Strong and Weak Acids and Bases

8.2

Many medicines are acids or bases. Tylenol (acetaminophen) and Aspirin (acetylsalicylic acid, or ASA) are two common painkillers that are acids. The pH of a 0.25 mol/L solution of acetaminophen is 5.3 while the pH of an equally concentrated solution of ASA is 2.0. Why the difference? Since the K_a of acetaminophen is 1.2×10^{-10} while the K_a of ASA is 3.27×10^{-4}, acetaminophen is much less likely than ASA to ionize to form $H^+(aq)$ ions in aqueous solution. This makes a difference to patients who need relief from pain but cannot tolerate additional acidity in their stomachs.

In earlier Chemistry courses you probably learned a little about strong and weak acids. You may also be aware that there are strong and weak bases. What decides the strength of an acid or a base? This property depends on the equilibrium position of the compound's ionization reaction. Stronger acids and bases have ionization equilibrium positions farther to the right. Weaker acids and bases have equilibrium positions farther to the left. Thus, in water, strong acids or bases will ionize more than weak acids or bases. ASA is a stronger acid than is acetaminophen.

Strong Acids and Weak Acids

When acids dissolve and ionize in water, they form a dynamic equilibrium between reactants and products. A **strong acid** is an acid for which the equilibrium position in an aqueous solution lies far to the right. This means that, at equilibrium, almost all the HA molecules have broken apart to produce ions (**Figure 1(a)**). **Table 1** summarizes some different ways of defining a strong acid. A **weak acid** is one for which the equilibrium position is far to the left. Most of the acid originally placed in the solution is HA molecules at equilibrium. That is, a weak acid ionizes only to a very small extent in aqueous solution, and so exists primarily as non-ionized molecules (**Figure 1(b)**).

strong acid an acid that ionizes almost 100 % in water, producing hydrogen ions

weak acid an acid that only partly ionizes in water, producing hydrogen ions

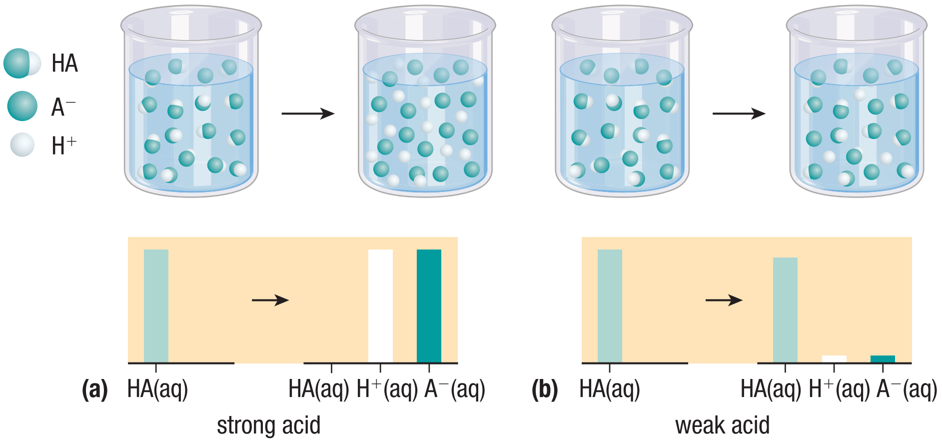


Figure 1 (a) A strong acid is almost completely ionized in water, resulting in relatively high concentrations of $H^+(aq)$ and $A^-(aq)$ ions and a much lower concentration of HA(aq) molecules. (b) A weak acid consists of a relatively high concentration of non-ionized HA(aq) molecules and much lower concentrations of $H^+(aq)$ and $A^-(aq)$ ions.

Table 1 Various Ways to Describe Acid Strength

Property	Strong acid	Weak acid
Value of acid ionization constant, K_a	K_a is large	K_a is small
Position of the ionization equilibrium	far to the right	far to the left
Equilibrium concentration of $H^+(aq)$ compared with original concentration of HA	$[H^+(aq)]_{equilibrium} \approx [HA(aq)]_{initial}$	$[H^+(aq)]_{equilibrium} << [HA(aq)]_{initial}$

As previously discussed, the acid ionization constant, K_a, is the equilibrium constant for the ionization of an acid. The general equation is

$$K_a = \frac{[H^+(aq)][A^-(aq)]}{[HA(aq)]}$$

where K_a always refers to the reaction of an acid, HA(aq), with water to form the conjugate base, A^-(aq), and the hydrogen ion, H^+(aq) (representing the hydronium ion, H_3O^+(aq)). Note that, since the concentration (density) of water is a constant, it is incorporated into the value of K_a. **Table 2** shows some values of K_a.

Table 2 Values of K_a for Some Common Acids

Acid formula	Name	Value of K_a
$HClO_4$(aq)	perchloric acid	very large
HNO_3(aq)	nitric acid	very large
HCl(aq)	hydrochloric acid	very large
HSO_4^-(aq)	hydrogen sulfate ion	1.2×10^{-2}
$HClO_2$(aq)	chlorous acid	1.2×10^{-2}
HF(aq)	hydrofluoric acid	6.6×10^{-4}
HNO_2(aq)	nitrous acid	4.6×10^{-4}
$HC_2H_3O_2$(aq)	ethanoic acid	1.8×10^{-5}
HClO(aq)	hypochlorous acid	3.5×10^{-8}
HCN(aq)	hydrocyanic acid	6.2×10^{-10}
NH_4^+(aq)	ammonium ion	5.8×10^{-10}
HCO_3^-(aq)	hydrogen carbonate ion	4.7×10^{-11}

increasing acid strength

There is an important connection between the strength of an acid and the strength of its conjugate base. The stronger an acid, the weaker its conjugate base, and conversely, the weaker an acid, the stronger its conjugate base (**Figure 2**).

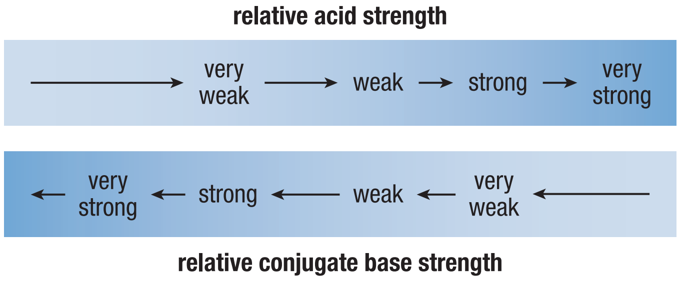


Figure 2 The relationship of acid strength and conjugate base strength for the reaction $HA(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + A^-(aq)$

Oxyacids and Organic Acids

oxyacid an acid in which the acidic hydrogen atom is attached to an oxygen atom

Most familiar acids are **oxyacids**, which have the acidic hydrogen (ionizable hydrogen) atom attached to an oxygen atom. Sulfuric acid is a typical example of a strong oxyacid. Many common weak acids, such as phosphoric acid, H_3PO_4(aq); nitrous acid, HNO_2(aq); and hypochlorous acid, HClO(aq), are also oxyacids.

You learned about organic acids—also called carboxylic acids—in Unit 1. An **organic acid** has a carbon backbone and a carboxyl group (**Figure 3**). Most organic acids are weak acids. Examples are ethanoic acid, $HC_2H_3O_2(aq)$, and benzoic acid, $HC_7H_5O_2(aq)$. Note that the acidic hydrogen atom is written at the beginning of the chemical formula. The remaining hydrogen atoms are not acidic—they do not form $H^+(aq)$ in water. This varies from how you wrote formulas for organic acids in Unit 1.

There are some important acids in which the acidic hydrogen atom is attached to an atom other than oxygen. The most significant of these are the acids of the halogens (specifically, $HF(aq)$, $HCl(aq)$, $HBr(aq)$, and $HI(aq)$).

organic acid an acid (except carbonic acid, $H_2CO_3(aq)$) containing carbon, oxygen, and hydrogen atoms; also called carboxylic acid

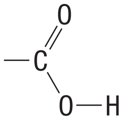


Figure 3 Organic acids (carboxylic acids) contain the carboxyl group. The hydrogen atom attached to oxygen is the ionizable hydrogen.

Strong Bases and Weak Bases

Like acids, bases may be either strong or weak, depending on the position of their equilibrium in solution. A strong base forms an equilibrium that lies farther to the right (toward products) when it reacts with water. A weak base forms an equilibrium that lies farther to the left (toward reactants) when it reacts with water.

Strong Bases

The bases sodium hydroxide, NaOH, and potassium hydroxide, KOH, dissociate completely in aqueous solution to form cations and hydroxide ions, leaving virtually no undissociated base entities in the solution.

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

$$KOH(s) \rightarrow K^+(aq) + OH^-(aq)$$

Thus, a 1.0 mol/L sodium hydroxide solution really contains 1.0 mol/L $Na^+(aq)$ and 1.0 mol/L $OH^-(aq)$. Since it dissociates completely, sodium hydroxide is called a **strong base**: a base that dissociates completely in aqueous solution.

strong base a base that dissociates completely in water, producing hydroxide ions

All the hydroxides of the Group 1 elements—LiOH, NaOH, KOH, RbOH, and CsOH—are strong bases. The Group 2 (alkaline earth) hydroxides—$Ca(OH)_2$, $Ba(OH)_2$, and $Sr(OH)_2$—are also strong bases. For the Group 2 bases, 2 mol of hydroxide ions are produced for every 1 mol of metal hydroxide dissolved:

$$\underset{1\text{ mol}}{Ca(OH)_2(aq)} \rightarrow \underset{1\text{ mol}}{Ca^{2+}(aq)} + \underset{2\text{ mol}}{2\,OH^-(aq)}$$

Although the alkaline earth hydroxides are strong bases, they are only slightly soluble. The low solubility of these bases can sometimes be an advantage. For example, many antacid medicines are suspensions of metal hydroxides, such as aluminum hydroxide, $Al(OH)_3(s)$, and magnesium hydroxide, $Mg(OH)_2(s)$. The low solubility of these compounds prevents a quick dissociation, which would release a high concentration of hydroxide ions that could harm the tissues lining the mouth, esophagus, and stomach. Although the hydroxide compounds in the antacid do not dissolve well in saliva, they readily dissolve in the highly acidic solution of the stomach. This is because, as dissolved $OH^-(aq)$ ions react with $H^+(aq)$ ions in stomach acid, the dissociation equilibrium position shifts to the right and more base dissociates.

Calcium hydroxide, $Ca(OH)_2$, often called slaked lime, is used in "scrubbers" to remove sulfur dioxide from the exhaust of power plants and refineries. Sulfur dioxide, if released into the air, may react with atmospheric moisture. This reaction produces sulfurous acid, leading to acid precipitation. In the scrubbing process, a suspension of slaked lime is sprayed into the stack gases to react with sulfur dioxide gas according to the following reaction steps:

$$SO_2(g) + H_2O(l) \rightleftharpoons H_2SO_3(aq)$$

$$Ca(OH)_2(aq) + H_2SO_3(aq) \rightleftharpoons CaSO_3(s) + 2\,H_2O(l)$$

The final products of these reactions—calcium sulfite and water—are less harmful to the environment than is sulfur dioxide.

Weak Bases

Many compounds are bases even though they do not contain the hydroxide ion. These compounds increase the concentration of hydroxide ions in aqueous solution because of their reaction with water. In other words, these bases are Brønsted–Lowry bases. For example, according to the Brønsted–Lowry theory, ammonia, $NH_3(aq)$, is a base because it reacts with water to form aqueous hydroxide ions:

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

In this reaction, water is a Brønsted–Lowry acid and ammonia is a Brønsted–Lowry base. Note that even though ammonia contains no hydroxide ions, it still increases the concentration of hydroxide ions in solution because of its reaction with water. Since the equilibrium position of this reaction is far to the left, ammonia is considered to be a **weak base**. Compounds that react with water as ammonia does are generally weak bases.

weak base a base that only partially reacts with water to produce hydroxide ions

Bases such as ammonia have at least one unshared pair of electrons that is capable of forming a coordinate covalent bond with a hydrogen ion. **Figure 4** shows the reaction of an ammonia molecule with a water molecule.

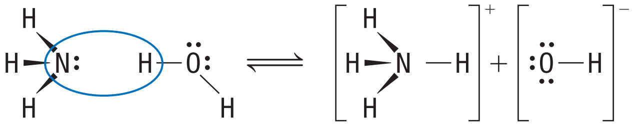
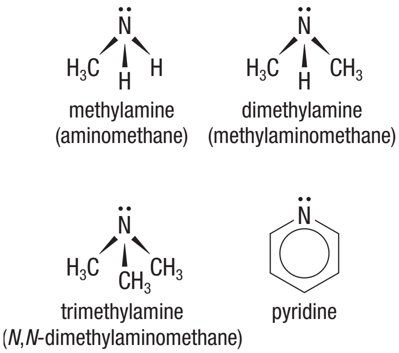


Figure 4 Ammonia and water react to form ammonium ions and hydroxide ions.

There are many bases that, like ammonia, produce hydroxide ions by reacting with water (**Figure 5**). In most of these bases, the lone pair of electrons is located on a nitrogen atom.

Recall the general equation for the reaction of a base, B, with water:

$$\underset{\text{base}}{B(aq)} + \underset{\text{acid}}{H_2O(l)} \rightleftharpoons \underset{\substack{\text{conjugate}\\ \text{acid}}}{BH^+(aq)} + \underset{\substack{\text{conjugate}\\ \text{base}}}{OH^-(aq)}$$



methylamine (aminomethane)

dimethylamine (methylaminomethane)

trimethylamine (*N*,*N*-dimethylaminomethane)

pyridine

Figure 5 Some examples of bases that do not contain hydroxide ions in their structures but produce hydroxide ions in aqueous solution by reaction with water

Brønsted-Lowry bases react with water to produce $OH^-(aq)$ ions and a conjugate acid, which together determine the acid–base properties of the aqueous solution. Weak bases, like weak acids, form dynamic equilibria in aqueous solutions. We can write an equilibrium constant for the ionization of a base.

For the reaction of a generic base with water, the equilibrium law equation, K, is written as follows:

$$K = \frac{[OH^-(aq)][BH^+(aq)]}{[B(aq)][H_2O(l)]}$$

However, since the concentration (density) of water is a constant, it can be incorporated into the value of K (just as it was in the equilibrium law equation for K_a). This yields a new constant, K_b, called the **base ionization constant**:

base ionization constant (K_b) the equilibrium constant for the ionization of a base; also called the base dissociation constant

$$K_b = \frac{[BH^+(aq)][OH^-(aq)]}{[B(aq)]}$$

For example, consider the ionization of ammonia in water. When ammonia reacts with water the equilibrium equation is as follows:

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

The K_b equation for this reaction is

$$K_b = \frac{[OH^-(aq)][NH_4^+(aq)]}{[NH_3(aq)]}$$

The equilibrium position of the reaction between ammonia and water lies far to the left, as is the case with all weak bases. Thus, the K_b values of ammonia and other weak bases tend to be small (for example, for ammonia, $K_b = 1.8 \times 10^{-5}$). **Table 3** lists some common bases and their K_b values. Weak bases are the conjugate bases of weak acids. For example, the ethanoate ion, $C_2H_3O_2^-(aq)$, is the conjugate base of ethanoic acid, $HC_2H_3O_2(aq)$. Similarly, the hypochlorite ion, $ClO^-(aq)$, is the conjugate base for hypochlorous acid, $HClO(aq)$.

UNIT TASK BOOKMARK

What is the K_a or K_b value of the consumer product that you have chosen for the Unit Task, outlined on page 582?

Table 3 K_b Values of Selected Weak Bases at 25 °C

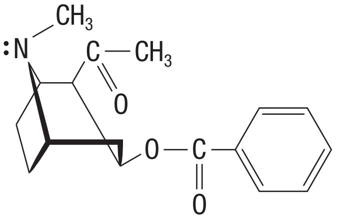
Name of base	Formula	K_b
dimethylamine	$(CH_3)_2NH(aq)$	9.6×10^{-4}
butylamine	$C_4H_9NH_2(aq)$	5.9×10^{-4}
methylamine	$CH_3NH_2(aq)$	4.4×10^{-4}
aniline	$C_6H_5NH_2(aq)$	4.1×10^{-10}
ammonia	$NH_3(aq)$	1.8×10^{-5}
hydrazine	$N_2H_4(aq)$	1.7×10^{-6}
morphine	$C_{17}H_{19}NO_3(aq)$	7.5×10^{-7}
hypochlorite ion	$ClO^-(aq)$	3.45×10^{-7}
pyridine	$C_5H_5N(aq)$	1.7×10^{-9}
ethanoate ion	$C_2H_3O_2^-(aq)$	5.6×10^{-10}
urea	$NH_2CONH_2(aq)$	1.5×10^{-14}

ORGANIC BASES

An organic compound that increases the concentration of hydroxide ions in aqueous solution is called an **organic base**. All organic bases contain carbon atoms and many also contain nitrogen atoms. One group of organic bases is called the alkaloids. Most alkaloids are derived from plants, fungi, and bacteria. Many drugs are based on alkaloids. These drugs include powerful painkillers such as codeine and morphine, and illicit drugs such as cocaine. Caffeine and nicotine are also alkaloids.

All of these compounds are weak organic bases because they contain at least one nitrogen atom with an unbonded pair of electrons that can accept a hydrogen ion, $H^+(aq)$, from water, leaving behind a hydroxide ion that makes the solution more basic. The diagram of cocaine in **Figure 6** shows the pair of electrons on the nitrogen atom.

organic base an organic compound that increases the concentration of hydroxide ions in aqueous solution



cocaine

Figure 6 Cocaine is an addictive alkaloid.

Water as an Acid and a Base

Recall from Section 8.1 that water is the most common amphiprotic substance: it can behave as either an acid or a base. Strangely, water can behave as both an acid and a base in the same reaction. This reaction is called the **autoionization of water** and involves the transfer of a hydrogen ion from one water molecule to another water molecule. The products are a hydroxide ion and a hydronium ion (**Figure 7**).

autoionization of water the transfer of a hydrogen ion from one water molecule to another

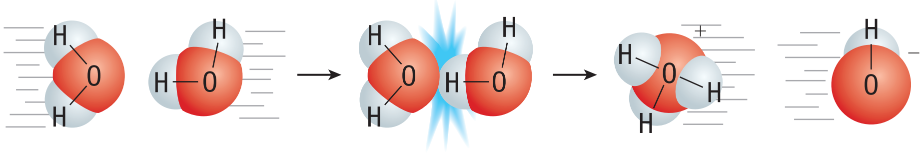


Figure 7 Two water molecules must collide in exactly the right orientation to form a hydronium ion and a hydroxide ion.

In this reaction, one water molecule acts as a Brønsted–Lowry acid by releasing a hydrogen ion, and the other acts as a Brønsted–Lowry base by accepting the hydrogen ion. We can write a chemical equation for the autoionization of water:

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

Since this is an equilibrium reaction, we can write an equilibrium law equation:

$$K = \frac{[H_3O^+(aq)][OH^-(aq)]}{[H_2O(l)]^2}$$

As always, we omit $[H_2O(l)]$, leaving the following simplified equation:

$$K = [H_3O^+(aq)][OH^-(aq)]$$

ion-product constant for water (K_w) the equilibrium constant for the autoionization of water

This constant is called the **ion-product constant for water, K_w**. We can also write K_w in an even simpler way if we use H^+ instead of H_3O^+:

$$K_w = [H^+(aq)][OH^-(aq)]$$

Experiments show that, at 25 °C in pure water,

$$[H^+(aq)] = 1.0 \times 10^{-7} \text{mol/L} \quad \text{and} \quad [OH^-(aq)] = 1.0 \times 10^{-7} \text{mol/L}$$

Therefore, we can calculate the value of K_w at 25 °C as follows:

$$\begin{aligned} K_w &= [H^+(aq)][OH^-(aq)] \\ &= (1.0 \times 10^{-7})(1.0 \times 10^{-7}) \\ K_w &= 1.0 \times 10^{-14} \end{aligned}$$

It is important to recognize the meaning of K_w. In any aqueous solution at 25 °C, no matter what the solution contains, the product of $[H^+(aq)]$ and $[OH^-(aq)]$ must always equal 1.0×10^{-14}. There are three possible situations:

- A neutral solution, where $[H^+(aq)] = [OH^-(aq)]$
- An acidic solution, where $[H^+(aq)] > [OH^-(aq)]$
- A basic solution, where $[OH^-(aq)] > [H^+(aq)]$

In each case, however, at 25 °C, $[H^+(aq)][OH^-(aq)] = 1.0 \times 10^{-14}$.

The Relationship between K_w, K_a, and K_b

The ionization reaction of a weak acid, HA(aq), is represented as

$$HA(aq) + H_2O(l) \rightleftharpoons A^-(aq) + H_3O^+(aq)$$

The acid ionization constant equation is

$$K_a = \frac{[H_3O^+(aq)][A^-(aq)]}{[HA(aq)]}$$

We know that A^-, the conjugate base of HA, is itself a base. As we did for other weak bases, we can write an ionization equation for the reaction of A^- with water:

$$A^-(aq) + H_2O(l) \rightleftharpoons HA(aq) + OH^-(aq)$$

And we can write the corresponding base ionization constant equation:

$$K_b = \frac{[HA][OH^-]}{[A^-]}$$

If we add together the ionization reactions for HA(aq) and $A^-(aq)$, as we did in Chapter 5 when solving Hess's law problems, we can obtain an overall equation:

$$\begin{array}{l} \cancel{HA(aq)} + H_2O(l) \rightleftharpoons \cancel{A^-(aq)} + H_3O^+(aq) \\ \underline{\cancel{A^-(aq)} + H_2O(l) \rightleftharpoons \cancel{HA(aq)} + OH^-(aq)} \\ \qquad 2\,H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq) \end{array}$$

Note that this overall reaction is exactly the same as the equation for the autoionization of water. Each of these equations represents an equilibrium with an associated equilibrium constant. For the ionization of the acid, HA(aq):

$$K_a = \frac{[H_3O^+(aq)][A^-(aq)]}{[HA(aq)]}$$

For the ionization of the conjugate base, $A^-(aq)$:

$$K_b = \frac{[HA(aq)][OH^-(aq)]}{[A^-(aq)]}$$

For the overall equation (autoionization of water):

$$K_w = [H_3O^+(aq)][OH^-(aq)]$$

Since there is a mathematical relationship between these three equations, there is also a mathematical relationship between their corresponding equilibrium constants, K_a, K_b, and K_w. If we multiply the K_a for the acid, HA(aq), by the K_b for its conjugate base, $A^-(aq)$, the product is K_w:

$$\begin{aligned} K_aK_b &= \frac{[H_3O^+(aq)]\cancel{[A^-(aq)]}}{\cancel{[HA(aq)]}} \times \frac{\cancel{[HA(aq)]}[OH^-(aq)]}{\cancel{[A^-(aq)]}} \\ &= [H_3O^+(aq)][OH^-(aq)] \\ K_aK_b &= K_w \end{aligned}$$

This relationship holds for all weak acids and bases: for a weak acid HA(aq) and its conjugate base $A^-(aq)$, or for a weak base B(aq) and its conjugate acid BH(aq),

$$\boxed{K_aK_b = K_w}$$

This relationship explains the trend we observed earlier: as the strength of the acid increases, the strength of its conjugate base decreases, and vice versa (**Figure 8**).

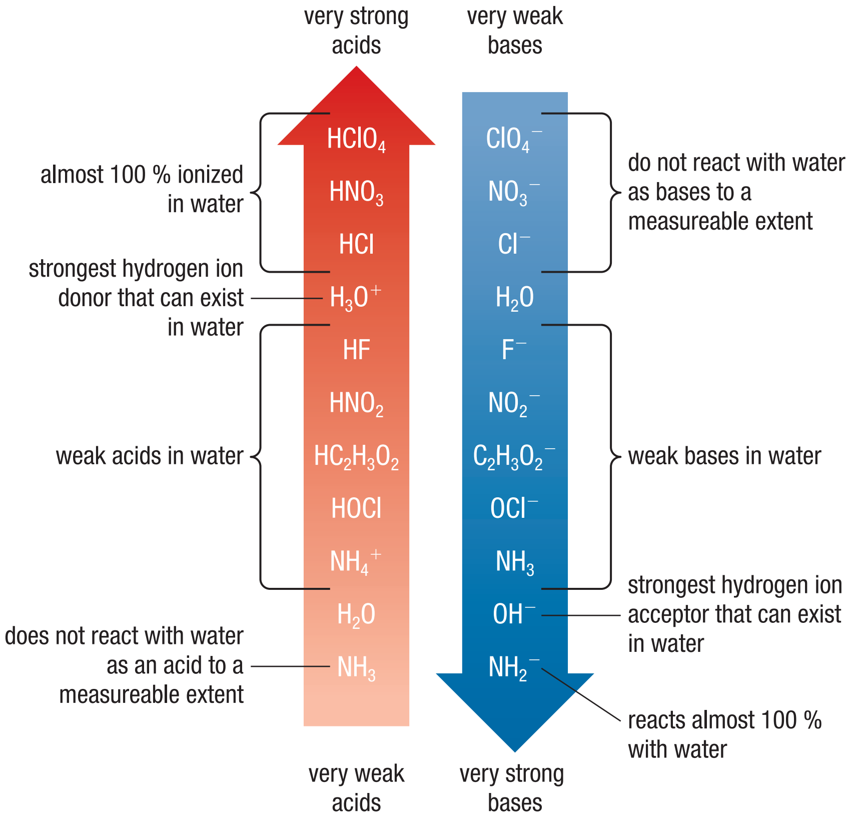


Figure 8 The relative strengths of acids and bases and their respective conjugate bases and conjugate acids. This illustrates which substances are likely to react with water.

Although a strong acid always has a very weak conjugate base, do not conclude that a weak acid has a strong conjugate base. For example, ethanoic acid is a weak acid ($K_a = 1.8 \times 10^{-5}$), but its conjugate base (ethanoate ion, $C_2H_3O_2^-(aq)$) is also a weak base ($K_b = 5.6 \times 10^{-10}$). In general, you can make the following assumptions:

- A strong acid or base has a very weak conjugate.
- A weak acid or base has a weak conjugate.
- A very weak acid or base has a strong conjugate.

Tutorial 1 Calculating K_b or K_a

You can use the equation $K_aK_b = K_w$ to calculate the value of K_a or K_b when given one or the other. For this reason, most tables of ionization constants list only K_a values for weak acids; the K_b values for their corresponding conjugate bases can be calculated.

Sample Problem 1: Calculating K_b from K_a

The hydrogen phosphate ion, $HPO_4^{2-}(aq)$, has a K_a of 1.3×10^{-13} at SATP. What is the base ionization constant, K_b, for the phosphate ion, $PO_4^{3-}(aq)$?

Given: K_a for $HPO_4^{2-}(aq) = 1.3 \times 10^{-13}$; $K_w = 1.0 \times 10^{-14}$

Required: K_b for $PO_4^{3-}(aq)$

Analysis: Use the equation $K_aK_b = K_w$ to solve for K_b, knowing that $K_w = 1.0 \times 10^{-14}$.

Solution:

$$\begin{aligned} K_aK_b &= K_w \\ K_b &= \frac{K_w}{K_a} \\ &= \frac{1.0 \times 10^{-14}}{4.2 \times 10^{-13}} \\ K_b &= 2.4 \times 10^{-2} \end{aligned}$$

Statement: The base ionization constant, K_b, for the phosphate ion is 2.4×10^{-2}.

Practice

1. The value of K_a for the ammonium ion, NH_4^+, is 5.8×10^{-10}. What is the value of the base ionization constant, K_b, for ammonia, NH_3, at SATP? K/U [ans: 1.7×10^{-5}]
2. K_b for the fluoride ion, F^-, is 1.5×10^{-11}. What is the value of K_a for HF at SATP? K/U [ans: 6.7×10^{-4}]

pH and pOH

Recall that, in pure water at 25 °C, the autoionization of water produces a hydrogen ion concentration of 1.0×10^{-7} mol/L and a hydroxide ion concentration of 1.0×10^{-7} mol/L. We may convert these very small concentration values into more convenient positive integer values by using logarithms. The negative logarithm of the hydrogen ion concentration is called **pH**. The negative logarithm of the hydroxide ion concentration is called **pOH**. Thus,

pH the negative logarithm of the concentration of hydrogen ions in an aqueous solution

pOH the negative logarithm of the concentration of hydroxide ions in an aqueous solution

$$pH = -\log[H^+(aq)] \quad \text{and} \quad pOH = -\log[OH^-(aq)]$$

Since pH is a logarithmic value based on 10, the pH changes by 1 for every 10-fold change in $[H^+(aq)]$. For example, a solution of pH 3 has a $H^+(aq)$ ion concentration 10 times greater than a solution of pH 4 and 100 times greater than a solution of pH 5. Also, note that since pH is defined as $-\log[H^+(aq)]$, pH decreases as $[H^+(aq)]$ increases and vice versa. The pH of common aqueous solutions at 25 °C ranges from 0 to 14. This range of pH values is called the pH scale and is shown in **Figure 9**.

We may use the pH equations to calculate the pH of pure water as follows:

$$\begin{aligned} [H^+(aq)] &= 1.0 \times 10^{-7}\ \text{mol/L} \\ pH &= -\log(1.0 \times 10^{-7}) \\ &= -(-7.00) \\ pH &= 7.00 \end{aligned}$$

We may use the pOH equation to calculate the pOH of pure water as follows:

$$
\begin{aligned}
[\mathrm{OH^-(aq)}] &= 1.0 \times 10^{-7}\ \mathrm{mol/L} \\
\mathrm{pOH} &= -\log(1.0 \times 10^{-7}) \\
&= -(-7.00) \\
\mathrm{pOH} &= 7.00
\end{aligned}
$$

In pure water, therefore, pH = 7 and pOH = 7.

This result does not only apply to pure (neutral) water, but to all neutral aqueous solutions. In all neutral aqueous solutions, both pH and pOH are equal to 7. Earlier you learned that, in pure water and all aqueous solutions, the product of $[\mathrm{H^+(aq)}]$ and $[\mathrm{OH^-(aq)}]$ always equals 1.0×10^{-14}, the value of K_w. For a solution to be neutral, the concentration of hydrogen ions must equal the concentration of hydroxide ions. These conditions can only be met if the concentration of hydrogen ions and hydroxide ions are both 1.0×10^{-7} mol/L. Thus, in pure water and all neutral aqueous solutions,

$$
\begin{aligned}
[\mathrm{H^+(aq)}] &= 1.0 \times 10^{-7}\ \mathrm{mol/L\ and\ pH} = 7 \\
[\mathrm{OH^-(aq)}] &= 1.0 \times 10^{-7}\ \mathrm{mol/L\ and\ pOH} = 7
\end{aligned}
$$

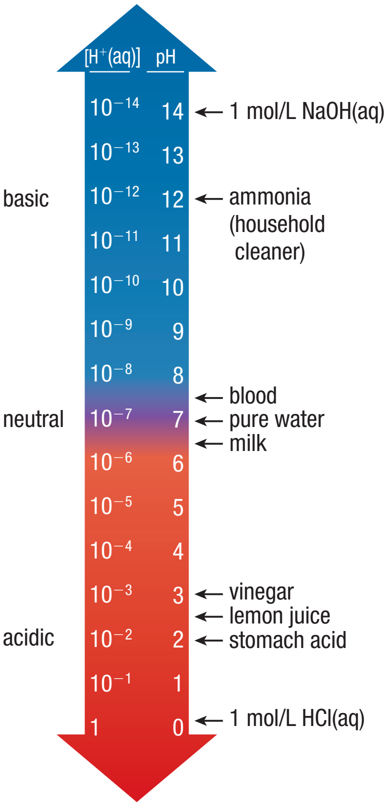


Figure 9 The pH scale showing hydrogen ion concentrations and the pH values of some common substances

pH and pOH of Acidic and Basic Solutions

Acidic and basic solutions are formed when acids and bases are dissolved in water. Recall that acids increase the concentration of $\mathrm{H^+(aq)}$ ions in solution and bases increase the concentration of $\mathrm{OH^-(aq)}$ ions in solution. If we add an acid to pure water, the $[\mathrm{H^+(aq)}]$ will increase to a value higher than 10^{-7} mol/L, and the pH will be lower than 7. For example, in a 0.010 mol/L HCl(aq) solution,

$$
\begin{aligned}
[\mathrm{H^+(aq)}] &= 1.0 \times 10^{-2}\ \mathrm{mol/L} \\
\mathrm{pH} &= -\log(1.0 \times 10^{-2}) \\
\mathrm{pH} &= 2
\end{aligned}
$$

Thus, when we dissolve an acid in water, there is an increase in $[\mathrm{H^+(aq)}]$ and a decrease in pH. However, while there is an increase in $[\mathrm{H^+(aq)}]$, there is also a proportional decrease in the concentration of hydroxide ions because, as mentioned earlier, in all aqueous solutions at 25 °C, $[\mathrm{H^+(aq)}][\mathrm{OH^-(aq)}] = 1.0 \times 10^{-14}$. Thus, since $[\mathrm{H^+(aq)}] = 1.0 \times 10^{-2}$ mol/L

and $[\mathrm{H^+(aq)}][\mathrm{OH^-(aq)}] = 1.0 \times 10^{-14}$

$$
\begin{aligned}
\text{then, } (1.0 \times 10^{-2})[\mathrm{OH^-(aq)}] &= 1.0 \times 10^{-14} \\
[\mathrm{OH^-(aq)}] &= \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-2}} \\
[\mathrm{OH^-(aq)}] &= 1.0 \times 10^{-12}\ \mathrm{mol/L}
\end{aligned}
$$

We may now calculate the pOH of this solution:

$$
\begin{aligned}
\mathrm{pOH} &= -\log(1.0 \times 10^{-12}) \\
\mathrm{pOH} &= 12
\end{aligned}
$$

Thus, for a 0.010 mol/L HCl(aq) solution,

$$
\begin{aligned}
[\mathrm{H^+(aq)}] &= 1.0 \times 10^{-2}\ \mathrm{mol/L\ and\ pH} = 2 \\
[\mathrm{OH^-(aq)}] &= 1.0 \times 10^{-12}\ \mathrm{mol/L\ and\ pOH} = 12
\end{aligned}
$$

As you can see in the example above, we may use the equation

$$[\mathrm{H^+(aq)}][\mathrm{OH^-(aq)}] = 1.0 \times 10^{-14}$$

to determine $[\mathrm{H^+(aq)}]$ or $[\mathrm{OH^-(aq)}]$ (and then pH and pOH) of any aqueous solution at 25 °C when the concentration of one ion or the other is known.

A useful equation in acid–base chemistry may be developed by using the negative logarithm of each component of the K_w equation, as follows:

$$[H^+(aq)][OH^-(aq)] = K_w$$

$$-\log([H^+(aq)][OH^-(aq)]) = -\log K_w$$

Since, mathematically, the logarithm of the product of two values is equal to the sum of the logarithms of the individual values, then

$$(-\log[H^+(aq)]) + (-\log[OH(aq)]) = -\log K_w$$
$$\text{pH} + \text{pOH} = -\log K_w$$

Since $K_w = 1.0 \times 10^{-14}$ and $-\log(1.0 \times 10^{-14}) = 14$ for all aqueous solutions at 25 °C, then pH + pOH = 14 for all aqueous solutions at 25 °C. This equation is useful because it allows us to calculate the pH or pOH of an aqueous solution at 25 °C if one or the other value is already known.

If, instead of adding an acid to pure water, we add a base, $[OH^-(aq)]$ will temporarily increase. As most of the hydroxide ions react with hydrogen ions, $[H^+(aq)]$ will decrease to a value below 1.0×10^{-7} mol/L and the pH will be greater than 7. At the same time, $[OH^-(aq)]$ will be higher than 1.0×10^{-7} mol/L and the pOH will be lower than 7. Using these relationships, we may state the following quantitative characteristics of solutions:

In neutral solutions,

- $[H^+(aq)] = 1.0 \times 10^{-7}$ mol/L and pH = 7
- $[OH^-(aq)] = 1.0 \times 10^{-7}$ mol/L and pOH = 7

In acidic solutions,

- $[H^+(aq)] > 1.0 \times 10^{-7}$ mol/L and pH < 7
- $[OH^-(aq)] < 1.0 \times 10^{-7}$ mol/L and pOH > 7

In basic solutions,

- $[H^+(aq)] < 1.0 \times 10^{-7}$ mol/L and pH > 7
- $[OH^-(aq)] > 1.0 \times 10^{-7}$ mol/L and pOH < 7

Tutorial 2 Calculating pH from pOH or pOH from pH

In this tutorial, you will learn to calculate pH or pOH when given the other value.

Sample Problem 1: Calculating pOH from pH

A sample of tap water has a measured pH value of 6.8 at 25 °C. What is its pOH?

Given: pH = 6.8

Required: pOH

Analysis: pH + pOH = 14.00

Solution:

$$\begin{aligned} \text{pH} + \text{pOH} &= 14.00 \\ \text{pOH} &= 14.00 - \text{pH} \\ &= 14.00 - 6.8 \\ \text{pOH} &= 7.2 \end{aligned}$$

Statement: The pOH of the solution is 7.2.

Sample Problem 2: Calculating pH from pOH

The pOH of a sample of human blood is 6.59 at 25 °C. What is its pH value?

Given: pOH = 6.59

Required: pH

Analysis: pH + pOH = 14.00

Solution:

$$
\begin{aligned}
\text{pH} + \text{pOH} &= 14.00 \\
\text{pH} &= 14.00 - \text{pOH} \\
&= 14.00 - 6.59 \\
\text{pOH} &= 7.41
\end{aligned}
$$

Statement: The pH of the blood sample is 7.41.

For guidance on using the correct number of significant digits in pH calculations, see Appendix A6.3.

Practice

1. Calculate the pOH of a solution that has a pH of 4. K/U [ans: 10]
2. Calculate the pH of a solution that has a pOH of 8.47. K/U [ans: 5.53]
3. Calculate the pH of a solution with a pOH of 2.41. Is this solution acidic, basic, or neutral? K/U [ans: 11.59; basic]

Measuring pH

Scientists in many different disciplines use pH meters to measure the pH of solutions. A **pH meter** is an electronic device with a probe that can be inserted into a solution of unknown pH. The probe contains an acidic aqueous solution enclosed by a special glass membrane that allows H^+(aq) ions to pass through. If the unknown solution has a different pH than the solution in the probe, the meter registers the resulting electric potential and displays the data as a pH reading (**Figure 10**). WEB LINK

pH meter a device that measures the acidity or alkalinity of a solution electronically and displays the result as a pH value

Figure 10 A pH meter shows the pH of a solution.

UNIT TASK **BOOKMARK**

Is the consumer product that you have chosen for the Unit Task (page 582) a strong or weak acid or base? What is its pH? What impact does this have on its effectiveness?

acid–base indicator a substance that changes colour within a specific pH range

Although a pH meter gives a very accurate measurement of a solution's pH, pH meters are not always available or practical. Another common way to determine the pH of a solution is to use an acid–base indicator. An **acid–base indicator** is a substance that has different colours in solutions with different pH values. Since the colour of an acid–base indicator varies with the pH of the solution, you can use an indicator to determine the approximate pH of a solution. For example, juice from red cabbage can range in colour from red to brown, depending on the pH of the solution with which it is mixed (**Figure 11**). Many plants produce naturally coloured substances that are acid–base indicators. For example, tea, red grape juice, and blueberries all change colour with pH. CAREER LINK

Litmus paper is another widely used acid–base indicator. It is a common indicator because it is readily available, inexpensive, and stores well. The dye used in litmus paper comes from lichen. After the water-soluble dye compound is extracted from the lichen, absorbent paper is soaked in the solution. When the paper dries, the litmus indicator is bonded to the paper. There are two types of litmus paper: blue and red. Acidic solutions turn blue litmus red; basic solutions turn red litmus blue. A neutral solution leaves red litmus red and blue litmus blue.

Figure 11 These beakers contain solutions with pH values ranging from 1 on the far left to 13 on the far right, along with a little boiled red cabbage and its juice.

Relating pH or pOH, and Ion Concentration

You may use acid–base indicators or a pH meter to determine the pH of a solution in an investigation. However, at times, you may need to determine the hydrogen or hydroxide ion concentration of a solution from the pH or pOH. Conversely, you may sometimes know the hydrogen ion or hydroxide ion concentration of a solution and need to determine the pH or pOH values.

The following equations allow you to calculate pH from $[H^+(aq)]$ and $[H^+(aq)]$ from pH:

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

Similarly, the following equations allow you to calculate pOH from $[OH^-(aq)]$ and $[OH^-(aq)]$ from pOH:

$$pOH = -\log[OH^-(aq)]$$
$$10^{-pOH} = [OH^-(aq)]$$

In the following tutorial, you will practice converting between pH and $[H^+(aq)]$ and between pOH and $[OH^-(aq)]$.

Tutorial 3 Calculating pH, pOH, $[H^+(aq)]$, or $[OH^-(aq)]$

You will often be given the concentration of an acid or base in solution and need to determine the pH of the solution from the concentration of hydrogen ions. You may also need to calculate pOH if you know the concentration of hydroxide ions. In this tutorial, you will convert between these different values.

Sample Problem 1: Calculating pH from $[H^+(aq)]$

A solution of NaOH has a $[H^+(aq)]$ of 5.2×10^{-11} mol/L. What is the pH of the solution?

Given: $[H^+(aq)] = 5.2 \times 10^{-11}$ mol/L

Required: pH

Analysis: $pH = -\log[H^+(aq)]$

Solution:

SKILLS HANDBOOK A6.2, A6.3

$$\begin{aligned}\text{pH} &= -\log[\text{H}^+(\text{aq})]\\ &= -\log(5.2 \times 10^{-11})\\ \text{pH} &= 10.284\end{aligned}$$

At this point you need to consider significant figures for logarithms. The number of decimal places in the log value must equal the number of significant figures in the original value.

Since there are 2 significant figures in the value of $[\text{H}^+(\text{aq})]$,

$$\begin{aligned}&\downarrow\\ [\text{H}^+(\text{aq})] &= 5.2 \times 10^{-11}\\ \text{pH} &= 10.28\\ &\uparrow\end{aligned}$$

the pH value should be rounded to 2 digits after the decimal.

$\text{pH} = 10.28$

Statement: The solution has a pH of 10.28.

Sample Problem 2: Calculating $[\text{H}^+(\text{aq})]$ from pH

A solution of ethanoic acid has a pH of 5.30. What is the concentration of hydrogen ions in the solution in mol/L?

Given: $\text{pH} = 5.30$

Required: $[\text{H}^+(\text{aq})]$

Analysis: $[\text{H}^+(\text{aq})] = 10^{-\text{pH}}$

Solution:

$$\begin{aligned}[\text{H}^+(\text{aq})] &= 10^{-\text{pH}}\\ &= 10^{-(5.30)} \text{ (2 digits following the decimal point)}\\ [\text{H}^+(\text{aq})] &= 5.0 \times 10^{-6} \text{ (2 significant digits)}\end{aligned}$$

Statement: The solution has a hydrogen ion concentration of 5.0×10^{-6} mol/L.

Sample Problem 3: Calculating pOH from $[\text{OH}^-(\text{aq})]$

A solution of ammonia has an $[\text{OH}^-(\text{aq})]$ of 8.2×10^{-12} mol/L. What is the pOH of the solution?

Given: $[\text{OH}^-(\text{aq})] = 8.2 \times 10^{-12}$ mol/L

Required: pOH

Analysis: $\text{pOH} = -\log[\text{OH}^-(\text{aq})]$

Solution:

$$\begin{aligned}\text{pOH} &= -\log[\text{OH}^-(\text{aq})]\\ &= -\log(8.2 \times 10^{-12}) \text{ (2 significant digits)}\end{aligned}$$

$\text{pOH} = 11.09$ (2 digits following the decimal point)

Statement: The solution has a pH of 11.09.

Sample Problem 4: Calculating pH and pOH from $[\text{H}^+(\text{aq})]$

A solution of sulfuric acid has a hydrogen ion concentration of 1.0×10^{-3} mol/L at 25 °C. Calculate the pH and pOH of this solution.

Given: $[\text{H}^+(\text{aq})] = 1.0 \times 10^{-3}$ mol/L

Required: pH, pOH

Analysis: $\text{pH} = -\log[\text{H}^+(\text{aq})]$; $K_\text{w} = [\text{H}^+(\text{aq})][\text{OH}^-(\text{aq})]$; $\text{pOH} = -\log[\text{OH}^-(\text{aq})]$

Solution: First, calculate the pH of the solution:

$$\text{pH} = -\log[\text{H}^+(\text{aq})]$$
$$= -\log(1.0 \times 10^{-3}) \text{ (2 significant digits)}$$
$$\text{pH} = 3.00 \text{ (2 digits following the decimal point)}$$

Next, use the equation for K_w to calculate the concentration of hydroxide ions:

$$K_w = [\text{H}^+(\text{aq})][\text{OH}^-(\text{aq})]$$
$$[\text{OH}^-(\text{aq})] = \frac{K_w}{[\text{H}^+(\text{aq})]}$$
$$= \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-3}}$$
$$[\text{OH}^-(\text{aq})] = 1.0 \times 10^{-11} \text{ mol/L}$$

Finally, calculate the pOH of the solution:

$$\text{pOH} = -\log[\text{OH}^-(\text{aq})]$$
$$= -\log(1.0 \times 10^{-11}) \text{ (2 significant digits)}$$
$$\text{pOH} = 11.00 \text{ (2 digits following the decimal point)}$$

Statement: The solution has a pH of 3.00 and a pOH of 11.00.

There is an alternative method of solving Sample Problem 4. Once you know the pH of the solution, you can use the equation pH + pOH = 14 to find pOH.

Practice

1. Calculate the pH of a solution with a $[\text{H}^+(\text{aq})]$ of 1.8×10^{-9} mol/L. K/U [ans: 8.74]
2. Calculate the $[\text{H}^+(\text{aq})]$ in a solution that has a pH of 2.5. K/U [ans: 3.2×10^{-3} mol/L]
3. Calculate the pOH of a solution with an $[\text{OH}^-(\text{aq})]$ of 3.2×10^{-4} mol/L. K/U [ans: 3.49]
4. Calculate the pH and pOH of a solution in which $[\text{OH}^-(\text{aq})] = 1.0$ mol/L. K/U [ans: pOH = 0.00; pH = 14.00]

Research This

Hazardous to Your Teeth

Skills: Researching, Analyzing, Evaluating, Communicating

SKILLS HANDBOOK A5

You should eat and drink a variety of foods and beverages to get all the nutrients you need in your diet. Unfortunately, some common foods and beverages may have adverse health effects on your teeth because something in those foods promotes the erosion of tooth enamel.

1. Conduct research and identify a food or beverage item that is associated with the erosion of tooth enamel.
2. Research the chemical process of tooth erosion.
3. Research how tooth erosion can be prevented.

A. What food or beverage item did you identify and what is its typical pH? K/U

B. Explain the chemical process of tooth erosion and what can be done to prevent it. T/I

C. Weigh the costs and benefits to overall health if a person were to stop consuming the food or beverage you identified. Would you recommend any change regarding consuming this food or beverage? A

D. Create a brief report to educate others about your findings and recommendation. Your report could be in any appropriate format of your choice. C A

WEB LINK

8.2 Review

Summary

- A strong acid or base completely ionizes in water. Its K_a or K_b value is very large.
- A weak acid or base ionizes only slightly in water. Its K_a or K_b value is relatively small.
- Water undergoes autoionization. The ion-product constant for the autoionization of water, K_w, is related to the concentration of hydrogen and hydroxide ions by the equation $K_w = [H^+(aq)][OH^-(aq)]$.
- K_w, K_a, and K_b are related by the equation $K_w = K_aK_b$.
- pH can be determined from the hydrogen ion concentration of a solution using the equation $pH = -\log[H^+(aq)]$. Similarly, $pOH = -\log[OH^-(aq)]$.
- pH and pOH are related by the equation $pH + pOH = 14$.
- The pH scale is used to describe the acidity or alkalinity of a solution.
- pH meters and acid–base indicators can be used to measure the pH of a solution.

Questions

1. Arrange the following sets of entities according to increasing strength as acids: H_2O, H_3O^+, HCl, $HC_2H_3O_2$, H_2S K/U
2. Arrange the following entities according to increasing strength as bases: F^-, Cl^-, NO_2^-, and CN^-. (Refer to Table 2 on page 496.) K/U
3. Calculate the K_b for the conjugate base of each of the following acids at SATP: K/U
   (a) methanoic acid, $HCHO_2(aq)$, $K_a = 1.8 \times 10^{-4}$
   (b) carbonic acid, $H_2CO_3(aq)$, $K_a = 4.4 \times 10^{-7}$
   (c) hypochlorous acid, $HClO(aq)$, $K_a = 3.5 \times 10^{-8}$
   (d) boric acid, $H_3BO_3(aq)$, $K_a = 5.8 \times 10^{-10}$
4. Calculate the K_a for the conjugate acid of the following bases at SATP: K/U
   (a) pyridine, $C_5H_5N(aq)$, $K_b = 1.7 \times 10^{-9}$
   (b) hydrazine, $N_2H_4(aq)$, $K_b = 1.7 \times 10^{-6}$
   (c) morphine, $C_{17}H_{19}NO_3(aq)$, $K_b = 7.5 \times 10^{-7}$
   (d) dimethylamine, $(CH_3)_2NH(aq)$, $K_b = 9.6 \times 10^{-4}$
5. Use the information given to state whether each solution is acidic or basic and to calculate its unknown $OH^-(aq)$ or $H^+(aq)$ concentration. K/U
   (a) 1.0×10^{-5} mol/L $OH^-(aq)$
   (b) 10.0 mol/L $H^+(aq)$
   (c) 5.0×10^{-7} mol/L $H^+(aq)$
   (d) 7.5×10^{-11} mol/L $OH^-(aq)$
   (e) 7.1×10^{-14} mol/L $H^+(aq)$
   (f) 1.2×10^{-6} mol/L $OH^-(aq)$
6. Determine the hydrogen and hydroxide ion concentrations in the following products: K/U
   (a) vinegar with a pH of 3.23
   (b) oven cleaner with a pH of 13.42
   (c) toilet bowl cleaner with a pH of 2.22
   (d) baking soda with a pH of 8.95
7. Calculate the pH of each of the following solutions at SATP: K/U
   (a) $[H^+(aq)] = 6.2 \times 10^{-11}$ mol/L
   (b) $[OH^-(aq)] = 7.1 \times 10^{-14}$ mol/L
   (c) $pOH = 4.98$
8. Calculate the pOH of each of the following solutions at SATP: K/U
   (a) $[OH^-(aq)] = 3.1 \times 10^{-4}$ mol/L
   (b) $[H^+(aq)] = 1.0 \times 10^{-7}$ mol/L
   (c) $pH = 3.84$
9. Calculate $[H^+(aq)]$ and $[OH^-(aq)]$ for each of the following solutions: K/U
   (a) $pH = 12.00$
   (b) $pH = 1.54$
   (c) $pOH = 11.00$
   (d) $pOH = 4.69$
10. Comment on this statement: "A strong acid is more dangerous than a weak acid." C A
11. While looking up acids on the Internet, a student came across the term "superacid." Research superacids. Write a short report outlining their production, uses, and hazards. C A

WEB LINK