The pH of a solution of a weak base can be calculated in a way which is very similar to that used for a weak acid. Instead of an acid constant $K_a$, a base constant $K_b$ must be used. If a weak base $B$ accepts protons from water according to the equation

$$\text{B} + \text{H}_2\text{O} \rightleftharpoons \text{BH}^+ + \text{OH}^- \label{1}$$

then the base constant is defined by the expression

$$K_b = \frac{[\text{BH}^+][\text{OH}^-]}{[\text{B}]} \label{2}$$

A list of $K_b$ values for selected bases arranged in order of strength is given in the table below. This table is part of our larger collection of acid-base resources.

**Table \(\PageIndex{1}\): The Base Constants for Some Bases at 25°C.**

<table>
<thead>
<tr>
<th>Base</th>
<th>Formula and Ionization Equation</th>
<th>$K_b$</th>
<th>Molecular Shape</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ammonia</td>
<td>$\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}^+_4 + \text{OH}^-$</td>
<td>$1.77 \times 10^{-5}$</td>
<td><img src="image" alt="Ammonia Molecular Shape" /></td>
</tr>
<tr>
<td>Aniline</td>
<td>$\text{C}_6\text{H}_5\text{NH}_2 + \text{H}_2\text{O} \rightleftharpoons \text{C}_6\text{H}_5\text{NH}^+_3 + \text{OH}^-$</td>
<td>$3.9 \times 10^{-10}$</td>
<td><img src="image" alt="Aniline Molecular Shape" /></td>
</tr>
<tr>
<td>Base</td>
<td>Formula and Ionization Equation</td>
<td>$K_b$</td>
<td>Molecular Shape</td>
</tr>
<tr>
<td>---------------</td>
<td>--------------------------------------------------------------------------------------------------</td>
<td>--------------------</td>
<td>-----------------</td>
</tr>
<tr>
<td>Carbonate ion</td>
<td>$(\text{CO}_3^{2–} + \text{H}_2\text{O}) \xrightarrow{\text{rightleftharpoons}} \text{HCO}^-_3 + \text{OH}^–$</td>
<td>$2.1 \times 10^{-4}$</td>
<td><img src="image1" alt="Molecule" /></td>
</tr>
<tr>
<td>Hydrazine</td>
<td>$(\text{N}_2\text{H}_4 + \text{H}_2\text{O}) \xrightarrow{\text{rightleftharpoons}} \text{N}_2\text{H}^+_5 + \text{OH}^–$</td>
<td>$K_1 = 1.2 \times 10^{-6}$&lt;br&gt; $K_2 = 1.3 \times 10^{-15}$</td>
<td><img src="image2" alt="Molecule" /></td>
</tr>
<tr>
<td>Base</td>
<td>Formula and Ionization Equation</td>
<td>$K_b$</td>
<td>Molecular Shape</td>
</tr>
<tr>
<td>--------------</td>
<td>---------------------------------------------------------------------</td>
<td>-----------------</td>
<td>-----------------</td>
</tr>
<tr>
<td>Hydride ion</td>
<td>$\text{(H}^- + \text{H}_2\text{O} \rightarrow \text{H}_2 + \text{OH}^-)}$</td>
<td>$\text{large}$</td>
<td></td>
</tr>
<tr>
<td>Phosphate ion</td>
<td>$\text{(PO}_4^{3-} + \text{H}_2\text{O} \rightleftharpoons \text{HPO}^{2-} + \text{OH}^-)}$</td>
<td>$5.9 \times 10^{-3}$</td>
<td></td>
</tr>
</tbody>
</table>

To find the pH we follow the same general procedure as in the case of a weak acid. If the stoichiometric concentration of the base is indicated by \(c_b\), the result is entirely analogous to equation 4 in the section on the pH of weak acids; namely,

\[
K_b = \frac{[\text{OH}^-]^2}{c_b - [\text{OH}^-]}
\]

Under most circumstances we can make the approximation

\[
c_b - [\text{OH}^-] \approx c_b
\]

in which case Equation \(\ref{3}\) reduces to the approximation

\[
[\text{OH}^-] \approx \sqrt{K_b c_b}
\]

which is identical to the expression obtained in the acid case (approximation shown in equation 6 in the section on the pH of weak acids) except that \(\text{OH}^-\) replaces \(\text{H}_3\text{O}^+\) and \(b\) replaces \(a\). Once we have found the hydroxide-ion concentration from this approximation, we can then easily find the pOH, and from it the pH.

Example \(\PageIndex{1}\): pH using \(K_b\)

Using the value for \(K_b\) listed in the table, find the pH of 0.100 \(M\) \(\text{NH}_3\).
Solution

It is not a bad idea to guess an approximate pH before embarking on the calculation. Since we have a dilute solution of a weak base, we expect the solution to be only mildly basic. A pH of 13 or 14 would be too basic, while a pH of 8 or 9 is too close to neutral. A pH of 10 or 11 seems reasonable. Using Eq. (4) we have

\[
[OH^-] = \sqrt{K_b \cdot c_b} = \sqrt{1.8 \times 10^{-5} \text{ mol L}^{-1} \cdot 0.100 \text{ mol L}^{-1}} = 1.34 \times 10^{-3} \text{ mol L}^{-1}
\]

Checking the accuracy of the approximation, we find

\[
\frac{[OH^-]}{c_b} = \frac{1.34 \times 10^{-3}}{0.100} \approx 1 \text{ percent}
\]

The approximation is valid, and we thus proceed to find the pOH.

\[
pOH = -\log\left(\frac{[OH^-]}{\text{mol L}^{-1}}\right) = -\log(1.34 \times 10^{-3}) = 2.87
\]

From which

\[
\text{pH} = 14.00 - \text{pOH} = 14.00 - 2.87 = 11.13
\]

This calculated value checks well with our initial guess.

Occasionally we will find that the approximation

\[
c_b - [OH^-] \approx c_b
\]

is not valid, in which case we must use a series of successive approximations similar to that outlined above for acids. The appropriate formula can be derived from Eq. (3) and reads

\[
[OH^-] = \sqrt{K_b \cdot (c_b - [OH^-])}
\]

Contributors

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