Acids, Bases, and Buffers

OVERVIEW

You're probably familiar with acids and bases in the products you use at home. Rust removers often contain phosphoric acid. Muriatic acid (a common name for hydrochloric acid) is used to clean swimming pools. Mild acids can be found in some foods that have a sour taste, such as lemon juice, which contains citric acid, and vinegar, which contains dilute acetic acid. Cleaning products like oven cleaners and drain cleaners often have a slippery feel. These are bases that contain the substance sodium hydroxide. What exactly do we mean when we say something is an acid or a base? Do the acids have anything in common with each other? What about the bases?

The definition of acids and bases is just one of several topics we'll explore in this Print Presentation. Once we understand what acids and bases are, we'll consider ways to quantify the amount of acid or base present in a solution—including an explanation of pH and the pH scale, a convenient way of expressing amounts of acid or base. Finally, we'll focus on weak acids and bases, with particular emphasis on some biologically relevant acids and bases.

ACIDS AND ACIDITY

Definitions of Acids and Bases

In the early 20th century, the Danish chemist Johannes Brønsted and the English chemist Thomas Lowry each came up with the definitions of acids and bases that are widely used today. Brønsted and Lowry proposed that an **acid** is a proton donor and a **base** is a proton acceptor.

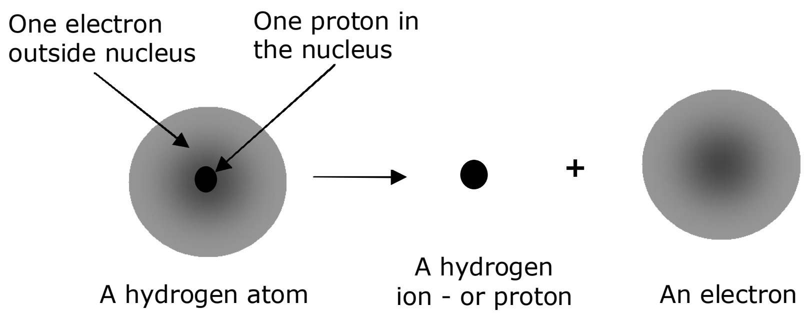
Acid:
A proton donor

Base:
A proton acceptor

Let's review what we mean by a proton, a term we learned when we studied atoms and molecules. A **proton** is a positively charged subatomic particle located in the nucleus. But when we're talking about acids and bases, there's a better way to think about a proton. Hydrogen's atomic number is 1, so it has one proton in its nucleus and one electron outside the nucleus. Its atomic mass is also 1, so it has no neutrons. When a hydrogen atom donates an electron to become a hydrogen ion, all that's left is a proton, as shown in Figure 1.

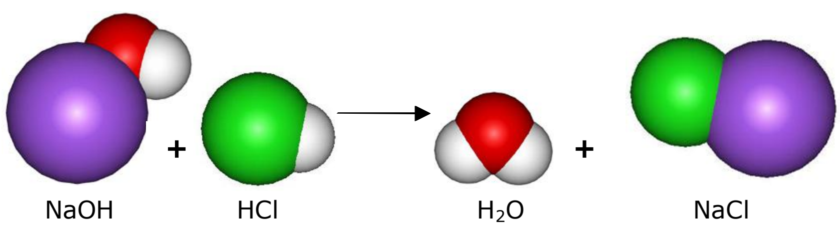
A hydrogen ion (H^+) is just a proton.

Figure 1. Hydrogen donates one electron to become a hydrogen ion.



Let's look at the reaction between hydrochloric acid (HCl) and sodium hydroxide (NaOH) to see how the definitions of acids and bases really work. As shown in Figure 2, HCl donates a proton to NaOH. This reaction makes HCl an acid and NaOH a proton acceptor, or base, according to the definitions.

Figure 2. In the reaction between HCl and NaOH, HCl is the proton donor (or acid), and NaOH is the proton acceptor (or base).



HCl is a proton donor and NaOH is a proton acceptor.

Most reactions between an acid and a base, like the one in Figure 2, don't take place between isolated molecules. Instead, they occur when both substances are dissolved in water. HCl and NaOH are both ionic compounds, so they dissolve in water to produce ions. The notation "aq" (for "aqueous" or water-based) after each chemical formula indicates that the substance is dissolved in water. Hydrogen ions and chloride ions are produced when HCl dissolves, and sodium ions and hydroxide ions are produced when NaOH dissolves.

$$HCl\ (aq) \rightarrow H^+\ (aq) + Cl^-\ (aq)$$

$$NaOH\ (aq) \rightarrow Na^+\ (aq) + OH^-\ (aq)$$

It's the concentration of these ions in solution that determines the extent to which a solution is acidic or basic.

The Mole

Chemists use a special unit of measure to count atoms, molecules, and other sub-microscopic objects. This unit of measure is called the mole. It's *not* the animal that digs tunnels through golf courses. It's also not a beauty mark!

In chemistry, a **mole** is defined as the amount of a substance that contains as many *particles* as there are *atoms* in exactly 12 grams of carbon-12. The particles being measured could be anything a chemist might be interested in: atoms, molecules, ions, even electrons. A mole measures substances by counting particles. You can think of it as the chemist's dozen, except that it's a lot bigger than a dozen. A mole of particles is actually a huge number of particles. Since atoms are so incredibly tiny, there's an enormous number of them in 12 grams of carbon. The definition of a mole doesn't actually state the number, but scientists have found that there are 6.02×10^{23} particles in 1 mole. This number of particles is called Avogadro's number, named after the Italian physicist Amedeo Avogadro.

Mole:
The amount of a substance that contains exactly as many particles as there are in exactly 12 grams of carbon-12

Molarity

Most chemical reactions in living systems take place in **solutions** formed when a substance, the **solute**, is dissolved in a liquid, the **solvent**. As discussed earlier, water is the "universal solvent" in living organisms, so most biological solutions are water-based or aqueous. The amount of solute dissolved in the solvent, also called the concentration, can vary from solution to solution, so we need a way to describe this quantity.

Solution:
A mixture of two or more substances that appears uniform at the macroscopic (visible) level

Chemists use a quantity called molarity to express the concentration of a solution. **Molarity** is the number of moles of solute per liter of solution. When used as an adjective, we say that a solution is 1 molar, or 1 M for short. A 1 M solution contains 1 mole of solute per liter of solution. This can be 1 mole of HCl, 1 mole of NaOH, or 1 mole of anything per liter of solution. Similarly, a 0.5 M solution contains 0.5 moles of solute per liter of solution. A 1 M solution is twice as concentrated as a 0.5 M solution, since it contains twice as many moles of solute per liter of solution. A substance's concentration is often represented by using square brackets around the formula. For example, a 0.5 M solution of HCl is sometimes indicated as [HCl] = 0.5 M.

Molarity (M):
The concentration of a solution in units of moles of solute per liter of solution

Aqueous solutions vary tremendously in their hydrogen ion, or proton, concentration. Some solutions have a very high hydrogen ion concentration, on the order of 1 M. Other solutions have extremely low hydrogen ion concentrations, perhaps as low as 1.0×10^{-14} M. Pure water has a hydrogen ion concentration of 1.0×10^{-7} M. The smaller the number, the lower the hydrogen ion concentration and the lower the acidity of the solution.

The pH Scale

Working with negative exponents like 1.0×10^{-14} isn't a very convenient way of describing the acidity of a solution. The **pH** scale is one way to express the acidity of a solution without using cumbersome numbers. The term pH is derived from the French for the "power of hydrogen." The power aspect comes from the fact that the pH scale is based on powers of 10.

pH:
- Scale that describes the acidity of a solution
- Derived from French for "power of hydrogen"
- Based on powers of 10
- Equals $-\log[H^+]$

Here's the formula for determining a pH value:

$$pH = -\log[H^+]$$

We can use this formula to convert any hydrogen ion concentration to a pH value. Suppose we have a sample of pure water, where $[H^+] = 1.0 \times 10^{-7}$ M. We can calculate its pH as follows:

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

We can also calculate the hydrogen ion concentration if we know the pH. Suppose we have a solution whose pH is 6. Its hydrogen ion concentration can be calculated as follows:

$$6 = -\log[H^+]$$
$$-6 = \log[H^+]$$
$$10^{-6} = [H^+]$$

As pH increases by one unit, $[H^+]$ decreases by a factor of 10.

As pH decreases by 1 unit, $[H^+]$ increases by a factor of 10.

Since pH is a logarithmic scale based on powers of 10, a pH change of 1 unit means a 10-fold change in the concentration of hydrogen ions. As we can see in the previous examples, a pH 6 solution has a hydrogen ion concentration 10 times higher than a pH 7 solution.

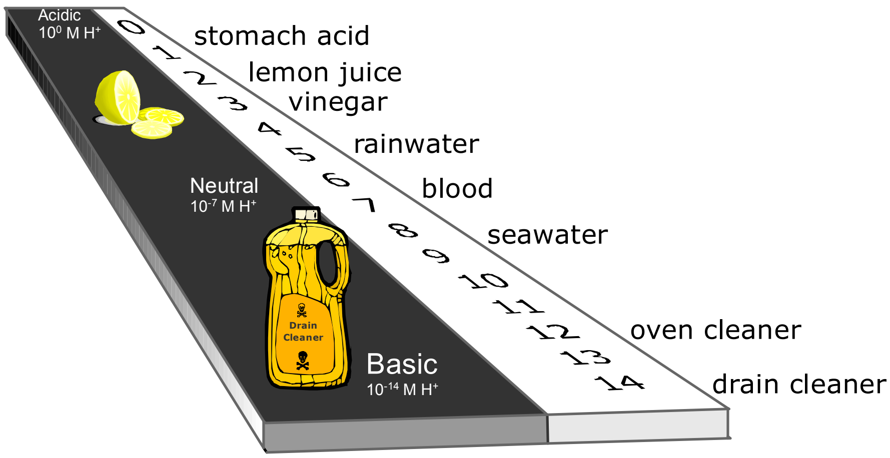
The pH scale ranges from 0 to 14, as shown in Figure 3. A neutral solution has a pH of 7. An acidic solution has a pH less than 7, and a basic solution has a pH greater than 7. Earlier we mentioned that lemon juice and vinegar are both acids. We'd expect them to have pH values less than 7, and that's just what we find when we measure their pH values in the lab. Oven cleaners and drain cleaners, both bases, have pH values greater than 7. There's also a wide range of pH values in biological solutions. Stomach acid, one of the most acidic body fluids, has a pH between 1 and 3. Normal rainwater has a pH just under 6, blood has a pH of about 7.4, and seawater has a pH between 7 and 8.5.

Neutral solutions have pH = 7.

Acidic solutions have pH < 7.

Basic solutions have pH > 7.

Figure 3. The pH scale ranges from 0 to 14.



Notice that the pH on the upper scale, is the absolute value of the negative exponent of the $[H^+]$ on the lower scale

WEAK ACIDS AND BASES

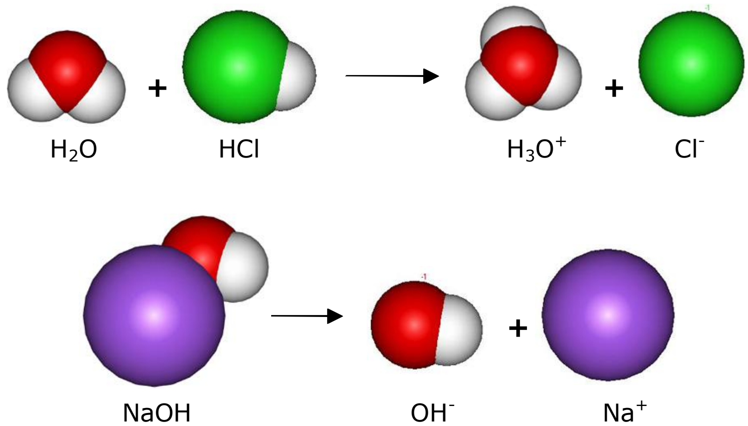
Strong Versus Concentrated

Which is more concentrated, a 1 M HCl solution or a 10^{-5} M HCl solution? As we've seen, the solution with the higher molarity is more concentrated, since there are more moles of solute, HCl, per liter of solution. The answer is obviously the 1 M solution.

Now suppose you were asked which has the higher pH, the 1 M HCl solution or the 10^{-5} M HCl solution? We can predict the answer because hydrochloric acid is a **strong acid**. This means that it dissociates, or separates, completely into its ions in solution. The released proton quickly reacts with any water molecules to form the hydronium ion, H_3O^+. Similarly, **strong bases** dissociate completely into their ions. Figure 4 illustrates the dissociation of HCl, a strong acid, and NaOH, a strong base.

Strong acid (base):
An acid (base) which dissociates completely into its ions in solution

Figure 4. Strong acids and bases dissociate completely into their ions.



The pH of a strong acid like HCl can be calculated from the molarity of the undissociated (original) acid. Therefore, the 1 M HCl solution has a pH of 0 and the 10^{-5} M HCl solution has a pH of 5. In other words, a more concentrated solution of a given acid will have a higher pH than a more dilute solution of the same acid.

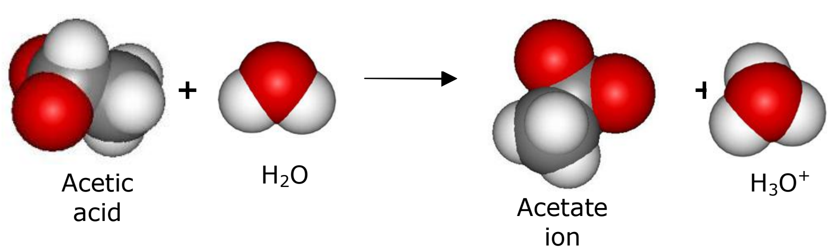
Weak Acid or Base Equilibrium

Now let's try something trickier: which will have the higher pH, a 1 M HCl solution or a 1 M acetic acid (vinegar) solution? If you guessed that they have the same pH, without having any more information about acetic acid, that would be a very good guess.

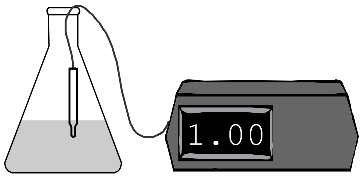
Acetic acid is an example of a weak acid. A **weak acid** or **weak base** is one that dissociates only partially in solution. Unlike strong acids, which undergo a complete, irreversible reaction, weak acids donate hydrogen ions to water in a reversible reaction. This reaction is shown in Figure 5, using acetic acid, CH_3COOH, as an example. The pair of arrows tells us the reaction is occurring in the forward and reverse directions simultaneously, a condition known as dynamic equilibrium, or just equilibrium.

Weak acid (base): An acid (base) that dissociates only partially in solution.

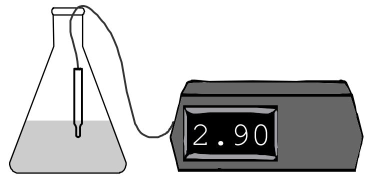
Figure 5. A weak acid dissociates only partially in solution.



Let's compare the pH of a 0.1 M HCl solution to the pH of a 0.1 M acetic acid solution.



0.1 M HCl



0.1 M CH_3COOH

Since all of the HCl dissociates, its pH is exactly what we'd predict if we used the original acid molarity in our calculations: 1.0. But why is the pH of the acetic acid solution so much higher? Let's calculate the hydrogen ion concentration from the pH and see if we can make sense of this observation:

$$2.9 = -\log[H^+]$$
$$-2.9 = \log[H^+]$$
$$[H^+] = 10^{-2.9} = 1.3 \times 10^{-3} \text{ M } H^+$$

Even though we started with 0.1 M acetic acid, we only got 1.3×10^{-3} M H^+ in solution. In other words, only about 1.3 % of the original acid dissociated. The rest remained as the undissociated acid when the reaction reached equilibrium:

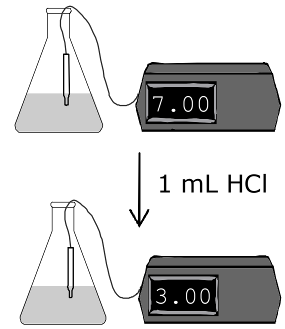
	$CH_3COOH + H_2O$ —	$CH_3COO^- + H_3O^+$
At equilibrium:	98.7 %	1.3 %

That's exactly what we mean when we say an acid dissociates only partially.

Weak acids can vary in the extent to which they dissociate. A quantitative term called the acid dissociation constant, or K_a, lets us compare the ionization of different acids. A larger value means that more of the acid will dissociate to produce hydrogen ions, and a smaller value means that less of the acid will dissociate. When you have two acid solutions at the same concentration, such as 0.1 M, the acid with the higher K_a value will dissociate more than the acid with the lower K_a value. This also means that the acid with the higher K_a will have the lower pH at equilibrium.

Next we'll consider what happens when we have a mixture of undissociated and dissociated forms of a weak acid or base.

Buffers

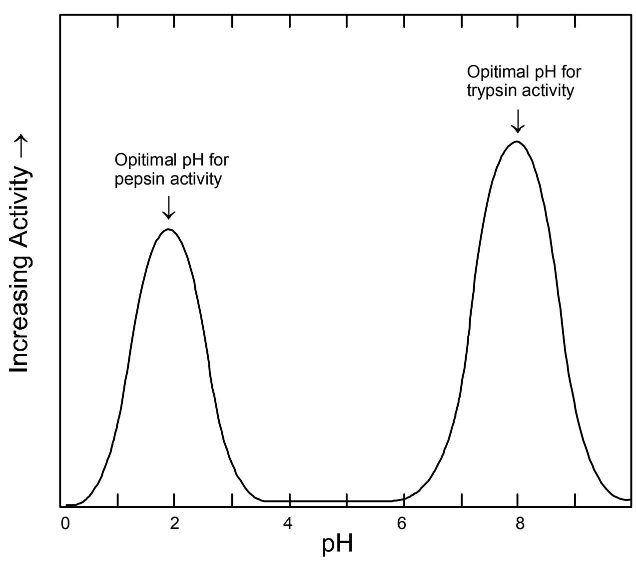


Let's think about what happens when we add 1 mL of 1 M HCl to 1 L of pure water. The pH of the water drops from 7 to 3 — that's a 10,000-fold increase in the acid concentration, just from adding this small amount of acid!

Biological systems can't tolerate such huge changes in pH. Most of the chemical reactions in our body are carried out by special proteins called enzymes. Enzymes function best in a narrow range of pH values, and don't function at all at pH values more than 1–2 units away from their optimal pH.

Figure 6 shows the activity of two digestive enzymes, pepsin and trypsin, as a function of pH. Pepsin is found in the stomach, which has a pH around 2, and this number shows that pepsin has evolved to function best close to pH 2. Compare that to trypsin, which is found in the intestines. Intestinal pH is usually between 7.5 and 8.5, and that's precisely the optimal value for trypsin's function. As Figure 6 illustrates, if the pH drops below 6 or rises above 10, trypsin won't function, and this could have fatal consequences for the organism.

Figure 6. Enzymes function best in a very narrow pH range.



There are other important biological reasons for maintaining pH in a very narrow range. For example, blood has an optimal pH of 7.4. The ability of hemoglobin to carry oxygen is dramatically affected by very small pH changes. A deviation of more than about 0.6 pH units can be fatal because hemoglobin can't carry oxygen to our cells. Let's think about what this means in terms of added acid. As we saw earlier, adding just 1 mL of acid to 1 L of water caused a pH change of 4 units. If we can let the pH drop by only 0.6 units, we can add no more than 4 x 10^{-7} mL of acid. That's a really small volume!

Maintaining a constant pH is critical for biological systems. Substances called buffers are present in most living cells and biological fluids. A **buffer** is a solution that resists changes in the concentration of hydrogen ions caused by the addition of an acid or a base. A buffer is a mixture of weak acid and the negative ion produced when the acid donates a proton. This ion has a special name: the conjugate base of the acid. A **conjugate base** can act as a base, or proton acceptor, when the reaction occurs in the reverse direction. In blood, one of the most important buffers is the carbonic acid - carbonate system:

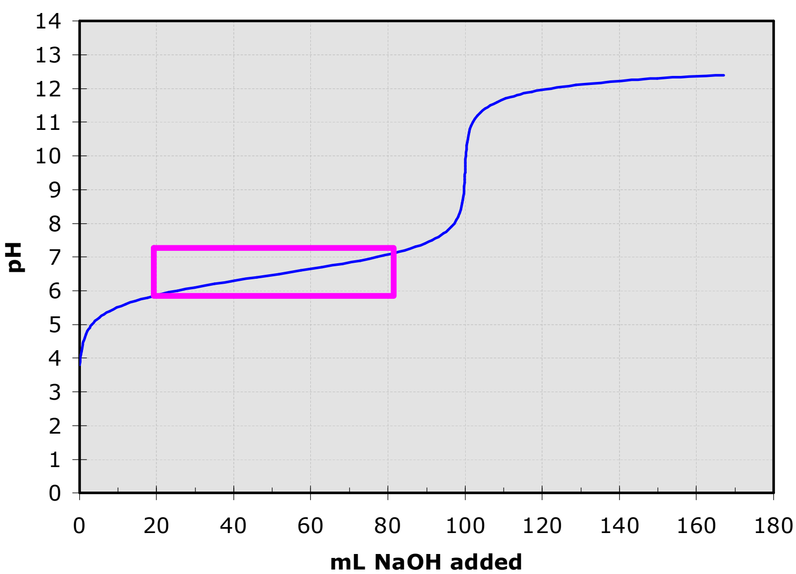
Buffer:
A solution that resists changes in the concentration of hydrogen ions caused by the addition of an acid or a base

$$H_2CO_3 \text{ — } HCO_3^- + H^+$$

Carbonic acid (acid) — **Carbonate ion** (conjugate base of carbonic acid)

A graph like the one in Figure 7 illustrates just how well a buffer can resist pH changes.

Figure 7. The carbonic acid-carbonate buffer system helps maintain the pH of blood.



In this example, we start with a 100 mL solution of 0.1 M carbonic acid at about pH 4. When we add 0.1 M NaOH, a base, a few milliliters at a time, the pH increases quickly until about 10 mL have been added. Then, as we continue to add base, the pH increases only gradually until about 100 mL have been added. The pH range over which there is only a small pH change upon the addition of an acid or base is called the **buffer**

Buffer zone:
The pH range over which there is only a small pH change upon the addition of acid or base to a buffer solution

zone. Different buffers have different buffer zones, based on their individual K_a values, so different buffers are used in different biological systems. The carbonate buffer system is common in blood, and the phosphate buffer system is common in many types of cells:

$$H_2PO_4^- \;—\; HPO_3^{2-} + H^+$$

Dihydrogen phosphate (acid)	**Monohydrogen phosphate** (conjugate base of dihydrogen phosphate)

Figure 7 illustrates what's meant by a buffer: the carbonic acid solution resists changes in pH when base is added. Consider what would have happened if we added NaOH directly to a solution at pH 4 that didn't contain a buffer. The pH would have jumped to 12.5 upon the addition of 50 mL of NaOH. If this pH change took place in a living organism, it probably wouldn't be living any more!

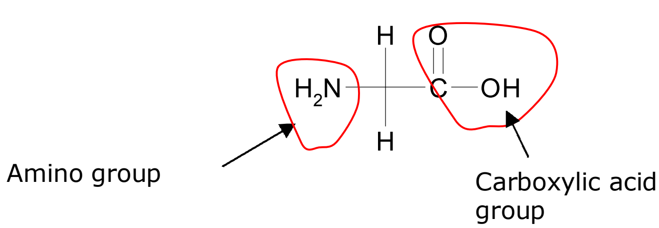
Zwitterions

So far we've considered only substances that are either proton donors or proton acceptors. Now let's look at some compounds that contain both acidic and basic groups. An important class of weak acids and bases are called **zwitterions**, ions that are simultaneously positively and negatively charged. The most familiar zwitterions in biology are the amino acids, which we'll learn much more about in the next chapter.

Zwitterion:
An ion that is simultaneously negatively and positively charged

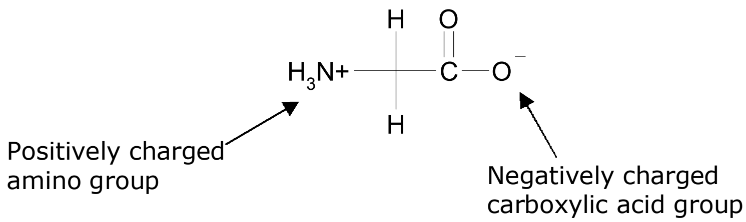
For now, let's focus on their acid-base properties using the simplest amino acid, glycine. Amino acids are named for the fact that they contain a group of atoms called an amino group, and a group of atoms called a carboxylic acid group (Figure 8).

Figure 8. Amino acids contain an amino group and a carboxylic acid group.



The amino group acts as a base or proton acceptor, and becomes positively charged. The carboxylic acid group acts as an acid or proton donor and becomes negatively charged, as shown in Figure 9.

Figure 9. The zwitterions form of an amino has a positively charged amino group and a negatively charged carboxylic acid group.



SUMMARY

Key Points about Acids and Acidity

- An acid is a proton (hydrogen ion) donor.
- A base is a proton (hydrogen ion) acceptor.
- A mole is the unit of measure used by chemists to count atoms, molecules, and other sub-microscopic objects.
- A mole is the amount of a substance that contains exactly as many particles as there are in exactly 12 grams of carbon-12.
- Molarity is the concentration of a solution in units of moles of solute per liter of solution.
- pH equals [is "equals" OK? "is determined by"? calculated by?] the negative logarithm of the hydrogen ion concentration ($-\log[H^+]$).
- As pH increases by 1 unit, $[H^+]$ decreases by a factor of 10, and as pH decreases by 1 unit, $[H^+]$ increases by a factor of 10.
- Acidic solutions have pH less than 7, neutral solutions have pH equal to 7, and basic solutions have pH greater than 7.

Key Points about Weak Acids and Bases

- A strong acid is one that dissociates completely into its ions.
- A weak acid is one that dissociates only partially in solution.
- The pH of a weak acid solution is much higher than the pH of a strong acid solution of the same concentration.
- A buffer is a solution that resists changes in the concentration of hydrogen ions caused by the addition of an acid or a base.
- A buffer contains a mixture of a weak acid and its conjugate base.
- Each buffer has a characteristic buffer zone, the pH range over which there is only a small pH change upon addition of acid or base.

- A zwitterion is an ion that is simultaneously negatively and positively charged.
- Amino acids are common biological examples of zwitterions.