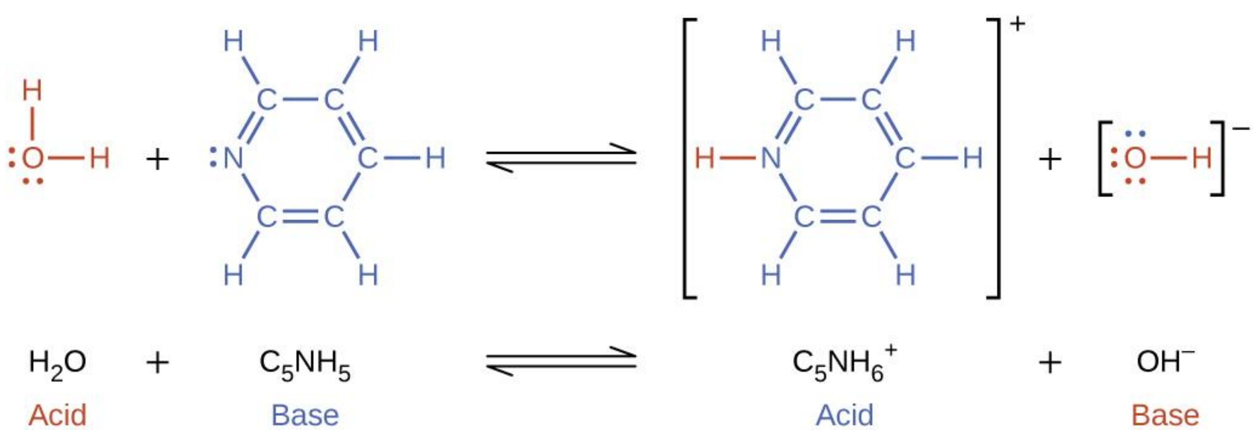


Figure 9.3 Illustration of pyridine molecules undergoing base ionization when dissolved in water, resulting in hydroxide and pyridinium ions. [credit: *Chemistry 2e.* Section 14.1. OpenStax. CC BY.]

The preceding ionization reactions suggest that water may function as both a base (as in its reaction with hydrogen fluoride) and an acid (as in its reaction with ammonia). Species capable of either donating or accepting protons are called **amphiprotic**, or more generally, **amphoteric**, a term that may be used for acids and bases per definitions other than the Brønsted-Lowry one.

The equations below show the two possible acid-base reactions for two amphiprotic species, bicarbonate ion and water:

$$\mathrm{HCO_3{}^-(aq) + H_2O(\mathit{l})}$$

$$\text{and} \quad \mathrm{CO_3{}^{2-}(aq) + H_3O^+(aq)}$$

$$\mathrm{HCO_3{}^-(aq) + H_2O(\mathit{l})}$$

$$\text{and} \quad \mathrm{H_2CO_3(aq) + OH^-(aq)}$$

The first equation represents the reaction of bicarbonate as an acid with water as a base, whereas the second represents reaction of bicarbonate as a base with water as an acid. When bicarbonate is added to water, both these equilibria are established simultaneously and the composition of the resulting solution may be determined through appropriate equilibrium calculations.

In the liquid state, molecules of an amphiprotic substance can react with one another as illustrated for water in Figure 9.4.

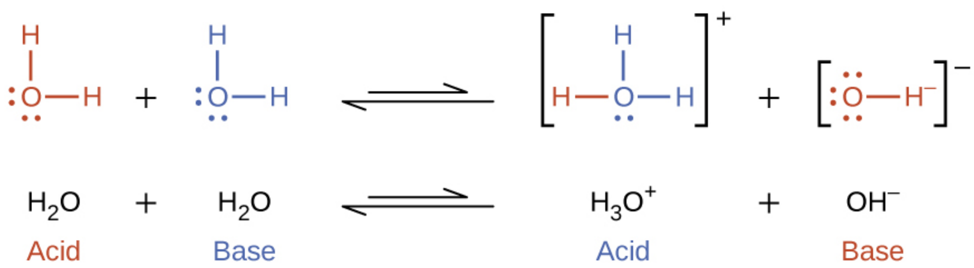


Figure 9.4 Water molecules of an amphiprotic substance reacting with one another. [credit: *Chemistry 2e.* Section 14.1. OpenStax. CC BY.]

The process in which molecules react to yield ions is called **autoionization**. Liquid water undergoes autoionization to a very slight extent; at 25 °C, approximately two out of every billion water molecules are ionized. The extent of the water autoionization process is reflected in the value of its equilibrium constant, the **ion-product constant for water, $\boldsymbol{K_\mathrm{w}}$**:

$$\mathrm{H_2O}(l) + \mathrm{H_2O}(l) \rightleftharpoons \mathrm{H_3O^+}(aq) + \mathrm{OH^-}(aq)$$

$$\text{and} \quad K_\mathrm{w} = [\mathrm{H_3O^+}][\mathrm{OH^-}]$$

The slight ionization of pure water is reflected in the small value of the equilibrium constant; at 25 °C, *K*w has a value of 1.0×10^{-14}. The process is endothermic, and so the extent of ionization and the resulting concentrations of hydronium ion and hydroxide ion increase with temperature. For example, at 100 °C, the value for *K*w is about 5.6×10^{-13}, roughly 50 times larger than the value at 25 °C.

Lewis Acids and Bases

In 1923, G. N. Lewis proposed a generalized definition of acid-base behavior in which acids and bases are identified by their ability to accept or to donate a pair of electrons and form a coordinate covalent bond.

A **coordinate covalent bond** (or dative bond) occurs when one of the atoms in the bond provides both bonding electrons. For example, a coordinate covalent bond occurs when a water molecule combines with a hydrogen ion to form a hydronium ion. A coordinate covalent bond also results when an ammonia molecule combines with a hydrogen ion to form an ammonium ion. Figure 9.5 illustrated both these equations.

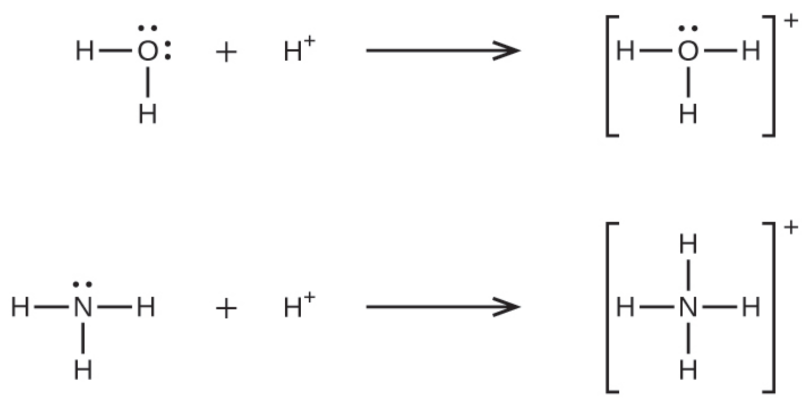


Figure 9.5 Equations for coordinate covalent bonds. [credit: *Chemistry 2e.* Section 15.2. OpenStax. CC BY.]

Reactions involving the formation of coordinate covalent bonds are classified as **Lewis acid-base chemistry**. The species donating the electron pair that compose the bond is a **Lewis base**, the species accepting the electron pair is a **Lewis acid**, and the product of the reaction is a **Lewis acid-base adduct**. As the two examples above illustrate, Brønsted-Lowry acid-base reactions represent a subcategory of Lewis acid reactions, specifically, those in which the acid species is H^+. A few examples involving other Lewis acids and bases are described below.

The boron atom in boron trifluoride, BF_3, has only six electrons in its valence shell. Being short of the preferred octet, BF_3 is a very good Lewis acid and reacts with many Lewis bases. Shown in Figure 9.6, a fluoride ion is the Lewis base in this reaction, donating one of its lone pairs:

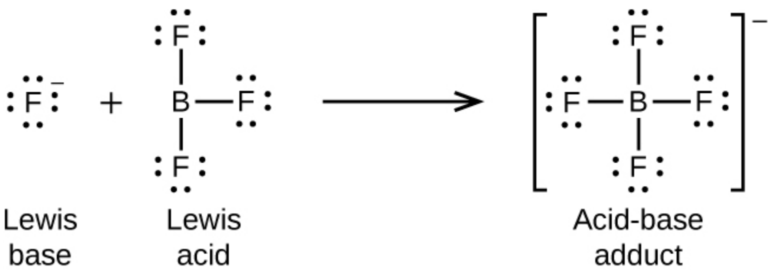


Figure 9.6 Fluoride ion donating one of its lone pairs. [credit: *Chemistry 2e.* Section 15.2. OpenStax. CC BY.]

In the following reaction, each of two ammonia molecules, Lewis bases, donates a pair of electrons to a silver ion, the Lewis acid shown in Figure 9.7:

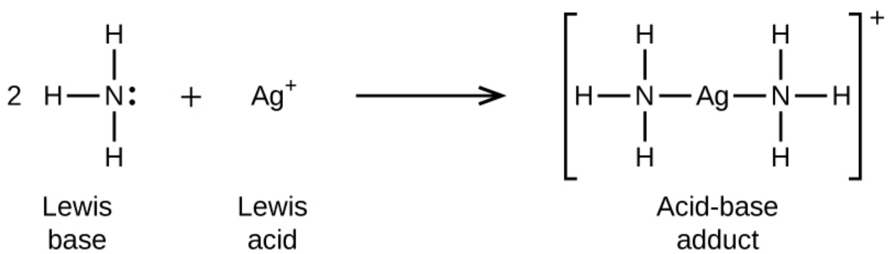


Figure 9.7 Ammonia molecules donating a pair of electrons to a silver ion. [credit: *Chemistry 2e.* Section 15.2. OpenStax. CC BY.]

Shown in Figure 9.8, nonmetal oxides act as Lewis acids and react with oxide ions, Lewis bases, to form oxyanions:

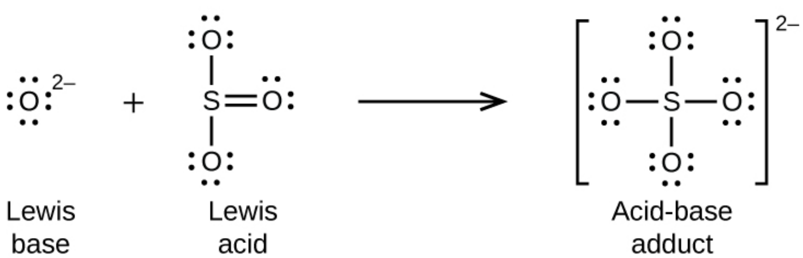


Figure 9.8 Equation illustration of the formation of oxyanions. [credit: *Chemistry 2e.* Section 15.2. OpenStax. CC BY.]

Many Lewis acid-base reactions are displacement reactions in which one Lewis base displaces another Lewis base from an acid-base adduct, or in which one Lewis acid displaces another Lewis acid. Figure 9.9 illustrates such displacement reactions.

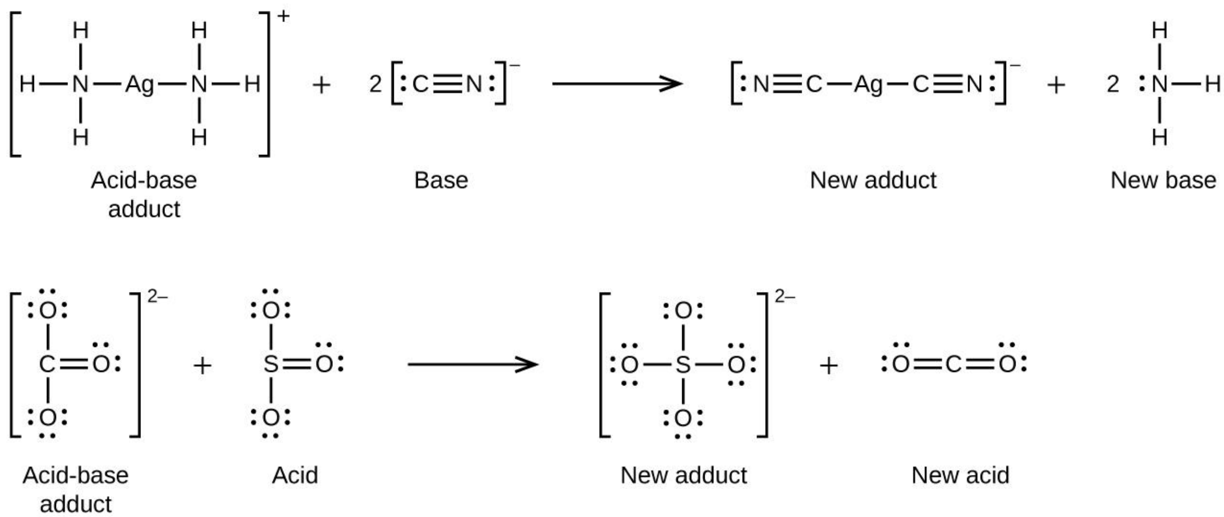


Figure 9.9 Equation illustration of Lewis acid-base displacement reactions. [credit: *Chemistry 2e.* Section 15.2. OpenStax. CC BY.]

Another type of Lewis acid-base chemistry involves the formation of a complex ion (or a coordination complex) comprising a central atom, typically a transition metal cation, surrounded by ions or molecules called **ligands**. These ligands can be neutral molecules like H_2O or NH_3, or ions such as CN^- or OH^-. Often, the ligands act as Lewis bases, donating a pair of electrons to the central atom.

Acids and Bases as Electrolytes

Acids and bases, like salts, dissociate in water into electrolytes. Acids and bases can very much change the properties of the solutions in which they are dissolved.

Acids

An **acid** is a substance that releases hydrogen ions (H^+) in solution (Figure 9.10a). Because an atom of hydrogen has just one proton and one electron, a positively charged hydrogen ion is simply a proton. This solitary proton is highly likely to participate in chemical reactions. Strong acids are compounds that release all of their H^+ in solution; that is, they ionize completely. Hydrochloric acid (HCl), which is released from cells in the lining of the stomach, is a strong acid because it releases all of its H^+ in the stomach's watery environment. This strong acid aids in digestion and kills ingested microbes. Weak acids do not ionize completely; that is, some of their hydrogen ions remain bonded within a compound in solution. An example of a weak acid is vinegar, or acetic acid; it is called acetate after it gives up a proton.

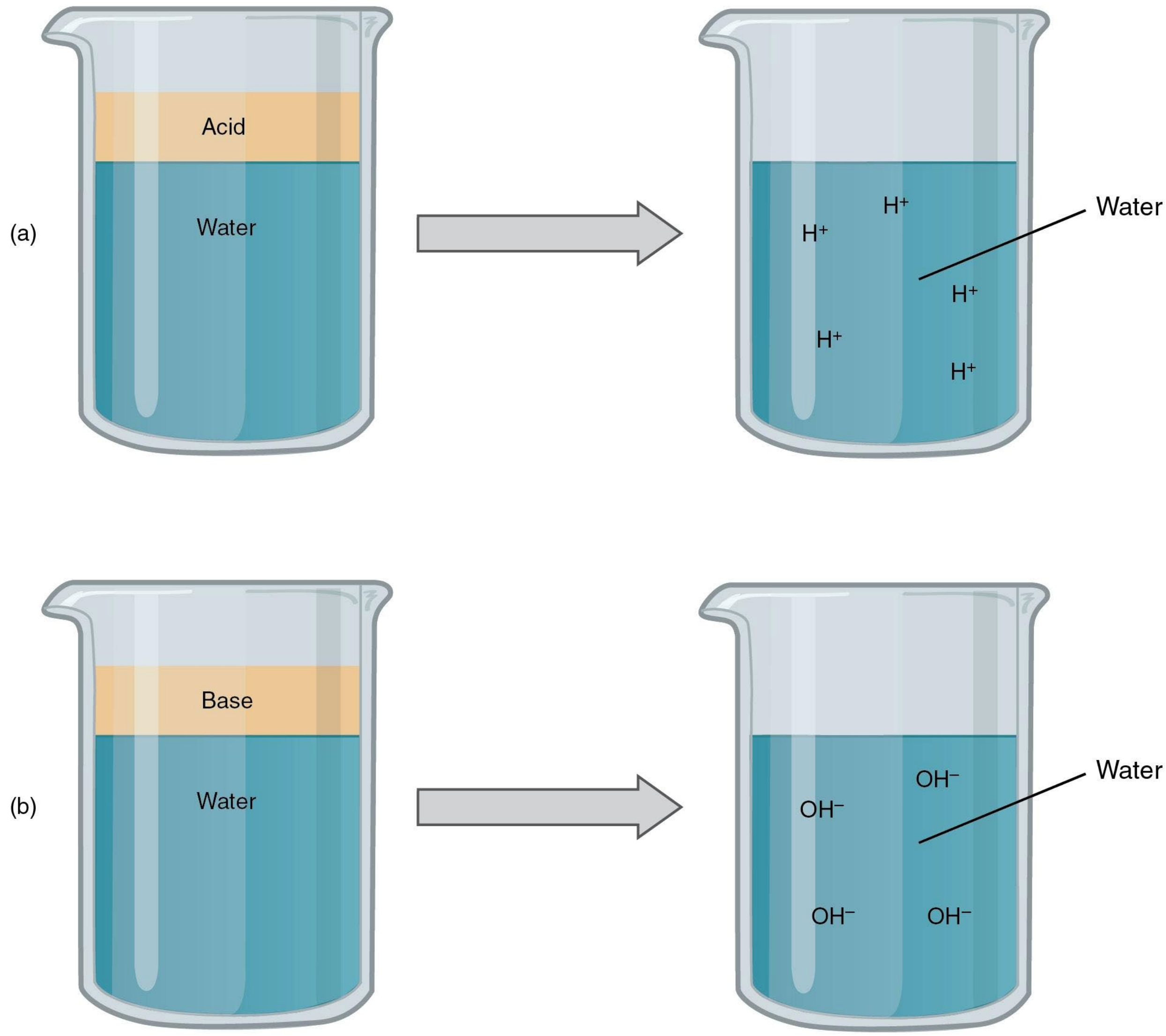


Figure 9.10 Acids and Bases (a) In aqueous solution, an acid dissociates into hydrogen ions (H^+) and anions. Nearly every molecule of a strong acid dissociates, producing a high concentration of H^+. (b) In aqueous solution, a base dissociates into hydroxyl ions (OH^-) and cations. Nearly every molecule of a strong base dissociates, producing a high concentration of OH^-. [credit: *Anatomy and Physiology.* Figure 2.16. OpenStax. CC BY.]

Bases

A **base** is a substance that releases hydroxyl ions (OH^-) in solution, or one that accepts H^+ already present in solution (Figure 9.10b). The hydroxyl ions (also known as hydroxide ions) or other basic substances combine with H^+ present to form a water molecule, thereby removing H^+ and reducing the solution's acidity. Strong bases release most or all of their hydroxyl ions; weak bases release only some hydroxyl ions or absorb only a few H^+. Food mixed with hydrochloric acid from the stomach would burn the small intestine, the next portion of the digestive tract after the stomach, if it were not for the release of bicarbonate (HCO_3^-), a weak base that attracts H^+. Bicarbonate accepts some of the H^+ protons, thereby reducing the acidity of the solution.

The Concept of pH

The relative acidity or alkalinity of a solution can be indicated by its pH. A solution's **pH** is the negative, base-10 logarithm of the hydrogen ion (H^+) concentration of the solution. As an example, a pH 4 solution has an H^+ concentration that is ten times greater than that of a pH 5 solution. That is, a solution with a pH of 4 is ten times more acidic than a solution with a pH of 5. The concept of pH will begin to make more sense when you study the pH scale. The scale consists of a series of increments ranging from 0 to 14. A solution with a pH of 7 is considered neutral—neither acidic nor basic. Pure water has a pH of 7. The lower the number below 7, the more acidic the solution, or the greater the concentration of H^+. The concentration of hydrogen ions at each pH value is 10 times different than the next pH. For instance, a pH value of 4 corresponds to a proton concentration of 10^{-4} M, or 0.0001M, while a pH value of 5 corresponds to a proton concentration of 10^{-5} M, or 0.00001M. The higher the number above 7, the more basic (alkaline) the solution, or the lower the concentration of H^+. Human urine, for example, is ten times more acidic than pure water, and HCl is 10,000,000 times more acidic than water.

pH Calculations

Hydronium and hydroxide ions are present both in pure water and in all aqueous solutions, and their concentrations are inversely proportional as determined by the ion product of water (K_w). The concentrations of these ions in a solution are often critical determinants of the solution's properties and the chemical behaviors of its other solutes, and specific vocabulary has been developed to describe these concentrations in relative terms. A solution is **neutral** if it contains equal concentrations of hydronium and hydroxide ions; **acidic** if it contains a greater concentration of hydronium ions than hydroxide ions; and **basic** if it contains a lesser concentration of hydronium ions than hydroxide ions.

A common means of expressing quantities that may span many orders of magnitude is to use a logarithmic scale. One such scale that is very popular for chemical concentrations and equilibrium constants is based on the p-function, defined as shown where "X" is the quantity of interest and "log" is the base-10 logarithm:

$$pX = -log\ X$$

The **pH** of a solution is therefore defined as shown here, where $[H_3O^+]$ is the molar concentration of hydronium ion in the solution,

$$\mathrm{pH} = -log[\mathrm{H_3O^+}]$$

Rearranging this equation to isolate the hydronium ion molarity yields the equivalent expression,

$$[\mathrm{H_3O^+}] = 10^{-\mathrm{pH}}$$

Likewise, the hydroxide ion molarity may be expressed as a p-function, or **pOH**:

$$\mathrm{pOH} = -log[\mathrm{OH^-}]$$

or

$$[\mathrm{OH^-}] = 10^{-\mathrm{pOH}}$$

Finally, the relation between these two-ion concentration expressed as p-functions is easily derived from the K_w expression:

$$K_\mathrm{w} = [\mathrm{H_3O^+}][\mathrm{OH^-}]$$

$$-logK_\mathrm{w} = -log([\mathrm{H_3O^+}][\mathrm{OH^-}]) = -log[\mathrm{H_3O^+}] + --\ log[\mathrm{OH^-}]$$

$$\mathrm{p}K_\mathrm{w} = \mathrm{pH} + \mathrm{pOH}$$

At 25 °C, the value of K_w is $1.0\ \times\ 10^{-14}$, and so,

$$14.00\ =\ \mathrm{pH} + \mathrm{pOH}$$

As will be shown later in Lab Example 9.5, the hydronium ion molarity in pure water (or any neutral solution) is 1.0×10^{-7} *M* at 25 °C. The pH and pOH of a neutral solution at this temperature are therefore:

$$\mathrm{pH} = -log[\mathrm{H_3O^+}] = -log(1.0 \times 10^{-7}) = 7.00$$

$$\mathrm{pOH} = -log[\mathrm{OH^-}] = -log(1.0 \times 10^{-7}) = 7.00$$

And so, *at this temperature*, acidic solutions are those with hydronium ion molarities greater than 1.0×10^{-7} *M* and hydroxide ion molarities less than 1.0×10^{-7} *M* (corresponding to pH values less than 7.00 and pOH values greater than 7.00). Basic solutions are those with hydronium ion molarities less than 1.0×10^{-7} M and hydroxide ion molarities greater than 1.0×10^{-7} M (corresponding to pH values greater than 7.00 and pOH values less than 7.00).

Since the autoionization constant K_w is temperature dependent, these correlations between pH values and the acidic/neutral/basic adjectives will be different at temperatures other than 25 °C. For example, the "Check Your Learning" exercise accompanying Lab Example 9.5 will later show the hydronium molarity of pure water at 80 °C is 4.9×10^{-7} M, which corresponds to pH and pOH values of:

$$\text{pH} = -log[\text{H}_3\text{O}^+] = -log(4.9 \times 10^{-7}) = 6.31$$

$$\text{pOH} = -log[\text{OH}^-] = -log(4.9 \times 10^{-7}) = 6.31$$

At this temperature, then, neutral solutions exhibit pH = pOH = 6.31, acidic solutions exhibit pH less than 6.31 and pOH greater than 6.31, whereas basic solutions exhibit pH greater than 6.31 and pOH less than 6.31. This distinction can be important when studying certain processes that occur at other temperatures, such as enzyme reactions in warm-blooded organisms at a temperature around 36–40 °C. Unless otherwise noted, references to pH values are presumed to be those at 25 °C (Table 9.1).

Table 9.1 Summary of Relations for Acidic, Basic, and Neutral Solutions

Classification	Relative Ion Concentrations	pH at 25 °C
acidic	$[H_3O^+] > [OH^-]$	$pH < 7$
neutral	$[H_3O^+] = [OH^-]$	$pH = 7$
basic	$[H_3O^+] < [OH^-]$	$pH > 7$

[credit: *Chemistry 2e.* Table 14.1. OpenStax. CC BY.]

Figure 9.11 shows the relationships between $[H_3O^+]$, $[OH^-]$, pH, and pOH for solutions classified as acidic, basic, and neutral.

$[H_3O^+]$ (M)	$[OH^-]$ (M)	pH	pOH	Sample Solution
10^{1}	10^{-15}	–1	15	
10^{0} or 1	10^{-14}	0	14	
10^{-1}	10^{-13}	1	13	
10^{-2}	10^{-12}	2	12	
10^{-3}	10^{-11}	3	11	
10^{-4}	10^{-10}	4	10	
10^{-5}	10^{-9}	5	9	
10^{-6}	10^{-8}	6	8	
10^{-7}	10^{-7}	7	7	
10^{-8}	10^{-6}	8	6	
10^{-9}	10^{-5}	9	5	
10^{-10}	10^{-4}	10	4	
10^{-11}	10^{-3}	11	3	
10^{-12}	10^{-2}	12	2	
10^{-13}	10^{-1}	13	1	
10^{-14}	10^{0} or 1	14	0	
10^{-15}	10^{1}	15	–1	

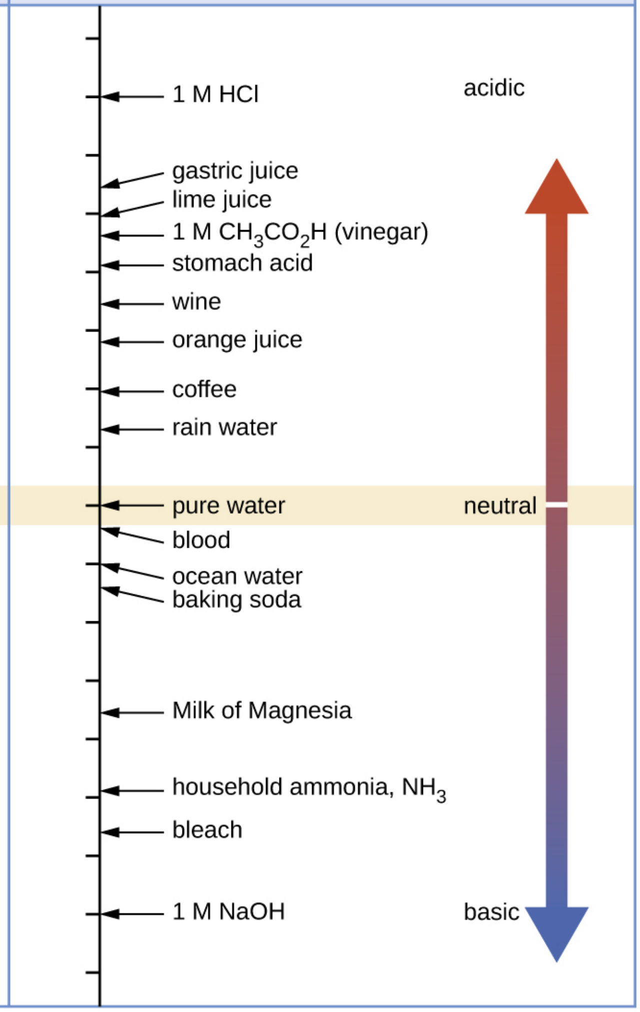


Figure 9.11 The pH and pOH scales represent concentrations of H_3O^+ and OH^-, respectively. The pH and pOH values of some common substances at 25 °C are shown in this chart. [credit: *Chemistry 2e.* Figure 14.2. OpenStax. CC BY.]

Acid and Base Ionization Constants

The relative strength of an acid or base is the extent to which it ionizes when dissolved in water. If the ionization reaction is essentially complete, the acid or base is termed *strong*; if relatively little ionization occurs, the acid or base is weak. As will be evident throughout the remainder of this chapter, there are many more weak acids and bases than strong ones. The most common strong acids and bases are listed in Figure 9.12.

6 Strong Acids		6 Strong Bases	
$HClO_4$	perchloric acid	LiOH	lithium hydroxide
HCl	hydrochloric acid	NaOH	sodium hydroxide
HBr	hydrobromic acid	KOH	potassium hydroxide
HI	hydroiodic acid	$Ca(OH)_2$	calcium hydroxide
HNO_3	nitric acid	$Sr(OH)_2$	strontium hydroxide
H_2SO_4	sulfuric acid	$Ba(OH)_2$	barium hydroxide

Figure 9.12 Some of the common strong acids and bases are listed here. [credit: *Chemistry 2e.* Figure 14.6. OpenStax. CC BY.]

The relative strengths of acids may be quantified by measuring their equilibrium constants in aqueous solutions. In solutions of the same concentration, stronger acids ionize to a greater extent, and so yield higher concentrations of hydronium ions than do weaker acids. The equilibrium constant for an acid is called the **acid-ionization constant, K_a**. For the reaction of an acid HA:

$$\mathrm{HA}(aq) + \mathrm{H_2O}(l) \rightleftharpoons \mathrm{H_3O^+}(aq) + \mathrm{A^-}(aq),$$

the acid ionization constant is written,

$$K_\mathrm{a} = \frac{[\mathrm{H_3O^+}][\mathrm{A^-}]}{[\mathrm{HA}]}$$

where the concentrations are those at equilibrium. Although water is a reactant in the reaction, it is the solvent as well, so we do not include $[H_2O]$ in the equation. The larger the K_a of an acid, the larger the concentration of H_3O^+ and A^- relative to the concentration of the nonionized acid, HA, in an equilibrium mixture, and the stronger the acid. An acid is classified as “strong” when it undergoes complete ionization, in which case the concentration of HA is zero and the acid ionization constant is immeasurably large ($K_a \approx \infty$). Acids that are partially ionized are called “weak,” and their acid ionization constants may be experimentally measured. A table of ionization constants for weak acids is provided in OpenStax *Chemistry 2e* Appendix H.

To illustrate this idea, three acid ionization equations and K_a values are shown below. The ionization constants increase from first to last of the listed equations, indicating the relative acid strength increases in the order $CH_3CO_2H < HNO_2 < HSO_4^-$:

$$CH_3CO_2H(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CH_3CO_2^-(aq), \qquad K_a = 1.8 \times 10^{-5}$$

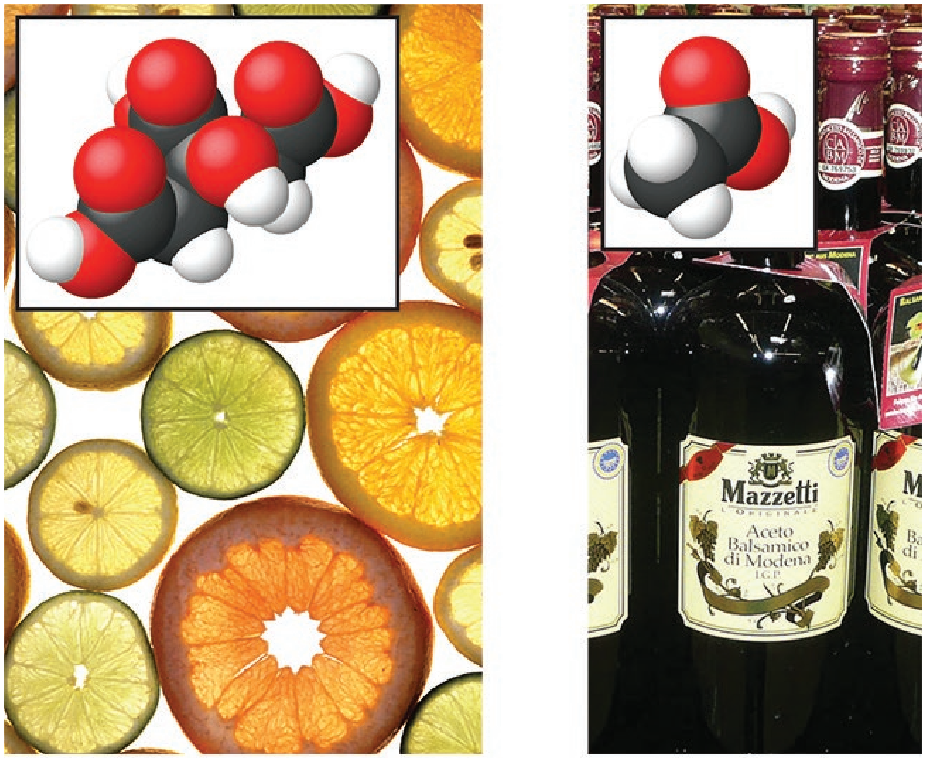
$$HNO_2(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + NO_2^-(aq), \qquad K_a = 4.6 \times 10^{-4}$$

$$HSO_4^-(aq) + H_2O(aq) \rightleftharpoons H_3O^+(aq) + SO_4^{2-}(aq), \qquad K_a = 1.2 \times 10^{-2}$$

The nature of HCl is such that its reaction with water as just described is essentially 100% efficient: Virtually every HCl molecule that dissolves in water will undergo this reaction. Acids that completely react in this fashion are called **strong acids**, and HCl is one among just a handful of common acid compounds that are classified as strong (Table 9.2). A far greater number of compounds behave as **weak acids** and only partially react with water, leaving a large majority of dissolved molecules in their original form and generating a relatively small amount of hydronium ions. Weak acids are commonly encountered in nature, being the substances partly responsible for the tangy taste of citrus fruits, the stinging sensation of insect bites, and the unpleasant smells associated with body odor. A familiar example of a weak acid is acetic acid, the main ingredient in food vinegars:

$$CH_3CO_2H(aq) + H_2O(l) \rightleftharpoons CH_3CO_2^-(aq) + H_3O^+(aq)$$

When dissolved in water under typical conditions, only about 1% of acetic acid molecules are present in the ionized form, $CH_3CO_2^-$ (Figure 9.13). (The use of a double-arrow in the equation above denotes the partial reaction aspect of this process.)



(a) (b)

Figure 9.13 (a) Fruits such as oranges, lemons, and grapefruit contain the weak acid citric acid. (b) Vinegars contain the weak acid acetic acid. [As cited in *Chemistry 2e.* Figure 4.6. OpenStax. CC BY, (a) modification of work by Scott Bauer, (b) modification of work by Brücke-Osteuropa/Wikimedia Commons.]

Table 9.2 Common Strong Acids

Compound Formula	**Name in Aqueous Solution**
HBr	hydrobromic acid
HCl	hydrochloric acid
HI	hydroiodic acid
HNO_3	nitric acid
$HClO_4$	perchloric acid
H_2SO_4	sulfuric acid

[credit: *Chemistry 2e.* Table 4.2. OpenStax. CC BY.]

A **base** is a substance that will dissolve in water to yield hydroxide ions, OH^-. The most common bases are ionic compounds composed of alkali or alkaline earth metal cations (groups 1 and 2) combined with the hydroxide ion—for example, $NaOH$ and $Ca(OH)_2$. Unlike the acid compounds discussed previously, these compounds do not react chemically with water; instead they dissolve and dissociate, releasing hydroxide ions directly into the solution. For example, KOH and $Ba(OH)_2$ dissolve in water and dissociate completely to produce cations (K^+ and Ba^{2+}, respectively) and hydroxide ions, OH^-. These bases, along with other hydroxides that completely dissociate in water, are considered **strong bases**.

Consider as an example the dissolution of lye (sodium hydroxide) in water,

$$\mathrm{NaOH}(s) \longrightarrow \mathrm{Na}^+(aq) + \mathrm{OH}^-(aq)$$

This equation confirms that sodium hydroxide is a base. When dissolved in water, $NaOH$ dissociates to yield Na^+ and OH^- ions. This is also true for any other ionic compound containing hydroxide ions. Since the dissociation process is essentially complete when ionic compounds dissolve in water under typical conditions, $NaOH$ and other ionic hydroxides are all classified as strong bases.

Unlike ionic hydroxides, some compounds produce hydroxide ions when dissolved by chemically reacting with water molecules. In all cases, these compounds react only partially and so are classified as **weak bases**. These types of compounds are also abundant in nature and important commodities in various technologies. For example, global production of the weak base ammonia is typically well over 100 metric tons annually, being widely used as an agricultural fertilizer, a raw material for chemical synthesis of other compounds, and an active ingredient in household cleaners (Figure 9.14). When dissolved in water, ammonia reacts partially to yield hydroxide ions, as shown here,

$$\mathrm{NH_3}(aq) + \mathrm{H_2O}(l) \rightleftharpoons \mathrm{NH_4}^+(aq) + \mathrm{OH}^-(aq)$$

This is, by definition, an acid-base reaction, in this case involving the transfer of H^+ ions from water molecules to ammonia molecules. Under typical conditions, only about 1% of the dissolved ammonia is present as NH_4^+ ions.

(a) (b)

Figure 9.14 Ammonia is a weak base used in a variety of applications. (a) Pure ammonia is commonly applied as an agricultural fertilizer. (b) Dilute solutions of ammonia are effective household cleansers. [As cited in *Chemistry 2e.* Figure 4.7. OpenStax. CC BY, (a) modification of work by National Resources Conservation Service, (b) modification of work by pat00139.]

Acid-Base Reactions

An **acid-base reaction** is one in which a hydrogen ion, H^+, is transferred from one chemical species to another. Such reactions are of central importance to numerous natural and technological processes, ranging from the chemical transformations that take place within cells and the lakes and oceans, to the industrial-scale production of fertilizers, pharmaceuticals, and other substances essential to society.

For purposes of this brief introduction, we will consider only the more common types of acid-base reactions that take place in aqueous solutions. In this context, an **acid** is a substance that will dissolve in water to yield hydronium ions, H_3O^+. As an example, consider the equation shown here:

$$\mathrm{HCl}(aq) + \mathrm{H_2O}(aq) \longrightarrow \mathrm{Cl}^-(aq) + \mathrm{H_3O}^+(aq)$$

The process represented by this equation confirms that hydrogen chloride is an acid. When dissolved in water, H_3O^+ ions are produced by a chemical reaction in which H^+ ions are transferred from HCl molecules to H_2O molecules (see Figure 9.15).

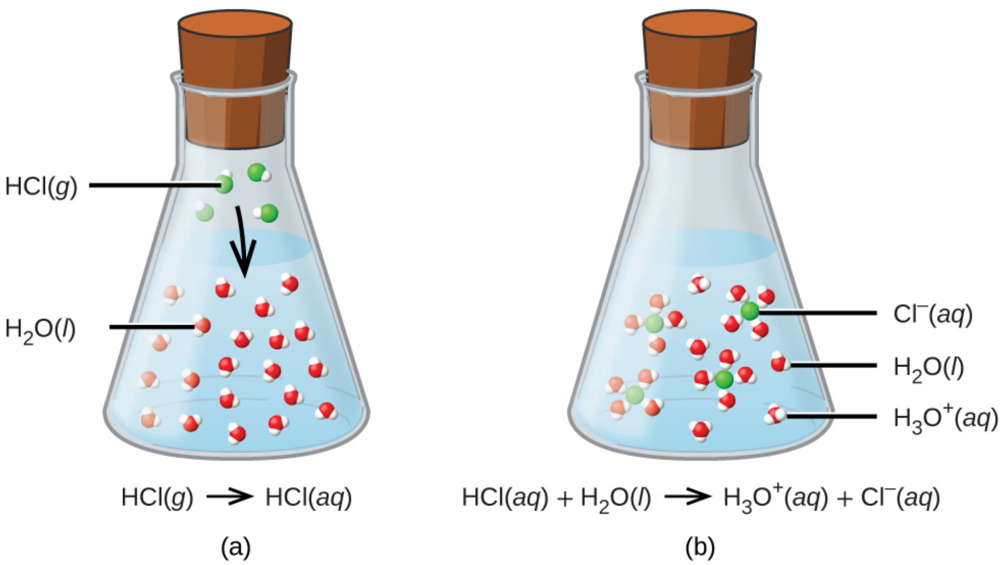


Figure 9.15 When hydrogen chloride gas dissolves in water, (a) it reacts as an acid, transferring protons to water molecules to yield (b) hydronium ions (and solvated chloride ions). [credit: *Chemistry 2e.* Figure 4.5. OpenStax. CC BY.]

A **neutralization reaction** is a specific type of acid-base reaction in which the reactants are an acid and a base (but not water), and the products are often a **salt** and water,

$$\text{acid} + \text{base} \longrightarrow \text{salt} + \text{water}$$

To illustrate a neutralization reaction, consider what happens when a typical antacid such as milk of magnesia (an aqueous suspension of solid $\mathrm{Mg(OH)_2}$ is ingested to ease symptoms associated with excess stomach acid (HCl):

$$\mathrm{Mg(OH)_2}(s) + 2\mathrm{HCl}(aq) \longrightarrow \mathrm{MgCl_2}(aq) + 2\mathrm{H_2O}(l)$$

Note that in addition to water, this reaction produces a salt, magnesium chloride.

Reaction of Metals with Acids

This is the most convenient laboratory method of producing hydrogen. Metals with lower reduction potentials reduce the hydrogen ion in dilute acids to produce hydrogen gas and metal salts. For example, as shown in Figure 9.16, iron in dilute hydrochloric acid produces hydrogen gas and iron(II) chloride,

$$\mathrm{Fe}(s) + 2\mathrm{H_3O^+}(aq) + 2\mathrm{Cl^-}(aq) \longrightarrow \mathrm{Fe^{2+}}(aq) + 2\mathrm{Cl^-}(aq) + \mathrm{H_2}(g) + 2\mathrm{H_2O}(l)$$

Figure 9.16 The reaction of iron with an acid produces hydrogen. Here, iron reacts with hydrochloric acid. [credit: Mark Ott, as cited in *Chemistry 2e.* Figure 18.27. OpenStax. CC BY.]

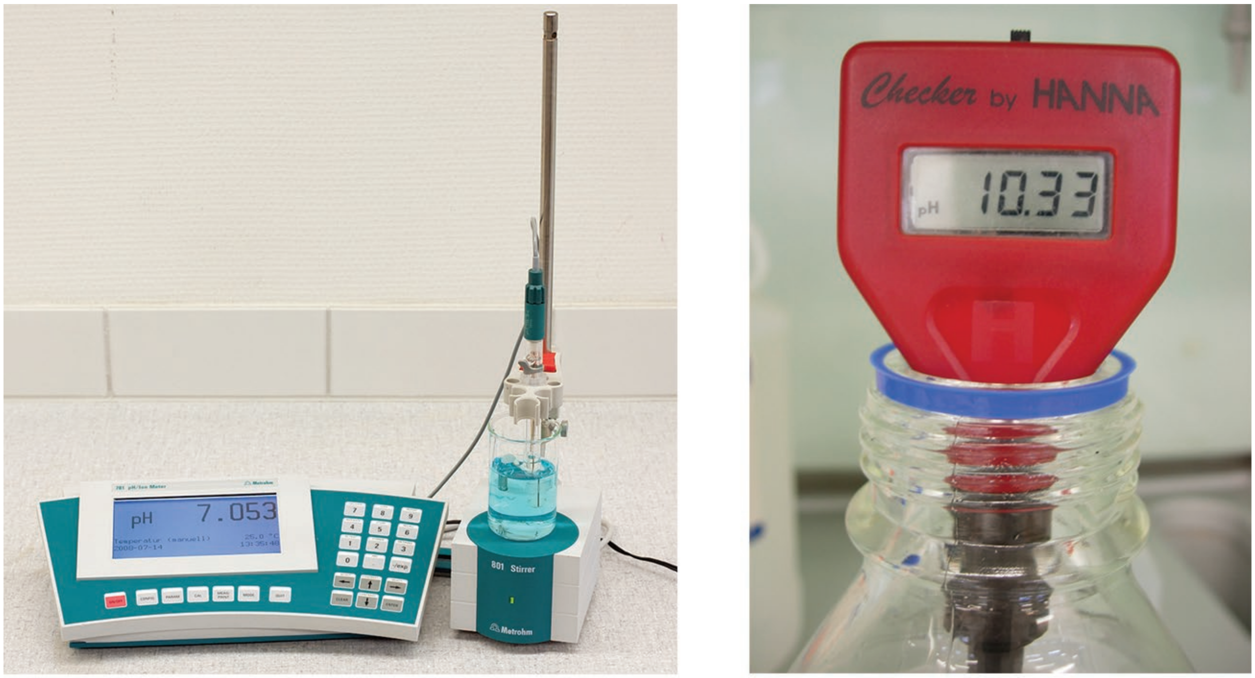
Buffers

The pH of human blood normally ranges from 7.35 to 7.45, although it is typically identified as pH 7.4. At this slightly basic pH, blood can reduce the acidity resulting from the carbon dioxide (CO_2) constantly being released into the bloodstream by the trillions of cells in the body. Homeostatic mechanisms (along with exhaling CO_2 while breathing) normally keep the pH of blood within this narrow range. This is critical, because fluctuations—either too acidic or too alkaline—can lead to life-threatening disorders.

All cells of the body depend on homeostatic regulation of acid–base balance at a pH of approximately 7.4. The body therefore has several mechanisms for this regulation, involving breathing, the excretion of chemicals in urine, and the internal release of chemicals collectively called buffers into body fluids. A **buffer** is a solution of a weak acid and its conjugate base. A buffer can neutralize small amounts of acids or bases in body fluids. For example, if there is even a slight decrease below 7.35 in the pH of a bodily fluid, the buffer in the fluid—in this case, acting as a weak base—will bind the excess hydrogen ions. In contrast, if pH rises above 7.45, the buffer will act as a weak acid and contribute hydrogen ions.

How to Measure pH of Solutions

The acidity of a solution is typically assessed experimentally by measurement of its pH. The pOH of a solution is not usually measured, as it is easily calculated from an experimentally determined pH value. The pH of a solution can be directly measured using a pH meter (Figure 9.17).



(a) (b)

Figure 9.17 (a) A research-grade pH meter used in a laboratory can have a resolution of 0.001 pH units, an accuracy of ± 0.002 pH units, and may cost in excess of $1000. (b) A portable pH meter has lower resolution (0.01 pH units), lower accuracy (± 0.2 pH units), and a far lower price tag. [credit: As cited in *Chemistry 2e.* Figure 14.4. OpenStax. CC BY, b) modification of work by Jacopo Werther].

The pH of a solution may also be visually estimated using colored indicators, as shown in Figure 9.18.

(a) (b)

Figure 9.18 (a) A solution containing a dye mixture, called universal indicator, takes on different colors depending upon its pH. (b) Convenient test strips, called pH paper, contain embedded indicator dyes that yield pH-dependent color changes on contact with aqueous solutions. [credit: modification of work by Sahar Atwa in *Chemistry 2e.* Figure 14.5. OpenStax. CC BY.]

Lab Examples

EXAMPLE 9.1: CALCULATION OF PH FROM H_3O

What is the pH of stomach acid, a solution of HCl with a hydronium ion concentration of 1.2×10^{-3} M?

Solution

$$\text{pH} = -log[\text{H}_3\text{O}^+]$$

$$= -log(1.2 \times 10^{-3})$$

$$= -(-2.92)$$

$$= 2.92$$

(The use of logarithms is explained in OpenStax *Chemistry 2e* Appendix B Essential Mathematics. When taking the log of a value, keep as many decimal places in the result as there are significant figures in the value.)

Check Your Learning
Water exposed to air contains carbonic acid, H_2CO_3, due to the reaction between carbon dioxide and water,

$$\text{CO}_2(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_2\text{CO}_3(aq)$$

Air-saturated water has a hydronium ion concentration caused by the dissolved CO_2 of 2.0×10^{-6} *M*, about 20-times larger than that of pure water. Calculate the pH of the solution at 25 °C.

ANSWER
5.70

EXAMPLE 9.2: CALCULATION OF HYDRONIUM ION CONCENTRATION FROM PH

Calculate the hydronium ion concentration of blood, the pH of which is 7.3.

Solution

$$\mathrm{pH} = -log[\mathrm{H_3O^+}] = 7.3$$

$$log[\mathrm{H_3O^+}] = -7.3$$

$$[\mathrm{H_3O^+}] = 10^{-7.3} \text{ or } [\mathrm{H_3O^+}] = \text{antilog of } -7.3$$

$$[\mathrm{H_3O^+}] = 5 \times 10^{-8} M$$

(On a calculator take the antilog, or the “inverse” log, of −7.3, or calculate $10^{-7.3}$.)

Check Your Learning

Calculate the hydronium ion concentration of a solution with a pH of −1.07.

ANSWER

12 *M*

EXAMPLE 9.3: ENVIRONMENTAL SCIENCE CONNECTION TO CHEMISTRY

Normal rainwater has a pH between 5 and 6 due to the presence of dissolved CO_2 which forms carbonic acid:

$$\mathrm{H_2O}(l) + \mathrm{CO_2}(g) \longrightarrow \mathrm{H_2CO_3}(aq)$$

$$\mathrm{H_2CO_3}(aq) \rightleftharpoons \mathrm{H^+}(aq) + \mathrm{HCO_3^-}(aq)$$

Acid rain is rainwater that has a pH of less than 5, due to a variety of nonmetal oxides, including CO_2, SO_2, SO_3, NO, and NO_2 being dissolved in the water and reacting with it to form not only carbonic acid, but sulfuric acid and nitric acid. The formation and subsequent ionization of sulfuric acid are shown here:

$$\mathrm{H_2O}(l) + \mathrm{SO_3}(g) \longrightarrow \mathrm{H_2SO_4}(aq)$$

$$\mathrm{H_2SO_4}(aq) \longrightarrow \mathrm{H^+}(aq) + \mathrm{HSO_4^-}(aq)$$

Carbon dioxide is naturally present in the atmosphere because most organisms produce it as a waste product of metabolism. Carbon dioxide is also formed when fires release carbon stored in vegetation or fossil fuels. Sulfur trioxide in the atmosphere is naturally produced by volcanic activity, but it also originates from burning fossil fuels, which have traces of sulfur, and from the process of "roasting" ores of metal sulfides in metal-refining processes. Oxides of nitrogen are formed in internal combustion engines where the high temperatures make it possible for the nitrogen and oxygen in air to chemically combine.

Acid rain is a particular problem in industrial areas where the products of combustion and smelting are released into the air without being stripped of sulfur and nitrogen oxides. In North America and Europe until the 1980s, it was responsible for the destruction of forests and freshwater lakes, when the acidity of the rain actually killed trees, damaged soil, and made lakes uninhabitable for all but the most acid-tolerant species. Acid rain also corrodes statuary and building facades that are made of marble and limestone (Figure 9.19). Regulations limiting the amount of sulfur and nitrogen oxides that can be released into the atmosphere by industry and automobiles have reduced the severity of acid damage to both natural and manmade environments in North America and Europe. It is now a growing problem in industrial areas of China and India. For further information on acid rain, visit this Acid Rain website hosted by the US Environmental Protection Agency.

(a)

(b)

Figure 9.19 (a) Acid rain makes trees more susceptible to drought and insect infestation and depletes nutrients in the soil. (b) It also corrodes statues that are carved from marble or limestone. [credit: As cited in *Chemistry 2e.* Figure 14.3. OpenStax. CC BY, (a) modification of work by Chris M Morris, (b) modification of work by "Eden, Janine and Jim"/Flickr.]

Example 9.4: Calculation of pOH

What are the pOH and the pH of a 0.0125-*M* solution of potassium hydroxide, KOH?

Solution

Potassium hydroxide is a highly soluble ionic compound and completely dissociates when dissolved in dilute solution, yielding $[OH-]$ = 0.0125 *M*,

$$\mathrm{pOH} = -log[\mathrm{OH}^-]$$

$$= -log\ 0.0125$$

$$= -(-1.903)$$

$$= 1.903$$

The pH can be found from the pOH:

$$\mathrm{pH} + \mathrm{pOH} = 14.00$$

$$\mathrm{pH} = 14.00 - \mathrm{pOH}$$

$$= 14.00 - 1.903$$

$$= 12.10$$

Check Your Learning

The hydronium ion concentration of vinegar is approximately 4×10^{-3} *M*. What are the corresponding values of pOH and pH?

ANSWER:

pOH = 11.6, pH = 2.4

Example 9.5: Ion Concentrations in Pure Water

What are the hydronium ion concentration and the hydroxide ion concentration in pure water at 25 °C?

Solution

The autoionization of water yields the same number of hydronium and hydroxide ions. Therefore, in pure water, $[H_3O^+] = [OH^-] = x$.

At 25 °C,

$$K_\mathrm{w} = [\mathrm{H_3O^+}][\mathrm{OH^-}] = (x)(x) = x^2 = 1.0 \times 10^{-14}$$

So,

$$x = [\mathrm{H_3O^+}] = [\mathrm{OH^-}] = \sqrt{1.0 \times 10^{-14}} = 1.0 \times 10^{-7}\ M$$

The hydronium ion concentration and the hydroxide ion concentration are the same, $1.0 \times 10^{-1}\ M.$

Check Your Learning
The ion product of water at 80 °C is 2.2×10^{-13}. What are the concentrations of hydronium and hydroxide ions in pure water at 80 °C?

ANSWER
$[H_3O^+] = [OH^-] = 4.9 \times 10^{-7}\ M$

Example 9.6: The Inverse Relation between [H_3O^+] and [OH^-]

A solution of an acid in water has a hydronium ion concentration of $2.0 \times 10^{-6} M$. What is the concentration of hydroxide ion at 25 °C?

Solution
Use the value of the ion-product constant for water at 25 °C to calculate the missing equilibrium concentration,

$$2\mathrm{H_2O}(l) \rightleftharpoons \mathrm{H_3O^+}(aq) + \mathrm{OH^-}(aq)$$

$$K_\mathrm{w} = [\mathrm{H_3O^+}][\mathrm{OH^-}] = 1.0 \times 10^{-14}$$

Rearrangement of the Kw expression shows that $[\mathrm{OH^-}]$ is inversely proportional to $[\mathrm{H_3O^+}]$:

$$[\mathrm{OH^-}] = \frac{K_\mathrm{w}}{[\mathrm{H_3O^+}]} = \frac{1.0 \times 10^{-14}}{2.0 \times 10^{-6}} = 5.0 \times 10^{-9}$$

Compared with pure water, a solution of acid exhibits a higher concentration of hydronium ions (due to ionization of the acid) and a proportionally lower concentration of hydroxide ions. This may be explained via Le Châtelier's principle as a left shift in the water autoionization equilibrium resulting from the stress of increased hydronium ion concentration.

Substituting the ion concentrations into the K_w expression confirms this calculation, resulting in the expected value:

$$K_w = [H_3O^+][OH^-]$$

$$= (2.0 \times 10^{-6})(5.0 \times 10^{-9})$$

$$= 1.0 \times 10^{-14}$$

Check Your Learning
What is the hydronium ion concentration in an aqueous solution with a hydroxide ion concentration of 0.001 *M* at 25 °C?

ANSWER
$[H_3O^+] = 1 \times 10^{-11} M$

Example 9.7: Representing the Acid-Base Behavior of an Amphoteric Substance

Write separate equations representing the reaction of HSO_3^-
a) as an acid with OH^-
b) as a base with HI

Solution
a) $HSO_3{}^-(aq) + OH^-(aq) \rightleftharpoons SO_3{}^{2-}(aq) + H_2O(l)$

b) $HSO_3{}^-(aq) + HI(aq) \rightleftharpoons H_2SO_3(aq) + I^-(aq)$

Check Your Learning
Write separate equations representing the reaction of $H_2PO_4{}^-$
a) as a base with HBr
b) as an acid with OH^-

ANSWER:
a) $H_2PO_4{}^-(aq) + HBr(aq) \rightleftharpoons H_3PO_4(aq) + Br^-(aq)$

b) $H_2PO_4{}^-(aq) + OH^-(aq) \rightleftharpoons HPO_4{}^{2-}(aq) + H_2O(l)$

EXAMPLE 9.8: WRITING EQUATIONS FOR ACID-BASE REACTIONS

Write balanced chemical equations for the acid-base reactions described here:
(a) the weak acid hydrogen hypochlorite reacts with water
(b) a solution of barium hydroxide is neutralized with a solution of nitric acid

Solution

(a) The two reactants are provided, $HOCl$ and H_2O. Since the substance is reported to be an acid, its reaction with water will involve the transfer of H^+ from $HOCl$ to H_2O to generate hydronium ions, H_3O^+ and hypochlorite ions, OCl^-.

$$\mathrm{HOCl}(aq) + \mathrm{H_2O}(l) \rightleftharpoons \mathrm{OCl^-}(aq) + \mathrm{H_3O^+}(aq)$$

A double arrow is appropriate in this equation because it indicates the $HOCl$ is a weak acid that has not reacted completely.

(b) The two reactants are provided, $Ba(OH)_2$ and HNO_3. Since this is a neutralization reaction, the two products will be water and a salt composed of the cation of the ionic hydroxide (Ba^{2+}) and the anion generated when the acid transfers its hydrogen ion (NO_3^-).

$$\mathrm{Ba(OH)_2}(aq) + 2\mathrm{HNO_3}(aq) \longrightarrow \mathrm{Ba(NO_3)_2}(aq) + 2\mathrm{H_2O}(l)$$

Check Your Learning

Write the net ionic equation representing the neutralization of any strong acid with an ionic hydroxide. (Hint: Consider the ions produced when a strong acid is dissolved in water.)

ANSWER

$\mathrm{H_3O^+}(aq) + \mathrm{OH^-}(aq) \longrightarrow 2\mathrm{H_2O}(l)$

EXAMPLE 9.9: CHEMISTRY IN EVERYDAY LIFE AND STOMACH ANTACIDS

Our stomachs contain a solution of roughly 0.03 *M* HCl, which helps us digest the food we eat. The burning sensation associated with heartburn is a result of the acid of the stomach leaking through the muscular valve at the top of the stomach into the lower reaches of the esophagus. The lining of the esophagus is not protected from the corrosive effects of stomach acid the way the lining of the stomach is, and the results can be very painful. When we have heartburn, it feels better if we reduce the excess acid in the esophagus by taking an antacid. As you may have guessed, antacids are bases. One of the most common antacids is calcium carbonate, $CaCO_3$. The reaction,

$$CaCO_3(s) + 2HCl(aq) \rightleftharpoons CaCl_2(aq) + H_2O(l) + CO_2(g)$$

not only neutralizes stomach acid, it also produces $CO_2(g)$, which may result in a satisfying belch.

Milk of Magnesia is a suspension of the sparingly soluble base magnesium hydroxide, $Mg(OH)_2$. It works according to the reaction,

$$Mg(OH)_2(s) \rightleftharpoons Mg^{2+}(aq) + 2OH^-(aq)$$

The hydroxide ions generated in this equilibrium then go on to react with the hydronium ions from the stomach acid, so that,

$$H_3O^+ + OH^- \rightleftharpoons 2H_2O(l)$$

This reaction does not produce carbon dioxide, but magnesium-containing antacids can have a laxative effect. Several antacids have aluminum hydroxide, $Al(OH)_3$, as an active ingredient. The aluminum hydroxide tends to cause constipation, and some antacids use aluminum hydroxide in concert with magnesium hydroxide to balance the side effects of the two substances.

EXAMPLE 9.10: CULINARY ASPECTS OF CHEMISTRY IN EVERYDAY LIFE

Examples of acid-base chemistry are abundant in the culinary world. One example is the use of baking soda, or sodium bicarbonate in baking. $NaHCO_3$ is a base. When it reacts with an acid such as lemon juice, buttermilk, or sour cream in a batter, bubbles of carbon dioxide gas are formed from decomposition of the resulting carbonic acid, and the batter “rises.” Baking powder is a combination of sodium bicarbonate, and one or more acid salts that react when the two chemicals come in contact with water in the batter.

Many people like to put lemon juice or vinegar, both of which are acids, on cooked fish (Figure 9.20). It turns out that fish have volatile amines (bases) in their systems, which are neutralized by the acids to yield involatile ammonium salts. This reduces the odor of the fish, and also adds a “sour” taste that we seem to enjoy.

$$CH_3COOH + NH_2CH_2CH_2CH_2CH_2NH_2 \longrightarrow CH_3COO^- + NH_3^+CH_2CH_2CH_2CH_2NH_2$$

Acetic acid + Putrescine ⟶ Acetate ion + Putrescinium ion

Figure 9.20 A neutralization reaction takes place between citric acid in lemons or acetic acid in vinegar, and the bases in the flesh of fish. [credit: *Chemistry 2e.* Figure 4.8. OpenStax. CC BY.]

Pickling is a method used to preserve vegetables using a naturally produced acidic environment. The vegetable, such as a cucumber, is placed in a sealed jar submerged in a brine solution. The brine solution favors the growth of beneficial bacteria and suppresses the growth of harmful bacteria. The beneficial bacteria feed on starches in the cucumber and produce lactic acid as a waste product in a process called fermentation. The lactic acid eventually increases the acidity of the brine to a level that kills any harmful bacteria, which require a basic environment. Without the harmful bacteria consuming the cucumbers they are able to last much longer than if they were unprotected. A byproduct of the pickling process changes the flavor of the vegetables with the acid making them taste sour.

Example 9.11: Simulation of the Microscopic World

Play in the Acid-Base Solutions Simulation to explore the microscopic world of strong and weak acids and bases.

EXAMPLE 9.12: SUGGESTED SIMULATION ON THE PH SCALE

Please use the following PhET pH-Scale Simulation and discover if the given solutions such as coffee, blood, water, orange juice, and soda are acidic or basic. Since H_3O^+ and OH^- ion ratios change according to the pH value of the solutions, observe the pH changes when the given solutions are diluted. You can also determine concentration of hydroxide and hydronium ions at the given pH. After testing solutions, you can place acids and bases in a relative order.

EXAMPLE 9.13 SUGGESTED SIMULATION FOR ACID-BASE SOLUTIONS

By exploring the PhET Acid-Base Solutions Simulation, you will be able to recognize important properties of acid and base solutions, and will be able to describe differences between a strong acid/base and weak acid/base, acid strength, and able to understand available ions present in acidic and basic solutions. You will learn tools are used to help determine if a solution is an acid or base.

Relations to Health Sciences

Acids and bases can be found in the human body. One of the sources of acid and bases in the human body is amino acids, which basically are proteins and common nutrients. Amino acids are needed to maintain our body's health and are defined as the building blocks of protein. Our body needs 20 different acids to stay healthy and function properly. For example, amino acids stimulate muscle growth to build collagen or produce hormones. Nine of these 20 amino acids are classified as essential. We can only get essential amino acids through our diet. The body can make 11 of the nonessential amino acids.

Fatty acids are another class of acids and are a major component of our cells and they are important for brain health and metabolism. Ascorbic acid, known as vitamin C, cannot be made by our body and has an important function. Our body uses ascorbic acid to make blood vessels, tendons, and skin. In addition, this ascorbic acid class fights against viruses and keeps our immune system strong.

Hydrochloric acid is an important acid and is formed naturally in our stomach to help digestion. Without enough stomach acid, our stomach cannot break down the food we eat, and the body cannot absorb the nutrients. The human body is built in a way that can maintain healthy acidity and alkalinity balance. This balance is the level of acids and bases in our body at which body works best. There are three mechanisms used to regulate this balance: buffers, our respiratory system, and our urinary system.

Acids and Bases in Industry

Acids and bases are used in manufacturing industries, such as cosmetics, medicines, plastics, fertilizers, and foods. For example, the steel industry uses hydrochloric acid to clean metals before processing. In other industries, nitric acid is used to create fertilizers and dyes. Boric acid pools are used to store nuclear waste produced in power plants. Phosphoric acid has many uses in different industries, such as medicine, chemicals, construction, and food. The food industry also uses citric acid as a preservative.

On the other hand, bases are used to manufacture soap products, refine petroleum, and make medicines. Bases are also used to neutralize acids in water supplies, as an antidote for food poisoning, and by farmers to neutralize harmful effects of acid rains.

Foods and Drinks

Foods we consume have a wide spectrum of pH-values. Bases are less common as food, but they are present in many household products. Many foods we think are acidic, such as limes and lemons, are actually alkaline when ingested. Once metabolized, limes and lemons become alkaline with a pH well above 7. Most viruses are pH sensitive and need a slightly acidic environment so they can penetrate to the cell. Treating viral infections by raising pH of the body increases the immune system’s function to kill the bacteria. Therefore, it is important to know which foods are alkaline after they are metabolized. Figure 9.21 shows pH values of the food and common household products.

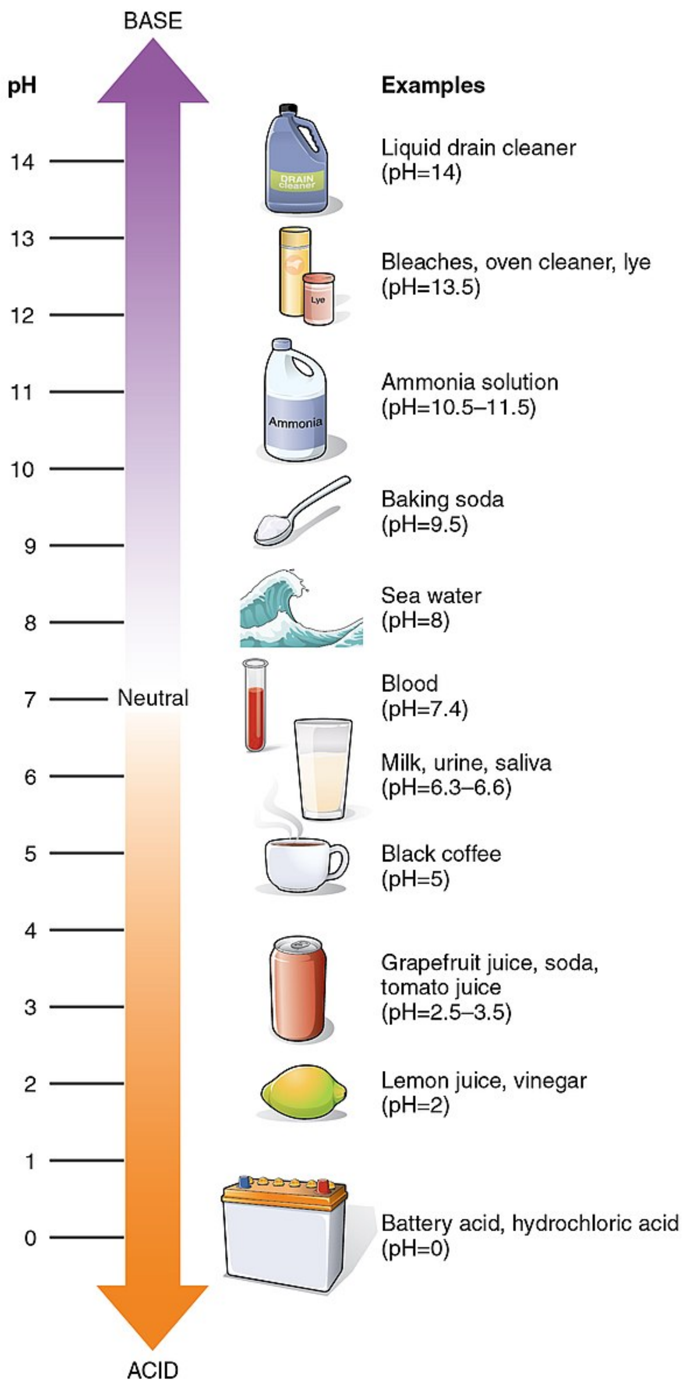


Figure 9.21 Range of pH values for commonly used household products. [credit: *Anatomy and Physiology*. Figure 2.17. OpenStax. CC BY.]

Mouth Chemistry

The pH range of healthy saliva is slightly acidic, ranging between 6.0-7.8 and depending on the person. The food and drink we consume change the pH level of saliva. If saliva is too acidic, it creates an environment that increases tooth demineralization, leading to tooth decay and cavities. Figure 9.22 shows demineralization starts when the pH is lower than 5.5.

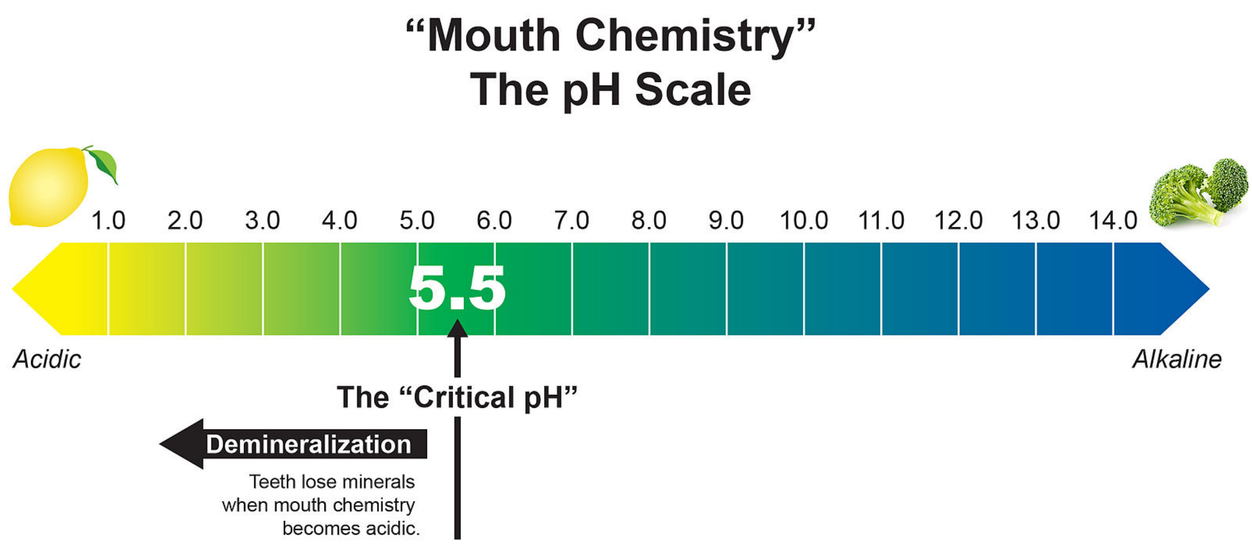


Figure 9.22 When the mouth pH is lower than 5.5, tooth demineralization occurs. [credit: Drmidei. (2013). De-mineralization that Occurs in Acidic Mouth. CC BY-SA.]

Rinsing out our mouth with water restores the pH level to neutral after eating carbohydrates rich, sugary foods, or consuming coffee and soda beverages. Another way to neutralize saliva is eating alkaline rich foods.

Fruits and Vegetables as Indicators

There are so many fruits and vegetables that can be used as pH indicators. They contain substances called anthocyanins. They are soluble pigments and can change color in different pH. As an example, Blue Butterfly Pea Flower Tea changes color when acidified with lemon juice (Figure 9.23).

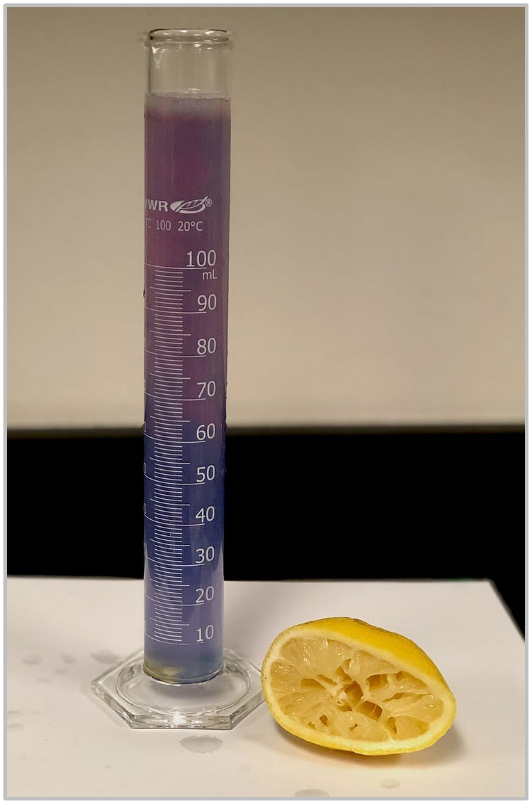


Figure 9.23 Blue Butterfly Pea Flower Tea can be used as an indicator to measure a pH value. In this image, the tea is being used to measure the pH value of lemon juice, as is indicated by the color change. [credit: Morgan Ouren. (2019). Acid in your Tea! CC BY.]

Acidosis and Alkalosis

Acidosis and alkalosis are not diseases. Rather, they are conditions formed for a variety of reasons. One example is our body's respiratory system. Our respiratory system contributes acid-base balance by regulating blood levels of carbonic acid produced in our lungs by reaction of CO_2 with water. In the balance condition, CO_2 and carbonic acid levels in the blood are in equilibrium. When the level of CO_2 in the blood increases, excess CO_2 reacts with water and makes additional carbonic acid, which lowers blood's pH level. An increase of acid concentration in the blood overwhelms buffer systems. Our blood then becomes acidic overtime and the condition of acidosis forms. The opposite of acidosis is alkalosis. Alkalosis is caused by excessive deep and continuous breathing, which reduces the level of CO_2 and carbonic acid in our blood. This imbalance results in an increase in blood pH, forming the alkalosis condition. In the acidosis and alkalosis conditions, kidneys try to affect blood pH by excreting excess acids or bases; however, kidneys make these adjustments more slowly than the lungs. Severe acidosis eventually results with coma. Figure 9.24 summarizes symptoms of acidosis and alkalosis conditions.

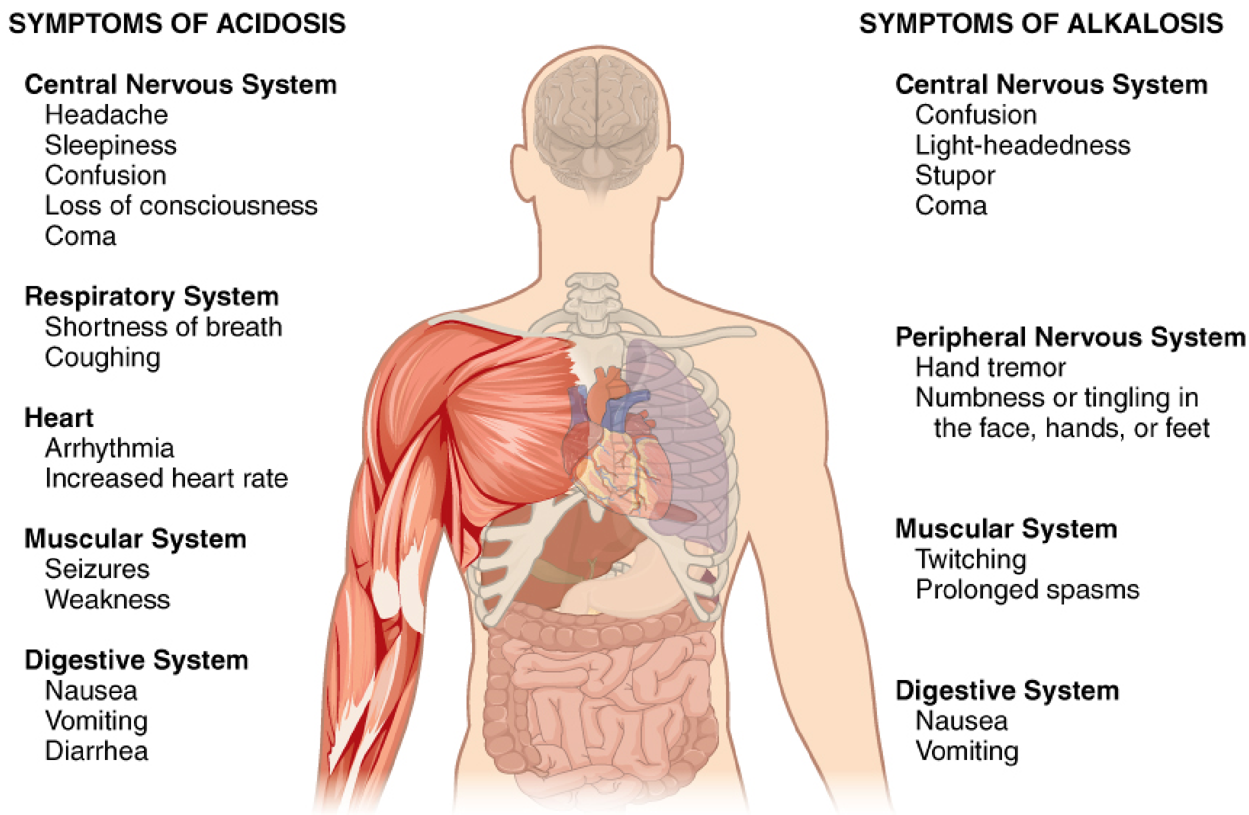


Figure 9.24 Symptoms of Acidosis (left). Symptoms of Alkalosis (right). [credit: *Anatomy and Physiology.* Figure 26.18. OpenStax. CC BY.]

Attributions

Betts, J. G., Young, K. A., Wise, J. A., Johnson, E., Poe, B., Kruse, D. H., Korol, O., Johnson, J. E., Womble, M., & DeSaix, P. (2013). *Anatomy and Physiology*. Houston, TX: OpenStax. CC BY. https://openstax.org/books/anatomy-and-physiology/pages/1-introduction

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