

## Chapter 11

### Sour and Bitter: Acids and Bases

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#### IN THIS CHAPTER

Discovering the properties of acids and bases

Finding out about the two acid-base theories

Determining what makes acids and bases strong

Using indicators to quickly identify bases and acids

Taking a look at the pH scale

Figuring out buffers

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Acids and bases are indicators of pH, and if you walk into any kitchen or bathroom, you find a multitude of each. In the refrigerator you find soft drinks full of carbonic acid. The pantry holds vinegar and baking soda, an acid and a base. Peek under the sink, and you notice the ammonia and other cleaners, most of which are bases. Check out that can of lye-based drain opener — it's highly basic. In the medicine cabinet, you find aspirin, an acid, and antacids, bases. The everyday world is full of acids and bases, and so is the everyday world of the industrial chemist. In this

chapter, I cover acids and bases and some good basic chemistry.

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# *Getting to Know the Properties of Acids and Bases:*

## *Macroscopic View*

Before you can fully grasp an appreciation for acids and bases, you first need a basic understanding of what they are. The following lists look at the properties of acids and bases that you can observe in everyday life.

Acids:

- Taste sour (but remember, in the lab, you test, not taste)
- Produce a painful sensation on the skin
- React with certain metals (magnesium, zinc, and iron) to produce hydrogen gas
- React with limestone and baking soda to produce carbon dioxide
- React with litmus paper and turn it red

Bases:

- Taste bitter (again, in the lab, you test, not taste)
- Feel slippery on the skin
- React with oils and greases
- React with acids to produce a salt and water

- React with litmus paper and turn it blue

Tables II-1 and II-2 show some common acids and bases found around the home.

**TABLE II-1 Common Acids Found in the Home**

<i>Chemical Name</i>	<i>Formula</i>	<i>Common Name or Use</i>
Hydrochloric acid	HCl	Muratic acid
Acetic acid	CH <sub>3</sub> COOH	Vinegar
Sulfuric acid	H <sub>2</sub> SO <sub>4</sub>	Auto battery acid
Carbonic acid	H <sub>2</sub> CO <sub>3</sub>	Carbonated water
Boric acid	H <sub>3</sub> BO <sub>3</sub>	Antiseptic; eye drops
Acetylsalicylic acid	C <sub>16</sub> H <sub>12</sub> O <sub>6</sub>	Aspirin

**TABLE II-2 Common Bases Found in the Home**

<i>Chemical Name</i>	<i>Formula</i>	<i>Common Name or Use</i>
Ammonia	$\text{NH}_3$	Cleaner
Sodium hydroxide	$\text{NaOH}$	Lye
Sodium bicarbonate	$\text{NaHCO}_3$	Baking soda
Magnesium hydroxide	$\text{Mg(OH)}_2$	Milk of magnesia
Calcium carbonate	$\text{CaCO}_3$	Antacid
Aluminum hydroxide	$\text{Al(OH)}_3$	Antacid

## Recognizing Acids and Bases: Microscopic View

If you look at Tables 11-1 and 11-2 closely, you may notice that all the acids contain hydrogen, and most of the bases contain the hydroxide ion ( $\text{OH}^-$ ). Two main theories of the structure of acids and bases use these facts in their descriptions of acids and bases and their reactions:

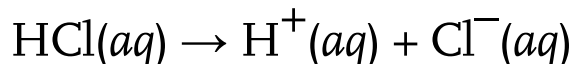
- Arrhenius theory
- Bronsted-Lowry theory

The following sections take a closer look at these two theories to help you gain a firmer understanding of acids and bases.

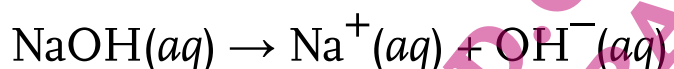
### *The Arrhenius theory: Must have water*

The Arrhenius theory was the first modern acid-base theory developed. In this theory, an acid is a substance that yields  $\text{H}^+$  (hydrogen) ions when dissolved in water, and a base is a substance that yields  $\text{OH}^-$  (hydroxide) ions when dissolved in water.  $\text{HCl(g)}$  can be considered a typical Arrhenius acid, because when this gas dissolves in water, it *ionizes* (forms ions) to give the  $\text{H}^+$  ion. (Chapter 13 is where you need to go

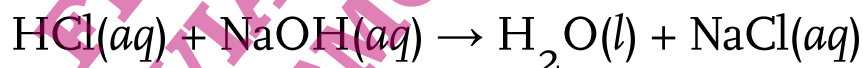
for the riveting details about ions.)



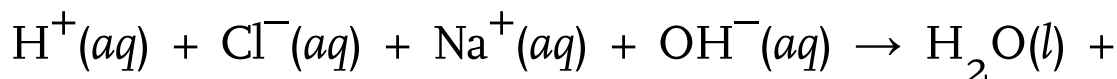
According to the Arrhenius theory, sodium hydroxide is classified as a base, because when it dissolves, it yields the hydroxide ion:

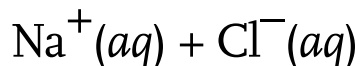


Arrhenius also classified the reaction between an acid and a base as a *neutralization* reaction, because if you mix an acidic solution with a basic solution, you end up with a neutral solution composed of water and a salt.

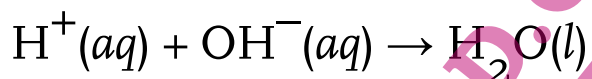


Look at the ionic form of this equation (the form showing the reaction and production of ions) to see where the water comes from:

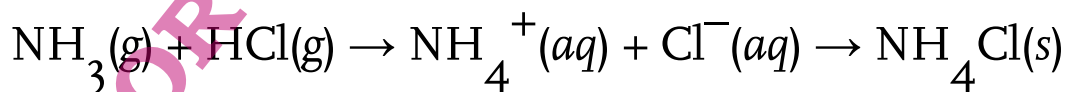




As you can see, the water is formed from combining the hydrogen and hydroxide ions. In fact, the net-ionic equation (the equation showing only those chemical substances that are changed during the reaction) is the same for all Arrhenius acid-base reactions:



The Arrhenius theory is still used quite a bit. But, like all theories, it has some limitations. It specifies that the reactions must take place in water and that bases must contain hydroxide ions, but many reactions that do not meet these stipulations resemble acid-base reactions. For example, look at the gas phase reaction between ammonia and hydrogen chloride gases:



The two clear, colorless gases mix, and a white solid of ammonium chloride forms. I show the intermediate



formation of the ions in the equation so that you can better see what's actually happening. The HCl transfers one  $\text{H}^+$  to the ammonia. That's basically the same thing that happens in the HCl/NaOH reaction, but the reaction involving the ammonia can't be classified as an acid-base reaction, because it doesn't occur in water and it doesn't involve the hydroxide ion. But again, the same basic process is taking place in both cases. In order to account for these similarities, a new acid-base theory was developed, the Bronsted-Lowry theory.

### *The Bronsted-Lowry acid-base theory: Giving and accepting*

The Bronsted-Lowry theory attempts to overcome the limitations of the Arrhenius theory by defining an acid as a proton ( $\text{H}^+$ ) donor and a base as a proton ( $\text{H}^+$ ) acceptor. The base accepts the  $\text{H}^+$  by furnishing a lone pair of electrons for a *coordinate-covalent* bond, which is a covalent bond (shared pair of electrons) in which one atom furnishes both of the electrons for the bond. Normally, one atom furnishes one electron for the bond and the other atom furnishes the second electron (see Chapter 14). In the coordinate-covalent

bond, one atom furnishes both bonding electrons.

Figure 11-1 shows the  $\text{NH}_3/\text{HCl}$  reaction using the electron-dot structures of the reactants and products. (Electron-dot structures are covered in Chapter 14, too.)



FIGURE 11-1: Reaction of  $\text{NH}_3$  with  $\text{HCl}$ .

$\text{HCl}$  is the acid, so it's the proton donor, and ammonia is the base, the proton acceptor. Ammonia has a lone pair of nonbonding electrons that it can furnish for the coordinate-covalent bond.

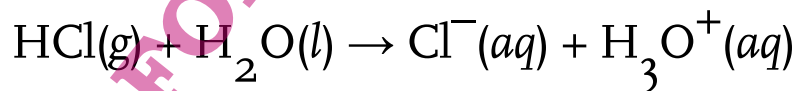
I discuss acid-base reactions under the Bronsted-Lowry theory in the section "Competing for protons: Bronsted-Lowry acid-base reactions," later in this chapter.

## Distinguishing between Strong and Weak Acids and Bases

I want to introduce you to a couple different categories of acids and bases — strong and weak. However, remember that acid-base strength is not the same as concentration. *Strength* refers to the amount of ionization or breaking apart that a particular acid or base undergoes. *Concentration* refers to the amount of acid or base that you initially have. You can have a concentrated solution of a weak acid, or a dilute solution of a strong acid, or a concentrated solution of a strong acid, or ... well, I'm sure you get the idea. The following sections point out the main differences between strong and weak acids and bases.

### *Ionizing completely: Strong acids*

Acids that ionize completely are considered strong. If you dissolve hydrogen chloride gas in water, the HCl reacts with the water molecules and donates a proton to them:



The  $\text{H}_3\text{O}^+$  ion is called the *hydronium ion*. This reaction goes

essentially to completion, meaning the reactants keep creating the product until they're all used up. In this case, all the HCl ionizes to  $\text{H}_3\text{O}^+$  and  $\text{Cl}^-$ ; no more HCl is present. Note that water, in this case, acts as a base, accepting the proton from the hydrogen chloride.

Because strong acids ionize completely, calculating the concentration of the hydronium ion and chloride ion in solution is easy if you know the initial concentration of the strong acid. For example, suppose that you bubble 0.1 moles (see Chapter 8 to get a firm grip on moles) of HCl gas into a liter of water. You can say that the initial concentration of HCl is 0.1 M (0.1 mol/L). *M* stands for molarity, and *mol/L* stands for moles of solute per liter. (For a detailed discussion of molarity and other concentration units, see Chapter 9.)

You can represent this 0.1 M concentration for the HCl in this fashion:  $[\text{HCl}] = 0.1$ . The brackets around the compound indicate molar concentration, or mol/L. Because the HCl completely ionizes, you see from the balanced equation that for every HCl that ionizes, you get one hydronium ion and one chloride ion. So the concentration of ions in that 0.1 M HCl solution is

$$[\text{H}_3\text{O}^+] = 0.1 \text{ M and } [\text{Cl}^-] = 0.1 \text{ M}$$

This idea is valuable when you calculate the pH of a solution. (And you can do just that in the section “Putting Coffee and Other Substances on the pH Scale,” later in this chapter.) Table II-3 lists the most common strong acids you’re likely to encounter.

**TABLE II-3 Common Strong Acids**

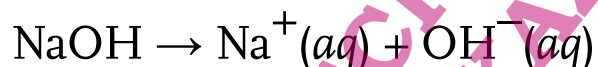
<i>Name</i>	<i>Formula</i>
Hydrochloric acid	HCl
Hydrobromic acid	HBr
Hydroiodic acid	HI
Nitric acid	$\text{HNO}_3$
Perchloric acid	$\text{HClO}_4$
Sulfuric acid (first ionization only)	$\text{H}_2\text{SO}_4$

Sulfuric acid is called a *diprotic* acid. It can donate 2 protons, but only the first ionization goes 100 percent. The other acids

listed in Table II-3 are *monoprotic* acids, because they donate only one proton.

### *Falling to pieces: Strong bases*

A strong base *dissociates* (breaks apart) completely in water. You normally see only one strong base, the hydroxide ion,  $\text{OH}^-$ . Calculating the hydroxide ion concentration is really straightforward. Suppose that you have a 1.5 M (1.5 mol/L) NaOH solution. The sodium hydroxide, a salt, completely dissociates into ions:



If you start with 1.5 mol/L NaOH, then you have the same concentration of ions:

$$[\text{Na}^+] = 1.5 \text{ M and } [\text{OH}^-] = 1.5 \text{ M}$$

### *Ionizing partway: Weak acids*

Acids that only partially ionize are called *weak acids*. One example is acetic acid ( $\text{CH}_3\text{COOH}$ ). If you dissolve acetic acid in water, it reacts with the water molecules, donating a proton

and forming hydronium ions. It also establishes an equilibrium in which you have a significant amount of unionized acetic acid. (In reactions that go to completion, the reactants are completely used up creating the products. But in equilibrium systems, two exactly opposite chemical reactions — one on each side of the reaction arrow — are occurring at the same place, at the same time, with the same speed of reaction.)

The acetic acid reaction with water looks like this:



The acetic acid that you added to the water is only partially ionized, so it's a weak acid. In the case of acetic acid, about 5 percent ionizes, and 95 percent remains in the molecular form. The amount of hydronium ion that you get in solutions of acids that don't ionize completely is much less than it is with a strong acid.

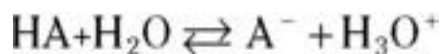


TIP

Calculating the hydronium ion concentration in weak acid solutions isn't as straightforward as it is in strong solutions, because not all the weak acid that dissolves initially has ionized. In order to calculate the hydronium

ion concentration, you must use the equilibrium constant expression for the weak acid. For weak acid solutions, you use a mathematical expression called the  $K_a$  — the *acid ionization constant*.

Take a look at the generalized ionization of some weak acid HA (hypothetical acid):



The  $K_a$  expression for this weak acid is

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$

Note that the  $[\text{HA}]$  represents the molar concentration of HA *at equilibrium*, not initially. Also, note that the concentration of water doesn't appear in the  $K_a$  expression, because there's so much water that it actually becomes a constant incorporated into the  $K_a$  expression.

Now go back to that acetic acid equilibrium. The  $K_a$  for acetic acid is  $1.8 \times 10^{-5}$ . The  $K_a$  expression for the acetic acid ionization is

$$K_a = 1.8 \times 10^{-5} = \frac{[\text{H}_3\text{O}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]}$$





Use this  $K_a$  to calculate the hydronium ion concentration in a 2.0 M solution of acetic acid.

1. Start by considering what information you have about the initial concentration and the products.

The initial concentration of acetic acid is 2.0 M, and you know that a little bit has ionized, forming a little hydronium ion and acetate ion. You also can see from the balanced reaction that for every hydronium ion that's formed, an acetate ion is also formed — so their concentrations are the same.

2. Using the information you've gathered from the balanced reaction, represent the amount of  $[H_3O^+]$  and  $[CH_3COO^-]$  as  $x$ :

$$[H_3O^+] = [CH_3COO^-] = x$$

3. In order to produce the  $x$  amount of hydronium and acetate ion, the same amount of

ionizing acetic acid is required. Represent the amount of acetic acid remaining at equilibrium as the amount you started with.

In this example, you started with 2.0 M, minus the amount that ionizes,  $x$ :

$$[\text{CH}_3\text{COOH}] = 2.0 - x$$

4. For the vast majority of situations, you can say that  $x$  is very small in comparison to the initial concentration of the weak acid. Therefore, you can often approximate the equilibrium concentration of the weak acid with its initial concentration.

In this case, you can say that  $2.0 - x$  is approximately equal to 2.0. The equilibrium constant expression now looks like this:

$$K_a = 1.8 \times 10^{-5} = \frac{[x][x]}{[2.0]} = \frac{[x]^2}{[2.0]}$$

5. Solve for  $x$ , which is the  $[\text{H}_3\text{O}^+]$ :
- $$(1.8 \times 10^{-5})(2.0) = [x]^2$$

$$\sqrt{3.6 \times 10^{-5}} = [x] = [\text{H}_3\text{O}^+]$$

$$6.0 \times 10^{-3} = [\text{H}_3\text{O}^+]$$

Refer to Table II-3 to see some common strong acids. Most of the other acids you encounter are weak.



One way to distinguish between strong and weak acids is to look for an acid ionization constant ( $K_a$ ) value. If the acid has a  $K_a$  value, then it's weak.

### *Finding equilibrium with water: Weak bases*

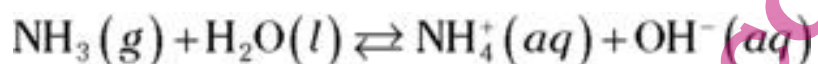
*Weak bases* also react with water to establish an equilibrium system. Ammonia is a typical weak base. It reacts with water to form the ammonium ion and the hydroxide ion:



Like a weak acid, a weak base is only partially ionized. The modified equilibrium constant expression for weak bases is  $K_b$ . You use it exactly the same way you use the  $K_a$  (see “Ionizing partway: Weak acids” for the details) except you solve for the  $[\text{OH}^-]$ .

## Competing for protons: Bronsted-Lowry acid-base reactions

With the Arrhenius theory, acid-base reactions are neutralization reactions. With the Bronsted-Lowry theory, acid-base reactions are a competition for a proton. For example, take a look at the reaction of ammonia with water:



Ammonia is a base (it accepts the proton), and water is an acid (it donates the proton) in the forward (left to right) reaction. But in the reverse reaction (right to left), the ammonium ion is an acid and the hydroxide ion is a base. If water is a stronger acid than the ammonium ion, then a relatively large concentration of ammonium and hydroxide ions are at equilibrium. If, however, the ammonium ion is a stronger acid, much more ammonia than ammonium ion is present at equilibrium.

Bronsted and Lowry said that an acid reacts with a base to form conjugate acid-base pairs, which differ by a single  $\text{H}^+$ .  $\text{NH}_3$  is a base, for example, and  $\text{NH}_4^+$  is its conjugate acid.  $\text{H}_2\text{O}$  is an acid in the reaction between ammonia and water, and  $\text{OH}^-$  is its conjugate base. In this reaction, the hydroxide ion is a strong base and ammonia is a weak base, so the

equilibrium is shifted to the left — not much hydroxide is present at equilibrium.

### *Playing both parts: Amphoteric water*

Water can act as either an acid or a base, depending on what it's combined with. When an acid reacts with water, water acts as a base, or a proton acceptor. But in reactions with a base (like ammonia; see the preceding section), water acts as an acid, or a proton donor. Substances that can act as either an acid or a base are called *amphoteric*.

But can water react with itself? Yes, it can. When two water molecules react with each other, one donates a proton and the other accepts it:



This reaction is an *equilibrium reaction*. A modified equilibrium constant, called the  $K_w$  (which stands for *water dissociation constant*), is associated with this reaction. The  $K_w$  has a value of  $1.0 \times 10^{-14}$  and has the following form:

$$1.0 \times 10^{-14} = K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$$

In pure water, the  $[\text{H}_3\text{O}^+]$  equals the  $[\text{OH}^-]$  from the balanced equation, so  $[\text{H}_3\text{O}^+] = [\text{OH}^-] = 1.0 \times 10^{-7}$ . The  $K_w$  value is a constant. This value allows you to convert from  $[\text{H}^+]$  to  $[\text{OH}^-]$ , and vice versa, in *any* aqueous solution, not just pure water.

In aqueous solutions, the hydronium ion and hydroxide ion concentrations are rarely going to be equal. But if you know one of them, the  $K_w$  allows you to figure out the other one.

Take a look at the 2.0 M acetic acid solution problem in the section “Ionizing partway: Weak acids,” earlier in this chapter.

You find that the  $[\text{H}_3\text{O}^+]$  is  $6.0 \times 10^{-3}$ . Now you have a way to calculate the  $[\text{OH}^-]$  in the solution by using the  $K_w$  relationship:

$$\begin{aligned}
 K_w &= 1.0 \times 10^{-14} = [\text{H}_3\text{O}^+][\text{OH}^-] \\
 &1.0 \times 10^{-14} = [6.0 \times 10^{-3}][\text{OH}^-] \\
 &1.0 \times 10^{-14} / 6.0 \times 10^{-3} = [\text{OH}^-] \\
 &1.7 \times 10^{-12} = [\text{OH}^-]
 \end{aligned}$$

## Identifying Acids and Bases with Indicators

*Indicators* are substances (organic dyes) that change color in the presence of an acid or base. You may be familiar with an acid-base indicator plant — the hydrangea. If it's grown in acidic soil, it turns pink; if it's grown in alkaline soil, it turns blue. Another common substance that acts as a good acid-base indicator is red cabbage. I have my students chop some up and boil it (most of them really love this part). They then use the leftover liquid to test substances. When mixed with an acid, the liquid turns pink; when mixed with a base, it turns green. In fact, if you take some of this liquid, make it slightly basic, and then exhale your breath into it through a straw, the solution eventually turns pink, indicating that the solution has turned slightly acidic. The carbon dioxide in your breath reacts with the water, forming carbonic acid:



Carbonated beverages are slightly acidic due to this reaction. Carbon dioxide is injected into the liquid to give it fizz. A little of this carbon dioxide reacts with the water to form carbonic acid. This reaction also explains why rainwater is slightly acidic. It absorbs carbon dioxide from the

atmosphere as it falls to earth.

In chemistry, indicators are used to indicate the presence of an acid or a base. Chemists have many indicators that change at slightly different pH levels. (You've probably heard the term *pH* used in various contexts. Me, I even remember it being used to sell deodorant and shampoo on TV. If you want to know what it actually stands for, check out the section "Putting Coffee and Other Substances on the pH Scale.") The two most commonly used indicators, which I discuss in the following sections, are

- Litmus paper
- Phenolphthalein

### *Taking a quick dip with litmus paper*

*Litmus* is a substance that is extracted from a type of lichen and absorbed into porous paper. (In case you're scheduled for a hot game of Trivial Pursuit this weekend, *lichen* is a plant that's made up of an alga and a fungus that live intimately together and mutually benefit from the relationship. Sounds kind of sordid to me.)

There are three different types of litmus:



- Red litmus is used to test for bases.
- Blue litmus is used to test for acids.
- Neutral litmus can be used to test for both.

If a solution is acidic, both blue and neutral litmus turn red. If a solution is basic, both red and neutral litmus turn blue. Litmus paper is a good, quick test for acids and bases. And you don't have to put up with the smell of boiling cabbage.

### *Titrating with phenolphthalein*

Phenolphthalein (pronounced fe-nul-*tha*-leen) is another commonly used indicator. Until a few years ago, phenolphthalein was used as the active ingredient in a popular laxative. In fact, I used to extract the phenolphthalein from the laxative by soaking it in either rubbing alcohol or gin (being careful not to drink it). I'd then use this solution as an indicator.

Phenolphthalein is clear and colorless in an acid solution and pink in a basic solution. It's commonly used in a procedure called a *titration*, where the concentration of an acid or base is determined by its reaction with a base or acid of known concentration.



Determine the molar concentration of an HCl solution.

To calculate, following these steps:

1. Place a known volume (say, 25.00 milliliters measured accurately with a pipette) in an Erlenmeyer flask and add a couple drops of phenolphthalein solution.

An *Erlenmeyer flask* is a flat-bottomed, conical-shaped container.

Because you're adding the indicator to an acidic solution, the solution in the flask remains clear and colorless.

2. Add small amounts of a standardized sodium hydroxide solution of known molarity (for example, 0.100 M) with a buret.

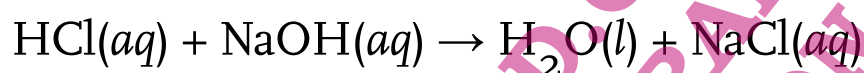
A *buret* is a graduated glass tube with a small opening and a stopcock, which helps you measure precise volumes of a solution.

3. Keep adding the base until the solution turns the

faintest shade of pink.

I call this the *endpoint* of the titration, the point in which the indicator shows that the acid has been exactly neutralized by the base. Figure 11-2 shows the titration setup.

Suppose that it takes 35.50 milliliters of the 0.100 M NaOH to reach the endpoint of the titration of the 25.00 milliliters of the HCl solution. Here's the reaction:



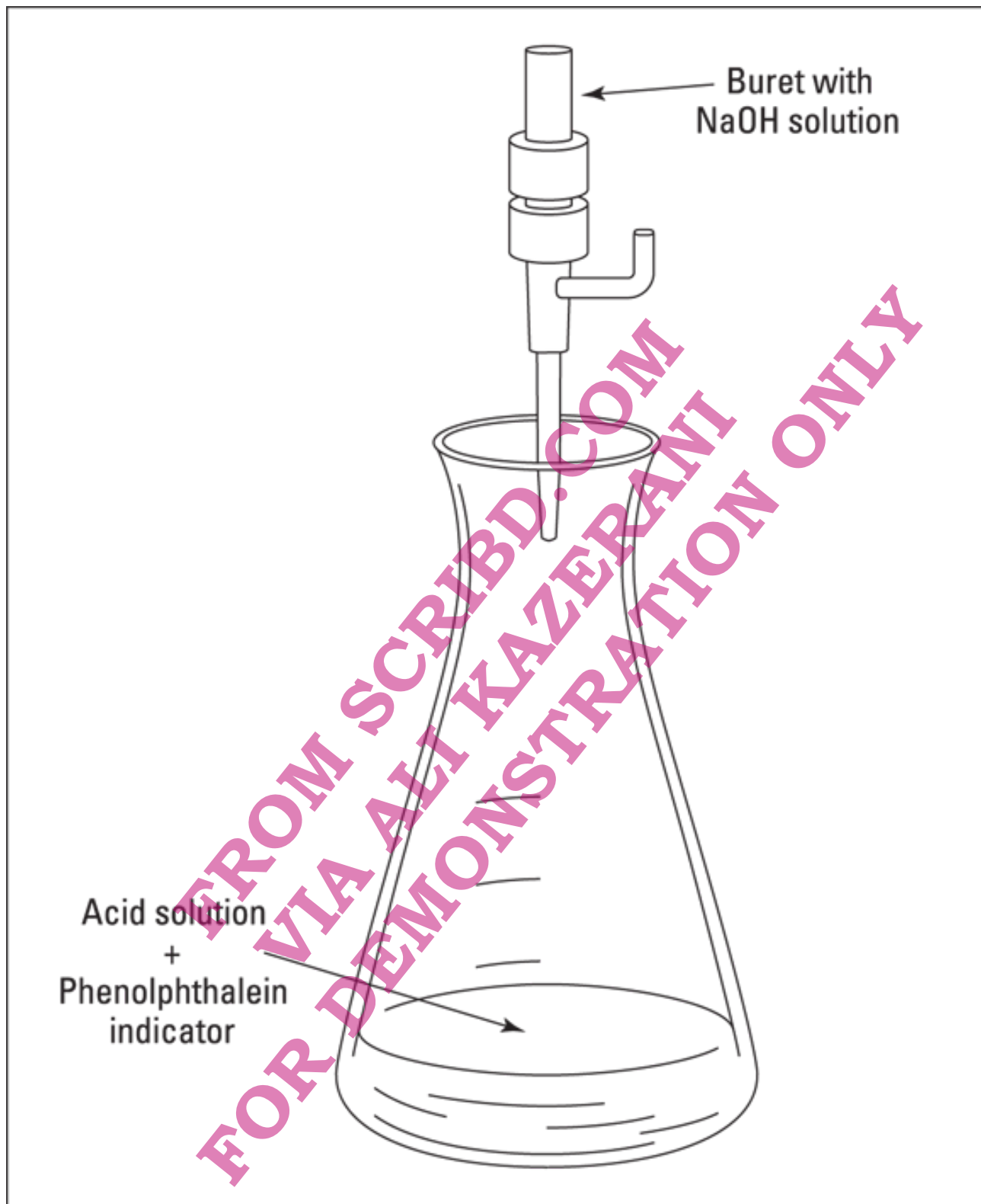


FIGURE II-2: Titration of an acid with a base.

From the balanced equation, you can see that the acid and base react in a 1:1 mole ratio. So if you can calculate the moles of bases added, you'll also know the number of moles of HCl present. Knowing the volume of the acid solution then allows you to calculate the molarity (note that you convert the milliliters to liters so that your units cancel nicely):

$$\frac{0.100\text{mol NaOH}}{1} \times \frac{0.03550\text{L}}{1} \times \frac{1\text{mol HCl}}{1\text{mol NaOH}} \times \frac{1}{0.02500\text{L}} = 0.142\text{M HCl}$$



TIP

You can calculate the titration of a base with a standard acid solution (one of known concentration) in exactly the same way, except the endpoint is the first disappearance of the pink color.

## Putting Coffee and Other Substances on the pH Scale

The amount of acidity in a solution is related to the concentration of the hydronium ion in the solution. The more acidic the solution is, the larger the concentration of the hydronium ion. In other words, a solution in which the  $[\text{H}_3\text{O}^+]$  equals  $1.0 \times 10^{-2}$  is more acidic than a solution in which the  $[\text{H}_3\text{O}^+]$  equals  $1.0 \times 10^{-7}$ . The *pH scale*, a scale based on the  $[\text{H}_3\text{O}^+]$ , was developed to more easily tell at a glance the relative acidity of a solution. *pH* is defined as the negative logarithm (abbreviated as *log*) of the  $[\text{H}_3\text{O}^+]$ . Mathematically, it looks like this:

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

Based on the water dissociation constant,  $K_w$  (see “Playing both parts: Amphoteric water,” earlier in this chapter), in pure water the  $[\text{H}_3\text{O}^+]$  equals  $1.0 \times 10^{-7}$ . Using this mathematical relationship, you can calculate the pH of pure water:

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

$$\text{pH} = -\log [1.0 \times 10^{-7}]$$

$$\text{pH} = -[-7]$$

$$\text{pH} = 7$$

The pH of pure water is 7. Chemists call this point on the pH scale *neutral*. A solution is called *acidic* if it has a larger  $[\text{H}_3\text{O}^+]$  than water and a smaller pH value than 7. A *basic* solution has a smaller  $[\text{H}_3\text{O}^+]$  than water and a larger pH value than 7.

The pH scale really has no end. You can have a solution of pH that registers less than 0. (A 10 M HCl solution, for example, has a pH of  $-1$ .) However, the 0 to 14 range is a convenient range to use for weak acids and bases and for dilute solutions of strong acids and bases. Figure 11-3 shows the pH scale.

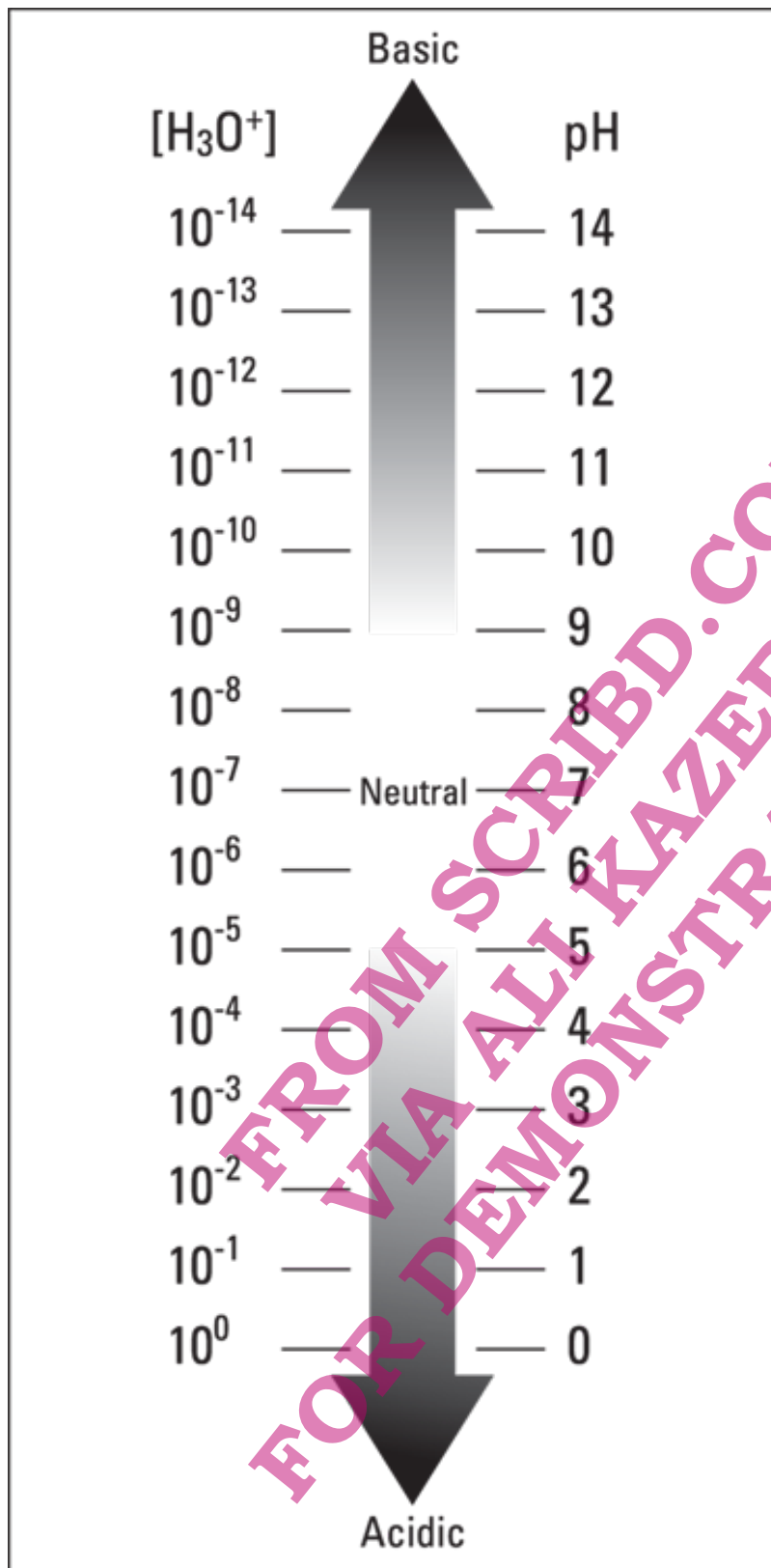


FIGURE II-3: The pH scale.



The  $[\text{H}_3\text{O}^+]$  of a 2.0 M acetic acid solution is  $6.0 \times 10^{-3}$ . Looking at the pH scale, you see that this solution is acidic. Now calculate the pH of this solution:

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

$$\text{pH} = -\log [6.0 \times 10^{-3}]$$

$$\text{pH} = -[-2.22]$$

$$\text{pH} = 2.22$$

In the section “Playing both parts: Amphoteric water,” I explain that the  $K_w$  expression enables you to calculate the  $[\text{H}_3\text{O}^+]$  if you have the  $[\text{OH}^-]$ . Another equation, called the *pOH*, can be useful in calculating the pH of a solution. The *pOH* is the negative logarithm of the  $[\text{OH}^-]$ . You can calculate the *pOH* of a solution just like the pH by taking the negative log of the hydroxide ion concentration. If you use the  $K_w$  expression and take the negative log of both sides, you get  $14 = \text{pH} + \text{pOH}$ . This equation makes going from *pOH* to pH quite easy.

Just as you can convert from  $[\text{H}_3\text{O}^+]$  to pH, you can also go from pH to  $[\text{H}_3\text{O}^+]$ . To do this, you use what’s called the *antilog relationship*, which is

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}}$$

Human blood, for example, has a pH of 7.3. Here's how you calculate the  $[\text{H}_3\text{O}^+]$  from the pH of blood:

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}}$$

$$[\text{H}_3\text{O}^+] = 10^{-7.3}$$

$$[\text{H}_3\text{O}^+] = 5.01 \times 10^{-8}$$

The same procedure can be used to calculate the  $[\text{OH}^-]$  from the pOH.

## ANTACIDS: GOOD, BASIC CHEMISTRY

Go to any drugstore or grocery store and look at the shelves upon shelves of antacids. They represent acid-base chemistry in action!

The stomach secretes hydrochloric acid in order to activate certain enzymes (biological catalysts) in the digestion process. But sometimes the stomach produces too much acid, or the acid makes its way up into the esophagus (leading to heartburn or acid reflux), and you need to neutralize the excess acid with — you guessed it — a base. The basic

formulations that are sold to neutralize this acid are called *antacids*. Antacids include the following compounds as active ingredients:

- **Bicarbonates** —  $\text{NaHCO}_3$  and  $\text{KHCO}_3$
- **Carbonates** —  $\text{CaCO}_3$  and  $\text{MgCO}_3$
- **Hydroxides** —  $\text{Al(OH)}_3$  and  $\text{Mg(OH)}_2$

Trying to select the “best” antacid for occasional use can be complicated. Certainly price is a factor, but the chemical nature of the bases can also be a factor. For example, individuals with high blood pressure may want to avoid antacids containing sodium bicarbonate because the sodium ion tends to increase blood pressure. Individuals concerned about loss of calcium from the bones, or *osteoporosis*, may want to use an antacid containing calcium carbonate. However, both calcium carbonate and aluminum hydroxide can cause constipation if used in large doses. On the other hand, large doses of both magnesium carbonate and magnesium hydroxide can act as laxatives. Selecting an antacid can really be a balancing act!

Substances commonly found in everyday life cover a wide range of pH values. Table II-4 lists some common substances and their pH values.

**TABLE II-4 Average pH Values of Some Common Substances**

<i>Substance</i>	<i>pH</i>
Oven cleaner	13.8
Hair remover	12.8
Household ammonia	11.0
Milk of magnesia	10.5
Chlorine bleach	9.5
Seawater	8.0
Human blood	7.3
Pure water	7.0
Milk	6.5
Black coffee	5.5

Soft drinks	3.5
Aspirin	2.9
Vinegar	2.8
Lemon juice	2.3
Auto battery acid	0.8

To sustain life, human blood must stay within about  $\pm 0.2$  pH units of 7.3, a narrow range. Many things in our environment, such as foods and hyperventilation, can act to change the pH of our blood. Buffers help to regulate blood pH and keep it in the 7.1 to 7.5 range.

## Controlling pH with Buffers

*Buffers*, or *buffer solutions* as they're sometimes called, resist a change in pH caused by the addition of acids or bases. Obviously, the buffer solution must contain something that reacts with an acid — a base. Something else in the buffer solution reacts with a base — an acid. In general, buffers come in two types:

- **Mixtures of weak acids and bases:** The mixtures of weak acids and bases may be conjugate acid-base pairs (such as  $\text{H}_2\text{CO}_3/\text{HCO}_3^-$ ) or nonconjugate acid-base pairs (such as  $\text{NH}_4^+/\text{CH}_3\text{COO}^-$ ). (For more info about conjugate acid-base pairs, see “Competing for protons: Bronsted-Lowry acid-base reactions,” earlier in this chapter.)

In the body, conjugate acid-base pairs are more common. In the blood, for example, the carbonic acid/bicarbonate pair helps to control the pH. This buffer can be overcome, though, and some potentially dangerous situations can arise. If a person exercises strenuously, lactic acid from the muscles is released into the

bloodstream. If there's not enough bicarbonate ion to neutralize the lactic acid, the blood pH drops, and the person is said to be in *acidosis*. Diabetes may also cause acidosis. On the other hand, if a person hyperventilates (breathes too fast), she breathes out too much carbon dioxide. The carbonic acid level in the blood is reduced, causing the blood to become too basic. This condition, called *alkalosis*, can be very serious.

- **Amphoteric species:** Amphoteric species may also act as buffers by reacting with an acid or a base. (For an example of an amphoteric species, see "Playing both parts: Amphoteric water," earlier in this chapter.) The bicarbonate ion ( $\text{HCO}_3^-$ ) and the monohydrogen phosphate ion ( $\text{HPO}_4^{2-}$ ) are amphoteric species that neutralize both acids and bases. Both of these ions are also important in controlling the blood's pH.

## ACIDS WITH BAD PRESS: AN INTRODUCTION TO ACID RAIN

Over the past decade or so, acid rain has emerged as a great environmental problem. Natural rainwater is somewhat acidic (around pH 5.6) due to the absorption of carbon

dioxide from the atmosphere and the creation of carbonic acid. However, when acid rain is mentioned in the press, it usually refers to rain in the pH 3 to 3.5 range.

The two major causes of acid rain are automotive and industrial pollution. In the automobile's internal combustion engine, nitrogen in the air is oxidized to various oxides of nitrogen. These nitrogen oxides, when released into the atmosphere, react with water vapor to form nitric acid ( $\text{HNO}_3$ ).

In fossil fuel power plants, oxides of sulfur are formed from the burning of the sulfur impurities commonly found in coal and petroleum. These oxides of sulfur, if released into the atmosphere, combine with water vapor to form both sulfuric and sulfurous acids ( $\text{H}_2\text{SO}_4$  and  $\text{H}_2\text{SO}_3$ ). Oxides of nitrogen are also produced in these power plants.

These acids fall to earth in the rain and cause a multitude of problems. They dissolve the calcium carbonate of marble statues and monuments. They decrease the pH of lake water to such a degree that fish can no longer live in the lakes. They cause whole forests to die or become stunted. They react with the metals in cars and buildings.

Industrial controls have been somewhat effective in



reducing the problem, but acid rain is still a major environmental issue. (See Chapter 18 for more info about acid rain.)

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