16.8: Finding the $[\mathrm{OH}^-]$ and pH of Strong and Weak Base Solutions

Like acids, bases can be strong or weak. Most strong bases are ionic compounds containing the OH^- ion. Most weak bases act as a base by accepting a proton from water. We examine each separately.

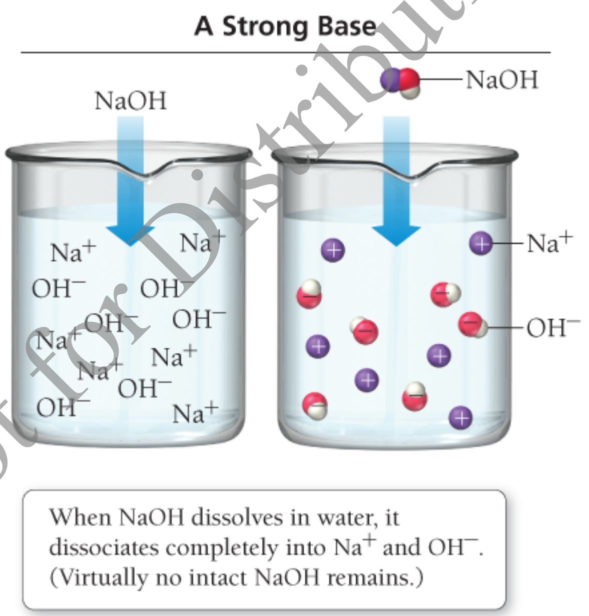
Strong Bases

Just as we define a strong acid as one that completely ionizes in solution, analogously we define a *strong base* as one that completely dissociates in solution. NaOH, for example, is a strong base:

$$\mathrm{NaOH}\,(aq) \rightarrow \mathrm{Na}^+\,(aq) + \mathrm{OH}^-\,(aq)$$

An NaOH solution contains no intact NaOH—it has all dissociated to form $\mathrm{Na}^+\,(aq)$ and $\mathrm{OH}^-\,(aq)$ (Figure 16.9). In other words, a 1.0 M NaOH solution has $[\mathrm{OH}^-] = 1.0$ M and $[\mathrm{Na}^+] = 1.0$ M Table 16.7 lists the common strong bases.

Figure 16.9 Ionization of a Strong Base



When NaOH dissolves in water, it dissociates completely into Na^+ and OH^-. (Virtually no intact NaOH remains.)

Table 16.7 Common Strong Bases

Lithium hydroxide (LiOH)	Strontium hydroxide $[\mathrm{Sr(OH)_2}]$
Sodium hydroxide (NaOH)	Calcium hydroxide $[\mathrm{Ca(OH)_2}]$
Potassium hydroxide (KOH)	Barium hydroxide $[\mathrm{Ba(OH)_2}]$

As Table 16.7 illustrates, most strong bases are group 1A or group 2A metal hydroxides. The group 1A metal hydroxides are highly soluble in water and can form concentrated base solutions. The group 2A metal hydroxides, however, are only slightly soluble, a useful property for some applications. Notice that the general formula for the group 2A metal hydroxides is $M(OH)_2$. When they dissolve in water, they produce 2 mol of OH^- per mole of the base. For example, $Sr(OH)_2$ dissociates as follows:

$$Sr(OH)_2(aq) \rightarrow Sr^{2+}(aq) + 2\,OH^-(aq)$$

Unlike diprotic acids, which ionize in two steps, bases containing two OH^- ions dissociate in a single step.

Weak Bases

Recall that a *weak base* is analogous to a weak acid. Unlike strong bases that contain OH^- and *dissociate* in water, the most common weak bases produce OH^- by accepting a proton from water, ionizing water to form OH^- according to the general equation:

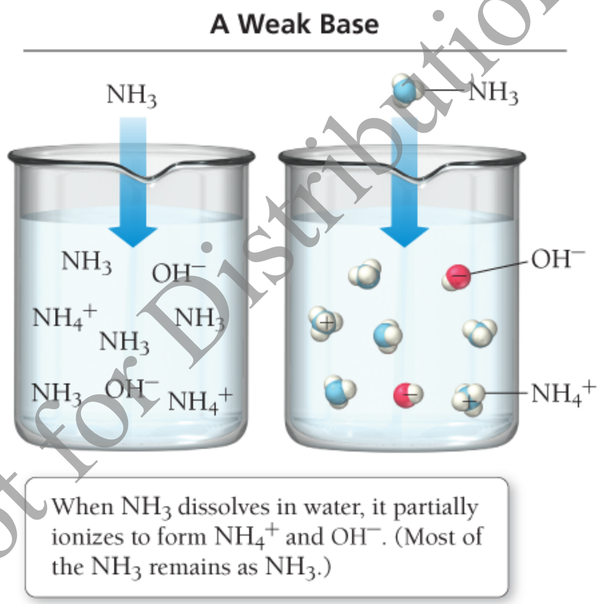
$$B(aq) + H_2O(l) \rightleftharpoons BH^+(aq) + OH^-(aq)$$

In this equation, B is a generic symbol for a weak base. Ammonia, for example, ionizes water as follows:

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

The double arrow indicates that the ionization is not complete. An NH_3 solution contains mostly NH_3 with only some NH_4^+ and OH^- (Figure 16.10). A 1.0 M NH_3 solution has $[OH^-]1.0\ M$

Figure 16.10 **Ionization of a Weak Base**



When NH_3 dissolves in water, it partially ionizes to form NH_4^+ and OH^-. (Most of the NH_3 remains as NH_3.)

We quantify the extent of ionization of a weak base with the **base ionization constant**, K_b. For the general reaction in which a weak base ionizes water, we define K_b as follows:

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

By analogy with K_a, the smaller the value of K_b, the weaker the base. Table 16.8 lists some common weak bases, their ionization reactions, and values for K_b. The "p" scale can also be applied to K_b, so that $pK_b = -\log K_b$

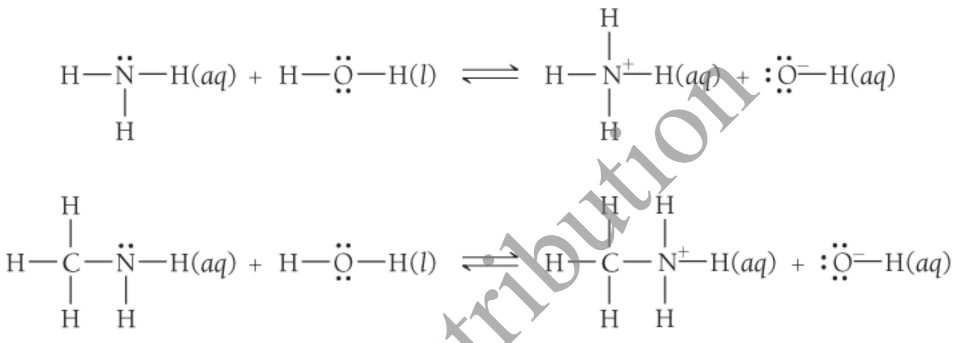
Table 16.8 **Some Common Weak Bases**

Weak Base	Ionization Reaction	K_b

Carbonate ion (CO_3^{2-})*	$CO_3^{2-}(aq) + H_2O(l) \rightleftharpoons HCO_3^-(aq) + OH^-(aq)$	1.8×10^{-4}
Methylamine (CH_3NH_2)	$CH_3NH_2(aq) + H_2O(l) \rightleftharpoons CH_3NH_3^+(aq) + OH^-(aq)$	4.4×10^{-4}
Ethylamine ($C_2H_5NH_2$)	$C_2H_5NH_2(aq) + H_2O(l) \rightleftharpoons C_2H_5NH_3^+(aq) + OH^-(aq)$	5.6×10^{-4}
Ammonia (NH_3)	$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$	1.76×10^{-5}
Bicarbonate ion (HCO_3^-)* (or hydrogen carbonate)	$HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq) + OH^-(aq)$	2.3×10^{-8}
Pyridine (C_5H_5N)	$C_5H_5N(aq) + H_2O(l) \rightleftharpoons C_5H_5NH^+(aq) + OH^-(aq)$	1.7×10^{-9}
Aniline ($C_6H_5NH_2$)	$C_6H_5NH_2(aq) + H_2O(l) \rightleftharpoons C_6H_5NH_3^+(aq) + OH^-(aq)$	3.9×10^{-10}

*The carbonate and bicarbonate ions must occur with a positively charged ion such as Na^+ that serves to balance the charge but does not have any part in the ionization reaction. For example, it is the bicarbonate ion that makes sodium bicarbonate ($NaHCO_3$) basic. We look more closely at ionic bases in Section 16.9.

All but two of the weak bases listed in Table 16.8 are either ammonia or *amines*, which we can think of as ammonia with one or more hydrocarbon groups substituted for one or more hydrogen atoms. All of these bases have a nitrogen atom with a lone pair (Figure 16.11). This lone pair acts as the proton acceptor that makes the substance a base, as shown in the reactions for ammonia and methylamine:

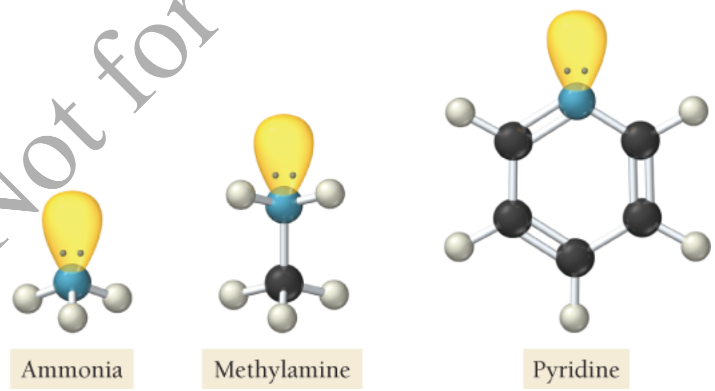


$$\mathrm{H-\underset{H}{\ddot{N}}-H}(aq) + \mathrm{H-\underset{\cdot\cdot}{\ddot{O}}-H}(l) \rightleftharpoons \mathrm{H-\overset{H}{\underset{H}{N^+}}-H}(aq) + :\underset{\cdot\cdot}{\ddot{O}}^{-}\mathrm{-H}(aq)$$

$$\mathrm{H-\overset{H}{\underset{H}{C}}-\underset{H}{\ddot{N}}-H}(aq) + \mathrm{H-\underset{\cdot\cdot}{\ddot{O}}-H}(l) \rightleftharpoons \mathrm{H-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{N^+}}-H}(aq) + :\underset{\cdot\cdot}{\ddot{O}}^{-}\mathrm{-H}(aq)$$

Figure 16.11 Lone Pairs in Weak Bases

Many weak bases have a nitrogen atom with a lone pair that acts as the proton acceptor.



Finding the $[OH^-]$ and pH of Basic Solutions

Finding the $[OH^-]$ and pH of a strong base solution is relatively straightforward, as shown in Example 16.11. As we did in calculating the $[H_3O^+]$ in strong acid solutions, we can neglect the contribution of the autoionization of water to the $[OH^-]$ and focus solely on the strong base itself.

Example 16.11 Finding the $[OH^-]$ and pH of a Strong Base Solution

What are the OH^- concentration and pH in each solution?

a. 0.225 M KOH

b. $0.0015\text{M}\ Sr(OH)_2$

SOLUTION

a. Since KOH is a strong base, it completely dissociates into K^+ and OH^- in solution. The concentration of OH^- is therefore the same as the given concentration of KOH.

Use this concentration and K_w to find $[H_3O^+]$.

Then substitute $[H_3O^+]$ into the pH expression to find the pH.

$$
\begin{aligned}
&KOH\,(aq) \rightarrow K^+\,(aq) + OH^-\,(aq)\\
&[OH^-] = 0.225\text{M}\\
&[H_3O^+]\,[OH^-] = K_w = 1.00 \times 10^{-14}\\
&[H_3O^+]\,(0.225) = 1.00 \times 10^{-14}\\
&[H_3O^+] = 4.44 \times 10^{-14}\text{M}\\
&\text{pH} = -\log[H_3O^+]\\
&\quad = -\log\left(4.44 \times 10^{-14}\right)\\
&\quad = 13.352
\end{aligned}
$$

b. Since $Sr(OH)_2$ is a strong base, it completely dissociates into 1 mol of Sr^{2+} and 2 mol of OH^- in solution. The concentration of OH^- is therefore twice the given concentration of $Sr(OH)_2$.

Use this concentration and K_w to find $[H_3O^+]$.

Substitute $[H_3O^+]$ into the pH expression to find the pH.

$$
\begin{aligned}
&Sr(OH)_2\,(aq) \rightarrow Sr^{2+}\,(aq) + 2\,OH^-\,(aq)\\
&[OH^-] = 2(0.0015)\text{M}\\
&\qquad = 0.0030\ \text{M}\\
&[H_3O^+]\,[OH^-] = K_w = 1.0 \times 10^{-14}\\
&[H_3O^+]\,(0.0030) = 1.0 \times 10^{-1}\\
&[H_3O^+] = 3.3 \times 10^{-12}\text{M}\\
&\text{pH} = -\log[H_3O^+]\\
&\quad = -\log\left(3.3 \times 10^{-12}\right)\\
&\quad = 11.48
\end{aligned}
$$

FOR PRACTICE 16.11 Find the $[OH^-]$ and pH of a 0.010 M $Ba(OH)_2$ solution.

Finding the $[OH^-]$ and pH of a *weak base* solution is analogous to finding the $[H_3O^+]$ and pH of a weak acid. Similarly, we can neglect the contribution of the autoionization of water to the $[OH^-]$ and focus solely on the weak base itself. We find the contribution of the weak base by preparing an ICE table showing the relevant concentrations of all species and then use the base ionization constant expression to find the $[OH^-]$. Example 16.12 demonstrates how to find the $[OH^-]$ and pH of a weak base solution.

Example 16.12 Finding the $[OH^-]$ and pH of a Weak Base Solution

Find the $[OH^-]$ and pH of a 0.100 M NH_3 solution.

SOLUTION

1. Write the balanced equation for the ionization of water by the base and use it as a guide to prepare an ICE table showing the given concentration of the weak base as its initial concentration. Leave room in the table for the changes in concentrations and for the equilibrium concentrations. (Note that you should list the OH^- concentration as approximately zero. Although a little OH^- is present from the autoionization of water, this amount is negligibly small compared to the amount of OH^- formed by the base.)

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

	$[NH_3]$	$[NH_4^+]$	$[OH^-]$
Initial	0.100	0.00	≈0.00
Change			
Equil			

2. Represent the change in the concentration of OH^- with the variable x. Define the changes in the concentrations of the other reactants and products in terms of x.

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

	$[NH_3]$	$[NH_4^+]$	$[OH^-]$
Initial	0.100	0.00	≈0.00
Change	$-x$	$+x$	$+x$
Equil			

3. Sum each column to determine the equilibrium concentrations in terms of the initial concentrations and the variable x.

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

	$[NH_3]$	$[NH_4^+]$	$[OH^-]$
Initial	0.100	0.00	≈0.00
Change	$-x$	$+x$	$+x$
Equil	$0.100 - x$	x	x

4. Substitute the expressions for the equilibrium concentrations (from Step 3) into the expression for the base ionization constant.
 In many cases, you can make the approximation that *x is small*.
 Substitute the value of the base ionization constant (from Table 16.8) into the K_b expression and solve for x.
 Confirm that the *x is small* approximation is valid by calculating the ratio of x to the number it was subtracted from in the approximation. The ratio should be less than 0.05 (or 5%).

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

$$= \frac{x^2}{0.100 - \cancel{x}} \text{ (x is small)}$$

$$1.76 \times 10^{-5} = \frac{x^2}{0.100}$$

$$\sqrt{1.76 \times 10^{-5}} = \sqrt{\frac{x^2}{0.100}}$$

$$x = \sqrt{(0.100)(1.76 \times 10^{-5})}$$

$$= 1.33 \times 10^{-3}$$

$$\frac{1.33 \times 10^{-3}}{0.100} \times 100\% = 1.33\%$$

Therefore the approximation is valid.

5. Determine the OH^- concentration from the calculated value of x.

Use the expression for K_w to find $[H_3O^+]$.

Substitute $[H_3O^+]$ into the pH equation to find pH.

$$\begin{aligned}
[OH^-] &= 1.33 \times 10^{-3} M \\
[H_3O^+][OH^-] &= K_w = 1.00 \times 10^{-14} \\
[H_3O^+](1.33 \times 10^{-3}) &= 1.00 \times 10^{-14} \\
[H_3O^+] &= 7.52 \times 10^{-12} M \\
pH &= -\log [H_3O^+] \\
&= -\log (7.52 \times 10^{-12}) \\
&= 11.124
\end{aligned}$$

FOR PRACTICE 16.12 Find the $[OH^-]$ and pH of a 0.33 M methylamine solution.

Interactive Worked Example 16.12 **Finding the [OH⁻] and pH of a Weak Base Solution**