ACIDS AND BASES

THE NATURE OF ACIDS AND BASES

1. The Arrhenius concept of acids and bases describes acids as substances that produce hydrogen ions in aqueous solution and bases as substances that produce hydroxide ions in aqueous solution.

2. The concept is limited because it applies only to aqueous solutions and allows for only one kind of base – the hydroxide ion.

3. In the **BREWSTER - LOWRY MODEL** an acid is a proton (H\(^+\)) donor and a base is a proton acceptor.

\[
\begin{align*}
\text{H}^+ \quad &\text{H}^+ + \quad H - C_l ^- \quad \rightarrow \quad [H^+ \quad C_l ^-] + [C_l ^-] \\
\end{align*}
\]

4. The proton is transferred from the HCl molecule to the water molecule to form \(H_3O^+\), which is called the **HYDROGEN ION**.

5. The reaction that occurs when an acid is dissolved in water is best expressed as:

\[
\begin{align*}
\text{HA} (aq) + \quad H_2O (l) &\rightleftharpoons \quad H_3O^+ (aq) + \quad A^- (aq) \\
\text{ACID} &\quad \text{BASE} &\quad \text{CONJUGATE ACID} &\quad \text{CONJUGATE BASE} \\
\end{align*}
\]

6. The **CONJUGATE BASE** is everything that remains of the acid molecule after a proton is lost.

7. The **CONJUGATE ACID** is formed when the proton is transferred to the base.

8. A **CONJUGATE ACID-BASE PAIR** consists of two substances related to each other by the donating and accepting of a single proton.

9. There is really a competition for the proton between the two bases H\(_2\)O and A\(^-\).

10. If water is a much stronger base than A\(^-\), the equilibrium position will be far to the right; most of the acid dissolved will be in the ionized form. If A\(^-\) is a much larger base than water, the equilibrium position will lie far to the left. In this case most of the acid dissolved will be present at equilibrium as HA.

11. The equilibrium expression for:

\[
\begin{align*}
HA (aq) + \quad H_2O (l) &\rightleftharpoons \quad H_3O^+ (aq) + \quad A^- (aq) \\
\text{HA} (aq) &\quad \text{H}_2O (l) &\quad \text{H}_3O^+ (aq) &\quad \text{A}^- (aq) \\
\end{align*}
\]

is:

\[
K_a = \frac{[H_3O^+] [A^-]}{[HA]} \]

\[
K_a = \frac{[H^+] [A^-]}{[HA]} \]

12. $K_a$ is called the **ACID DISSOCIATION CONSTANT**

13. While water plays an important role in causing the acid to ionize, it is not included in the equilibrium expression.

14. $K_a$ is used to represent the equilibrium constant for the reaction in which a proton is removed from HA to form the conjugate base $A$.

**** Write the dissociation reaction and the corresponding $K_a$ equilibrium expression for each of the following acids in water:

a. $\text{HC}_2\text{H}_3\text{O}_2$

\[ \text{HC}_2\text{H}_3\text{O}_2(aq) \rightleftharpoons H^+(aq) + \text{C}_2\text{H}_3\text{O}_2^-(aq) \]

\[ K_a = \frac{[H^+] \cdot [\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]} \]

b. $\text{Co(H}_2\text{O)}_6^{3+}$

\[ \text{Co(H}_2\text{O)}_6^{2+}(aq) \rightleftharpoons H^+(aq) + \text{Co(H}_2\text{O)}_5\text{OH}^{3+}(aq) \]

\[ K_a = \frac{[H^+] \cdot [\text{Co(H}_2\text{O)}_5\text{OH}^{3+}]}{[\text{Co(H}_2\text{O)}_6^{2+}]} \]

c. $\text{CH}_3\text{NH}_3^+$

\[ \text{CH}_3\text{NH}_3^+(aq) \rightleftharpoons H^+(aq) + \text{CH}_3\text{NH}_2^- (aq) \]

\[ K_a = \frac{[H^+] \cdot [\text{CH}_3\text{NH}_2^-]}{[\text{CH}_3\text{NH}_3^+]} \]
15. The Bronsted-Lowry model is not limited to aqueous solutions; it can be extended to reactions in the gas phase:

\[ \text{HCl}_2 (g) + \text{HCl} (g) \rightleftharpoons \text{HCl}_3 (g) \]

\[ + \text{Cl}^- \quad \text{H}^+ \quad \text{Cl}^- \]

**ACID STRENGTH**

1. The strength of an acid is defined by the equilibrium position of its dissociation (ionization) reaction.

2. A **strong acid** is one for which this equilibrium lies far to the right.

3. A strong acid yields a weak conjugate base; one that has a low affinity for a proton. In fact, the conjugate base is a much weaker base than water.

4. A **weak acid** is one for which the equilibrium lies far to the left. Most of the acid placed in the solution is still present as HA at equilibrium.

5. A weak acid has a conjugate base that is a much stronger base than water. A weak acid yields a relatively strong conjugate base.

6. The common strong acids are:
   \[ \text{H}_2 \text{SO}_4 \quad \text{H}_2 \text{SO}_3 \quad \text{HNO}_3 \quad \text{HClO}_3 \]
   \[ \text{HCl} \quad \text{HBr} \quad \text{HClO}_4 \]

7. Sulfuric acid is a **diprotic acid**, an acid having two acidic protons.

8. The dissociation of sulfuric acid looks like:

   \[ \text{H}_2 \text{SO}_4 (aq) \rightarrow \text{H}^+(aq) + \text{HSO}_4^-(aq) \]

   This reaction is virtually 100% complete, thus \( \text{H}_2 \text{SO}_4 \) is a strong acid.

   The next step:

   \[ \text{HSO}_4^- (aq) \rightleftharpoons \text{H}^+(aq) + \text{SO}_4^{2-}(aq) \]

   Does not go far because \( \text{HSO}_4^- \) is a weak acid.
9. Most acids are \textit{oxyacids} in which the acidic proton is attached to an oxygen atom.

10. Many common weak acids are oxyacids.

11. Organic acids contain the \textit{carboxyl} group.

\begin{align*}
\text{C}_2\text{H}_3\text{O}_2^- & \quad \text{ACETIC ACID} \\
\text{C}_6\text{H}_5\text{COO}^- & \quad \text{BENZOIC ACID}
\end{align*}

12. Another important acid is one in which the acidic proton is attached to an atom other than oxygen. The most significant of these are hydrohalic acids \[ \text{HX} \]. Where X represents a halogen atom.

13. When a strong acid molecule is placed in water the position of the dissociation equilibrium lies so far to the right that \([\text{HA}]\) cannot be measured accurately. This prevents an accurate calculation of \(K_a\).

**** Use Table 14.2 to order the following from the strongest to the weakest base:

\[ \text{H}_2\text{O} > \text{NO}_3^- > \text{Cl}^- > \text{H}_2\text{O} > \text{I}^- \]

Water is a stronger base than the conjugate base of a strong acid but a weaker base than the conjugate base of a weak acid.

The strength of an acid is inversely related to the strength of its conjugate base.
WATER AS AN ACID AND A BASE

1. A substance is said to be **amphoteric** if it can behave either as an acid or as a base.

2. Water is the most common amphoteric substance as can be seen in the **autoionization** of water:

   \[
   H_2O + H_2O \rightleftharpoons H^+ + OH^- + H_2O
   \]

3. The same process can occur for liquid ammonia.

4. The autoionization reaction for water

   \[
   2H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)
   \]

   Leads to the equilibrium expression:

   \[
   K_w = [H_3O^+] [OH^-] = [H^+] [OH^-]
   \]

   Where \( K_w \) is called the **ion-product constant** or the **dissociation constant** for water.

5. Experimental evidence shows that:

   \[
   [H^+] = [OH^-] = 1.0 \times 10^{-7} M
   \]

   Which means that:

   \[
   K_w = [H^+] [OH^-] = (1.0 \times 10^{-7})(1.0 \times 10^{-7}) = 1.0 \times 10^{-14}
   \]

6. In any aqueous solution at 25°C, no matter what it contains, the product of \([H^+][OH^-]\) must equal \(1.0 \times 10^{-14}\).
7. There are three possible situations:

a. A neutral solution, where $[H^+] = [OH^-]$

b. An acidic solution, where $[H^+] > [OH^-]$

c. A basic solution, where $[OH^-] > [H^+]$

** Calculate the $[H^+]$ or $[OH^-]$ as required for each of the following solutions at 25°C and state whether the solution is neutral, acidic, or basic.

a. $1.0 \times 10^{-8} \text{ M OH}^-$

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

$$[H^+] = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-8}} = 1.0 \times 10^{-6}$$

$b. 1.0 \times 10^{-7} \text{ M OH}^-$

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

$$[OH^-] = \frac{1.0 \times 10^{-7}}{1.0 \times 10^{-14}} = 1.0 \times 10^7$$

$c. 10.0 \text{ M H}^+$

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

$$[H^+] = 1.0 \times 10^{-14}$$

THE pH SCALE

1. Because $[H^+]$ in an aqueous solution is usually a small number, the pH scale is used to represent solution acidity.

2. $\rho H = -\log [H^+]$

The number of decimal places in the log is equal to the number of significant figures in the original number.

3. Other log scales are:

$$\rho OH = -\log [OH^-]$$

$$\rho K = -\log K$$
4. Since pH is a log scale based on 10, the pH changes by 1 for every power of ten change in [H⁺]. Also because pH is defined as \(-\log ([H^+]\)), then pH decreases as [H⁺] increases.

<table>
<thead>
<tr>
<th>[H⁺]</th>
<th>pH</th>
</tr>
</thead>
<tbody>
<tr>
<td>10⁻¹⁴</td>
<td>14</td>
</tr>
<tr>
<td>10⁻¹³</td>
<td>13</td>
</tr>
<tr>
<td>10⁻¹²</td>
<td>12</td>
</tr>
<tr>
<td>10⁻¹¹</td>
<td>11</td>
</tr>
<tr>
<td>10⁻¹⁰</td>
<td>10</td>
</tr>
<tr>
<td>10⁻⁹</td>
<td>9</td>
</tr>
<tr>
<td>10⁻⁸</td>
<td>8</td>
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<tr>
<td>10⁻⁷</td>
<td>7</td>
</tr>
<tr>
<td>10⁻⁶</td>
<td>6</td>
</tr>
<tr>
<td>10⁻⁵</td>
<td>5</td>
</tr>
<tr>
<td>10⁻⁴</td>
<td>4</td>
</tr>
<tr>
<td>10⁻³</td>
<td>3</td>
</tr>
<tr>
<td>10⁻²</td>
<td>2</td>
</tr>
<tr>
<td>10⁻¹</td>
<td>1</td>
</tr>
<tr>
<td>1</td>
<td>0</td>
</tr>
</tbody>
</table>

\(\rho H + \rho OH = 14.00\)

**** Calculate the pH and pOH of the following solutions:

a. \([OH^-] = 3.6 \text{ M}\)
\[\rho OH = \log([OH^-]) = -\log(3.6) = 0.56\]
\[\rho H = 14.00 - \rho OH = 14.00 - (0.56) = 13.44\]

b. \([OH^-] = 9.7 \times 10^{-9} \text{ M}\)
\[\rho OH = -\log([OH^-]) = -\log(9.7 \times 10^{-9}) = 8.01\]
\[\rho H = 14.00 - \rho OH = 14.00 - 8.01 = 5.99\]

c. \([OH^-] = 2.2 \times 10^{-3} \text{ M}\)
\[\rho OH = -\log([OH^-]) = -\log(2.2 \times 10^{-3}) = 2.66\]
\[\rho H = 14.00 - \rho OH = 14.00 - 2.66 = 11.34\]

d. \([OH^-] = 1.0 \times 10^{-7} \text{ M}\)
\[\rho OH = 7.00\]
\[\rho H = 7.00\]
CALCULATING THE pH OF STRONG ACID SOLUTIONS

1. When we deal with acid-base equilibria, we must focus on the solution components and their chemistry.

2. Determine the species present in the solution.

3. Determine which components are significant and which can be ignored. Focus on the major species, those solution components in relatively large amounts.

4. The key to solving these problems successfully is writing the major species in the solution.

**** What are the major species present in 0.250 M solution of each of the following acids? Calculate the pH of each of these solutions.

a. HClO₄

\[ \text{HClO}_4 (aq) \rightarrow \text{H}^+ (aq) + \text{ClO}_4^- (aq) \]
\[ \rho H = -\log [\text{H}^+] = -\log 0.250 = 0.602 \]

b. HNO₃

\[ \text{HNO}_3 (aq) \rightarrow \text{H}^+ (aq) + \text{NO}_3^- (aq) \]
\[ \rho H = -\log [\text{H}^+] = -\log 0.250 = 0.602 \]

CALCULATING THE pH OF WEAK ACID SOLUTIONS

1. As always, write the major species in the solution.

2. Look at hydrofluoric acid.

***** Find the pH of a 1.00 M solution of HF, \( K_a = 7.2 \times 10^{-4} \)

Since \( K_a \) is very small, we know that hydrofluoric acid is a weak acid and will be dissociated to only a slight extent.

\[ \text{HF} (aq) \rightleftharpoons \text{H}^+ (aq) + \text{F}^- (aq) \]

The major species in solution are: \( \text{HF} \quad \text{H}_2\text{O} \quad \text{H}_3\text{O} \)

The next step is to determine which of the major species can furnish H⁺ ions.

\[ \text{HF} (aq) \rightleftharpoons \text{H}^+ (aq) + \text{F}^- (aq) \quad K_a = 7.2 \times 10^{-4} \]
\[ \text{H}_2\text{O} \rightarrow \text{H}^+ (aq) + \text{OH}^- (aq) \quad K_w = 1.0 \times 10^{-14} \]
In aqueous solutions one source of $\text{H}^+$ can be singled out as dominant.

Compare $K_a$ and $K_w$. One can see that $K_a > K_w$.

Therefore, HF will be the dominant source of $\text{H}^+$.

$$\text{HF} (a_q) \rightleftharpoons \text{H}^+ (a_q) + \text{F}^- (a_q)$$

Set up the ICE table:

<table>
<thead>
<tr>
<th></th>
<th>$\text{H}^+$</th>
<th>$\text{F}^-$</th>
<th>$\text{HF}$</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\text{HF}$</td>
<td>+x</td>
<td></td>
<td></td>
</tr>
<tr>
<td>$\text{H}^+$</td>
<td></td>
<td>+x</td>
<td></td>
</tr>
<tr>
<td>$\text{F}^-$</td>
<td>x</td>
<td></td>
<td></td>
</tr>
<tr>
<td>$\text{HF}$</td>
<td>1-x</td>
<td>1-x</td>
<td>0</td>
</tr>
</tbody>
</table>

$$7.2 \times 10^{-4} = \frac{(x)(1-x)}{1-x}$$

Assume $x$ is small,

$$\chi^2 = 7.2 \times 10^{-4}$$

How valid is the approximation that $[\text{HF}] = 1.00 \text{ M}$? The validity of the approximation depends on how much accuracy we will demand for the calculation of $[\text{H}^+]$.

Since $K_a$ values are known to plus or minus 5%, it is reasonable to expect the same when determining the validity of the approximation.

$$\frac{[\text{HA}]}{[\text{HA}]}_0 - \chi \approx \frac{[\text{HA}]}{[\text{HA}]}_0$$

First calculate the value of $x$ by making the approximation.

$$K_a = \frac{x^2}{[\text{HA}]}_0 \approx \frac{x^2}{[\text{HA}]}_0$$

$$\chi^2 = K_a [\text{HA}]_0$$

Compare sizes of $x$ and $[\text{HA}]_0$. If $(x - [\text{HA}]_0) \times 100 \% \leq 5\%$, then the approximation is considered valid.

$$\chi = 2.7 \times 10^{-2} \text{ M}$$

$$\frac{\chi}{[\text{HA}]_0} \times 100 = \frac{2.7 \times 10^{-2}}{1.0} \times 100 = 2.7 \%$$

The approximation is valid.

$$\chi = [\text{H}^+] = 2.7 \times 10^{-2} \text{ M}$$

$$\varphi \text{H} = -\log (2.7 \times 10^{-2}) = 1.57$$
What are the major species present in 0.250M solution of HOCl₆H₃? Calculate the pH.

\[ K_a = 1.6 \times 10^{-10} \]

Major species: HOC₆H₅⁻ and H₂O

\[ HOC₆H₅⁻ (aq) \rightleftharpoons H^+ (aq) + OC₆H₅^- (aq) \]

\[
\begin{array}{c|c|c|c|c}
\text{I} & 0.250 & 0 & 0 & 4.33 \times 10^{-6} \\
\text{C} & 2 \chi & \chi & \chi & 0.250 (100) \\
\end{array}
\]

\[ \chi^2 = (0.250)(1.6 \times 10^{-10}) = \frac{x}{0.250} \approx \chi \text{ (small)} \]

\[ \chi = 6.32 \times 10^{-6} \]

\[ \text{pH} = 5.20 \]

---

**EXTRA EXAMPLES**

A solution is prepared by adding 50.0 mL of 0.050 M HCl to 150.0 mL of 0.10 M HNO₃.

Calculate the concentration of all species in the solution. What is the pH?

\[ \text{HCl}(aq) \rightarrow H^+ (aq) + Cl^- (aq) \]

mol HCl : (0.050 L)(0.05 mol/L) = 0.0025 mol

mol HNO₃ : (0.10 L)(0.10 mol/L) = 0.015 mol

\[ [H^+] = \frac{0.0025 + 0.015}{0.200} = 0.0875 \text{M} \]

\[ [Cl^-] = \frac{0.0025}{0.200} = 0.0125 \text{M} \]

\[ [NO_3^-] = \frac{0.015}{0.200} = 0.075 \text{M} \]

\[ \text{pH} = -\log [0.0875] = 1.06 \]
What mass of HNO₃ is present in 250.0 mL of a nitric acid solution having a pH = 5.10?

\[
\left[ H^+ \right] = 10^{-5.10} = 7.94 \times 10^{-6} \text{ M}
\]

\[
\text{HNO}_3(a_q) \rightarrow H^+(a_q) + NO_3^-(a_q)
\]

\[
\left[ \text{HNO}_3 \right] = 7.94 \times 10^{-6} \text{ M}
\]

\[
g \text{HNO}_3 = \frac{250 \text{ mL}}{1 \text{ L}} \times \frac{7.94 \times 10^{-6} \text{ mol}}{1 \text{ L}} \times \frac{63.02 \text{ g HNO}_3}{1 \text{ mol}}
\]

\[
= 1.28 \times 10^{-4} \text{ g}
\]

***** Calculate the concentration of all species present and the pH of a 0.0020 M HF solution. \(K_a = 7.2 \times 10^{-4}\)

\[
\text{HF}(a_q) \leftrightharpoons H^+(a_q) + F^-(a_q)
\]

\[
\begin{array}{c|c|c}
\text{I} & \text{0.002} & \text{0} \\
\text{C} & \chi & \chi \\
\text{E} & \text{0.002} - \chi & \chi
\end{array}
\]

\[
K_a = \frac{[H^+][F^-]}{[HF]} = 7.2 \times 10^{-4} = \frac{(\chi)(\chi)}{0.002 - \chi}
\]

\[
7.2 \times 10^{-4} = \frac{\chi^2}{0.002}
\]

\[
\chi^2 = 1.44 \times 10^{-6}
\]

\[
\chi = 0.012
\]

\[
\frac{0.0012}{0.002} = 60\% \text{ fail}
\]

\[
1.44 \times 10^{-6} - 7.2 \times 10^{-4} \chi = \chi^2
\]

\[
\chi^2 + 7.2 \times 10^{-4} \chi - 1.44 \times 10^{-6} = 0
\]

\[
\chi = \frac{-7.2 \times 10^{-4} \pm \sqrt{(7.2 \times 10^{-4})^2 - 4(-1.44 \times 10^{-6})}}{2(1)}
\]

\[
\chi = 8.93 \times 10^{-4}
\]

\[
[H^+] = [F^-] = 8.93 \times 10^{-4} \text{ M}
\]

\[
[HF] = 0.00111 \text{ M}
\]

\[
\text{pH} = 3.05
\]
A typical aspirin tablet contains 325 mg of acetylsalicylic acid (HC₉₀₄). Calculate the pH of a solution that is prepared by dissolving two aspirin tablets in one cup (237 mL) of solution. Assume the aspirin tablets are pure acetylsalicylic acid, \( K_a = 3.3 \times 10^{-4} \).

\[
\left[ \text{HC}_9\text{H}_7\text{O}_4 \right] = \frac{2 \text{ TABLETS}}{1 \text{ TABLET}} \times \frac{3.25 \text{ g HC}_9\text{H}_7\text{O}_4}{1 \text{ mL HC}_9\text{H}_7\text{O}_4} \times \frac{1 \text{ mL}}{0.157 \text{ L}} = 0.0152 \text{ M}
\]

\[
\text{HC}_9\text{H}_7\text{O}_4 (a.f.) \rightleftharpoons H^+ (a.g.) + \text{C}_9\text{H}_7\text{O}_4^- (a.g.)
\]

\[
3.3 \times 10^{-4} = \frac{\chi^2}{0.0152 - \chi}
\]

\[
\chi = 0.00224
\]

\[
\frac{0.0224}{0.0152} (\text{100}) = 14.7 \%
\]

The pH of weak acid mixtures

1. The approach used for a single weak acid is also used for a mixture of weak acids.

Calculate the pH of a solution that contains 1.00M HCN (\( K_a = 6.2 \times 10^{-10} \)) and 5.00M HNO₃ (\( K_a = 4.0 \times 10^{-4} \)). Also calculate the concentration of cyanide ion (CN⁻) in the solution equilibrium.

The major species are: HCN, HNO₃, H₂O

\[
\text{HCN (a.f.)} \rightleftharpoons H^+ (a.g.) + \text{CN}^- (a.g.) \quad K_a = 6.2 \times 10^{-10}
\]

\[
\text{HNO}_3 (a.g.) \rightleftharpoons H^+ (a.g.) + \text{NO}_3^- (a.g.) \quad K_a = 4.0 \times 10^{-4}
\]

\[
\text{H}_2\text{O} (l) \rightleftharpoons H^+ (a.g.) + \text{OH}^- (a.g.) \quad K_w = 1.0 \times 10^{-14}
\]

\[
\text{[CN}^-\text{]} = \frac{[\text{HCN}][\text{CN}^-]}{[\text{H}^+]} = \frac{1.0 \times 10^{-3} \times 1.0 \times 10^{-3}}{1.0 \times 10^{-4}} = 1.0 \times 10^{-2} \text{M}
\]

\[
\text{pH} = 2.168
\]
This could be complicated, but it isn’t because HNO₂, although a weak acid, is much stronger than the other two acids. HNO₂ is the dominant H⁺ producer and we can focus on:

\[ K_a = 4.0 \times 10^{-4} = \frac{[H^+][NO_2^-]}{[HNO_2]} \]

\[ \text{HNO}_2(a_q) \rightleftharpoons H^+(aq) + NO_2^-(aq) \]

\[ \chi = \frac{[H^+]}{[HNO_2]} = 4.5 \times 10^{-2} \]

\[ \rho H = 1.35 \]

\[ K_a = 4.0 \times 10^{-4} = \frac{[H^+][NO_2^-]}{[HNO_2]} \]

\[ [\text{HNO}_2] = \frac{[\text{H}^+][\text{NO}_2^-]}{K_a} = \frac{4.5 \times 10^{-2}}{4.0 \times 10^{-4}} = 1125 \]

\[ \text{PERCENT DISSOCIATION} = \frac{\text{AMOUNT DISSOCIATED}}{\text{INITIAL CONCENTRATION}} \times 100\% \]

1. The \textbf{PERCENT DISSOCIATION} is defined as follows:

\[ \text{PERCENT DISSOCIATION} = \frac{\text{AMOUNT DISSOCIATED}}{\text{INITIAL CONCENTRATION}} \times 100\% \]

2. For a given weak acid, the percent dissociation increases as the acid becomes more dilute.

**** Using the \( K_a \) values in Table 14.2, calculate the percent dissociation in a 0.20M solution of nitrous acid.

\[ \text{HNO}_2(a_q) \rightleftharpoons H^+(aq) + NO_2^-(aq) \]

\[ K_a = 4.0 \times 10^{-4} \]

\[ \chi = \frac{[H^+]}{[HNO_2]} = 8.9 \times 10^{-3} \]

\[ \rho H = 1.35 \]

\[ \text{PERCENT DISSOCIATION} = \frac{\text{AMOUNT DISSOCIATED}}{\text{INITIAL CONCENTRATION}} \times 100\% \]
3. For solutions of any weak acid \( HA \), \([H^+]\) decreases as \([HA]_0\) decreases, but the percent dissociation increases as \([HA]_0\) decreases.

***** In a 0.100M solution of HF, the percent dissociation is 8.1%. Calculate \( K_a \).

\[
\text{HF (aq)} \leftrightharpoons \text{H}^+ (\text{aq}) + \text{F}^- (\text{aq})
\]

<table>
<thead>
<tr>
<th>( C )</th>
<th>( x )</th>
<th>( \text{HF} )</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.10M</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>( C - x )</td>
<td>( x )</td>
<td>( x )</td>
</tr>
</tbody>
</table>

\[
K_a = \frac{[\text{H}^+][\text{F}^-]}{[\text{HF}]} = \frac{x^2}{0.10 - x}
\]

\[
\chi = \frac{[\text{H}^+]}{[\text{HF}]} = \frac{[\text{F}^-]}{[\text{HF}]} = \frac{(0.081)(0.100)}{8.1 \times 10^{-3}} = 9.4 M
\]

\[
[H_F] = 0.100 - 8.1 \times 10^{-3} = 0.0917 M
\]

**

\[
K_a = \frac{(8.1 \times 10^{-3})^2}{0.0917} = 7.1 \times 10^{-5}
\]

BASSES

1. According to the Arrhenius concept, a base is a substance that produces \( \text{OH}^- \) ions in aqueous solution. According to the Bronsted-Lowry model, a base is a proton acceptor.

2. If a base completely dissociates in solution then it is a **strong base**.

3. The Group 1 and Group 2 elements all form strong bases. The alkaline earth hydroxides are not very soluble.

***** Calculate the \( \text{pH} \) of the following solutions:

a. 0.0062M \( \text{Sr(OH)}_2 \)

\[
[\text{OH}^-] = 2(0.0062) = 0.012
\]

\[
\rho \text{OH} = -\log [\text{OH}^-] = -\log (0.012) = 1.92
\]

\[
\rho \text{H} = 14.00 - \rho \text{OH} = 14.00 - 1.92 = 12.08
\]

b. 0.75M \( \text{Sr(OH)}_2 \)

\[
[\text{OH}^-] = 2(0.75) = 1.50 M
\]

\[
\rho \text{OH} = -\log [\text{OH}^-] = -\log 1.50 = -0.18
\]

\[
\rho \text{H} = 14.00 - (-0.18) = 14.18
\]

c. 5.0 \times 10^{-5}M \( \text{Sr(OH)}_2 \)

**This is extremely dilute. The 1.0 \times 10^{-7}M \text{ OH}^- \text{ donated by } \text{H}_2\text{O will dominate.}**

\[
\rho \text{OH} = 7.00 = \rho \text{H}
\]
4. Many types of proton acceptors do not contain the hydroxide ion. However, when dissolved in water, these substances increase the concentration of hydroxide ion because of their reaction with water.

\[
\text{NH}_3 (aq) + H_2O (l) \rightleftharpoons \text{NH}_4^+ (aq) + OH^- (aq)
\]

5. Bases such as ammonia typically have at least one unshared pair of electrons that is capable of forming a bond with a proton.

6. The general reaction between a base and water is given by:

\[
\text{Base} + H_2O (l) \rightleftharpoons BH^+ (aq) + OH^- (aq)
\]

The equilibrium constant expression is:

\[
K_b = \frac{[BH^+][OH^-]}{[B]}
\]

Where \( K_b \) always refers to the reaction of a base with water to form the conjugate acid and the hydroxide ion.

7. pH calculations for solutions of weak bases are very similar to those for weak acids.

**** For the reaction of hydrazine \((N_2H_4)\) in water,

\[
H_2NNH_2 (aq) + H_2O (l) \leftrightarrow H_2NNH_3^+ (aq) + OH^- (aq) \quad K_b = 3.0 \times 10^{-6}
\]

Calculate the concentration of all species and the pH of a 2.0M solution of hydrazine in water.

\[
\begin{align*}
H_2NNH_2 (aq) + H_2O (l) & \rightleftharpoons H_2NNH_3^+ (aq) + OH^- (aq) \\
I & = 2.0 \\
C & = x \\
E & = 2 - x \\
K_b & = 3.0 \times 10^{-6} = \frac{x^2}{2} \\
[OH^-] & = \frac{[H_2NNH_3^+][OH^-]}{[H_2NNH_2]} = \frac{x^2}{2-x} \\
[OH^-] & = 2.4 \times 10^{-3} M \\
\rho OH & = 14.00 - 2.62 = 11.38 \\
\rho H & = 2.4 \times 10^{-3} M \\
\frac{2.4 \times 10^{-3}}{2} (10^2) & = 0.12 \gamma_o \\
[H_2NNH_2] & = 2.0 M \\
\text{and} \quad \text{pH} & = 10^{-11.38} = 4.2 \times 10^{-12} M
\end{align*}
\]
POLYPROTIC ACIDS

1. Some acids can furnish more than one proton and are called **polyprotic acids**.

2. A polyprotic acid always dissociates in a stepwise manner, one proton at a time.

3. The conjugate base of the first dissociation step becomes the acid in the next step.

   \[
   \begin{align*}
   H_2CO_3 (aq) & \rightleftharpoons H^+ (aq) + HCO_3^- (aq) \\
   K_{a1} & = \frac{[H^+][HCO_3^-]}{[H_2CO_3]} = 4.3 \times 10^{-7} \\
   HCO_3^- (aq) & \rightleftharpoons H^+ (aq) + CO_3^{2-} (aq) \\
   K_{a2} & = \frac{[H^+][CO_3^{2-}]}{[HCO_3^-]} = 5.6 \times 10^{-11}
   \end{align*}
   \]

4. For a typical weak polyprotic acid:

   \[K_{a1} > K_{a2} > K_{a3}\]

5. The acid involved in each step of the dissociation is successively weaker. The loss of a second or third proton occurs less readily than the loss of the first proton. As the negative charge on the acid increases, it becomes more difficult to remove the positively charged proton.

6. For a typical polyprotic acid in water, only the first dissociation step is important in determining the pH.

**** Calculate the pH of a 5.0 M \(H_3PO_4\) solution and the equilibrium concentrations of the species \(H_3PO_4\), \(H_2PO_4^-\), \(HPO_4^{2-}\), and \(PO_4^{3-}\):

\[
H_3PO_4 (aq) \rightleftharpoons H^+ (aq) + H_2PO_4^- (aq)
\]

\[
\begin{array}{c|c|c|c}
I & S & 0
\end{array}
\]

\[
\begin{array}{c|c|c|c}
C & -\chi & +\chi & +\gamma \\
E & S-\chi & \chi & \chi
\end{array}
\]

\[
K_{a1} = \frac{[H^+][H_2PO_4^-]}{[H_3PO_4]} = 7.5 \times 10^{-3} = \frac{\chi^2}{S-\chi}
\]

\[
\chi = 1.9 \times 10^{-1}
\]

\[
\rho H = \rho 7.3
\]

\[\frac{[H^+]}{[H_3PO_4]} = 0.19
\]

\[\frac{[H_2PO_4^-]}{[H_3PO_4]} = 0.19
\]
7. Sulfuric acid is a strong acid so the first dissociation sets the pH. In dilute H₂SO₄ solutions, the second dissociation step contributes significantly to [H⁺].

ACID-BASE PROPERTIES OF SALTS

1. Salt is simply another name for \textit{ionic compound}.

2. When a salt dissolves in water it breaks up into ions. Under certain conditions, these ions can behave as acids or bases.

\textbf{Salts That Produce Neutral Solutions}

1. Salts that consist of cations of strong bases and the anions of strong acids have no effect on the [H⁺] when dissolved in water.

2. Neither species has any affinity for H⁺ nor can they produce H⁺.

3. This means that aqueous solutions of salts such as KCl, NaCl, NaNO₃ and KNO₃ are neutral.

\textbf{Salts That Produce Basic Solutions}

1. For sodium acetate, NaC₂H₃O₂, the major species are: \( \text{Na}^+ \), \( \text{C}_2\text{H}_3\text{O}_2^- \), and \( \text{H}_2\text{O} \).

\( \text{Na}^+ \) has no affinity for H⁺

\( \text{C}_2\text{H}_3\text{O}_2^- \) is the conjugate base of acetic acid and has an affinity for H⁺

\( \text{H}_2\text{O} \) is amphoteric

\[ \text{C}_2\text{H}_3\text{O}_2^- (aq) + \text{H}_2\text{O} \rightarrow \text{HC}_2\text{H}_3\text{O}_2 (aq) + \text{OH}^- (aq) \]

The above reaction is a base reacting with water to produce a hydroxide ion and a conjugate acid.

\[ K_a = 1.8 \times 10^{-5} \quad K_a \times K_b = K_w \quad K_b = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10} \]

2. For any salt whose cation has neutral properties (Na⁺ or K⁺) and whose anion is the conjugate base of a weak acid, the aqueous solution will be basic.
Calculate the pH of a 0.30M NaF solution. The $K_a$ value for HF is $7.2 \times 10^{-4}$.

Major species are: $Na^+, F^-$, and $H_2O$

HF is a weak acid; $F^-$ has a strong affinity for protons.

$$F^-(aq) + H_2O(l) \leftrightarrow HF(aq) + OH^-(aq)$$

$$K_b = \frac{[HF][OH^-]}{[F^-]}$$

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{7.2 \times 10^{-4}} = 1.4 \times 10^{-10}$$

$$[OH^-] = \chi = 2.0 \times 10^{-6} M$$

$$\rho OH = 5.69$$

$$\rho H = 8.31$$

$$\chi = \frac{x^2}{3} \Rightarrow x = \frac{\chi^2}{3}$$

In this reaction, $CN^-$ appears to be a weak base.

3. In the reaction of cyanide ion with water, the cyanide ion is competing with the hydroxide ion for $H^+$ instead of competing with water.

**Salts That Produce Acidic Solutions**

1. Some salts produce acidic solutions when dissolved in water. $NH_4Cl$

$$NH_4^+(aq) \overset{H_2O}{\rightarrow} NH_3(aq) + H^+(aq)$$

The $Cl^-$ ion, having virtually no affinity for $H^+$ in water, does not affect the pH of the solution.

3. Salts in which the anion is not a base and the cation is the conjugate acid of a weak base produce more acidic solutions.
Calculate the pH of a 0.10M NH₄Cl solution. \( K_b \) for NH₃ is 1.8 \times 10^{-5}.

Major species in the solution are: \( \text{NH}_4^+ \), \( \text{Cl}^- \) and \( \text{H}_2\text{O} \)

\[
\text{NH}_4^+ \rightleftharpoons \text{NH}_3 + \text{H}^+
\]

\[
K_a = \frac{[\text{H}^+][\text{NH}_3]}{[\text{NH}_4^+]} = \frac{[\text{H}^+][\text{NH}_3]}{K_b} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10}
\]

\( \text{NH}_4^+ \) is a very weak acid, but it is stronger than \( \text{H}_2\text{O} \) so it will dominate in the production of \( \text{H}^+ \).

\[
\text{NH}_4^+(\text{aq}) \rightleftharpoons \text{NH}_3(\text{aq}) + \text{H}^+(\text{aq})
\]

\[
K_a = \frac{[\text{H}^+][\text{NH}_3]}{[\text{NH}_4^+]} = \frac{(x)(x)}{0.10} = \frac{x^2}{0.10} \approx \frac{x^2}{1}
\]

\[
[\text{H}^+] = x = 7.5 \times 10^{-5} \text{M}
\]

\[
\rho H = 5.13
\]

2. A second type of salt that produces an acidic solution is one that contains a highly charged metal ion.

3. Dissolving \( \text{AlCl}_3 \) in water results in an acidic solution. This is due to the hydrated ion being a weak acid:

\[
\text{Al}^3+(\text{aq}) \rightleftharpoons \text{Al}(\text{aq}) + 3 \text{H}^+(\text{aq})
\]

5. The high charge on the metal ion polarizes the \( \text{O} - \text{H} \) bonds in the attached water molecules making the hydroniums more acidic.

6. The higher the charge on the metal ion, the stronger the acidity of the hydrated ion.

Calculate the pH of a 0.10M CoCl₃ solution. The \( K_a \) value for \( \text{Co}^{3+} \) is 1.0 \times 10^{-5}

Major species: \( \text{Co}^{3+}(\text{aq}) \), \( \text{Cl}^- \), \( \text{H}_2\text{O} \)

\( \text{Co}^{3+}(\text{aq}) \) will determine the pH since it is a stronger acid than water.

\[
\text{Co}^{3+}(\text{aq}) \rightleftharpoons \text{Co}(\text{aq}) + 3 \text{H}^+(\text{aq})
\]

\[
K_a = 1.0 \times 10^{-5} = \frac{[\text{Co}(\text{aq})][\text{H}^+]}{[\text{Co}^{3+}(\text{aq})]} = \frac{(x)(x)}{0.10} \approx \frac{x^2}{1}
\]

\[
[\text{H}^+] = x = 1.0 \times 10^{-3} \text{M}
\]

\[
\rho H = 3.00
\]
1. We have seen that when a substance is dissolved in water it produces an acidic solution if it can donate protons and a basic solution if it can accept protons.

2. Any molecule containing a hydrogen atom is potentially an acid. However, many such molecules show no acidic properties.

3. There are two main factors that determine whether a molecule containing an $X - H$ bond will behave as a Bronsted-Lowry acid:
   (a) **THE STRENGTH OF THE BOND**
   (b) **THE POLARITY OF THE BOND**

4. Based on these criteria, HF should be a very strong acid. In fact, HF is the only weak acid in the hydrogen halide series. This is because the $H - F$ bond is difficult to break when dissolved in water. HF does not dissociate easily.

5. Oxyacids are another important class of acids: $H - O - X$.

6. The acid strength of oxyacids increases with an increase in the number of oxygen atoms attached to the central atom. This happens because the very electronegative oxygen atoms are able to draw electrons away from the halide atom and the $O - H$ bond.

7. The $O - H$ bond is both polarized and weakened. This means that a proton is most readily produced by the molecule with the largest number of attached oxygen atoms.

8. This type of behavior is also observed for hydrated metal ions. The acidity of the water molecules attached to the metal ion is increased by the attraction of electrons to the positive metal ion.

9. The greater the charge on the metal ion, the more acidic the hydrated ion becomes.

---

### Table 14.6 Acid–Base Properties of Various Types of Salts

<table>
<thead>
<tr>
<th>Type of Salt</th>
<th>Examples</th>
<th>Comment</th>
<th>pH of Solution</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cation is from strong base; anion is from strong acid</td>
<td>KCl, KNO₃, NaCl, NaNO₃</td>
<td>Neither acts as an acid or a base</td>
<td>Neutral</td>
</tr>
<tr>
<td>Cation is from strong base; anion is from weak acid</td>
<td>Na₃C₆H₅O₇, KCN, NaF</td>
<td>Anion acts as a base; cation has no effect on pH</td>
<td>Basic</td>
</tr>
<tr>
<td>Cation is conjugate acid of weak base; anion is from strong acid</td>
<td>NH₄Cl, NH₄NO₃</td>
<td>Cation acts as acid; anion has no effect on pH</td>
<td>Acidic</td>
</tr>
<tr>
<td>Cation is conjugate acid of weak base; anion is conjugate base of weak acid</td>
<td>NH₄C₆H₅O₇, NH₄CN</td>
<td>Cation acts as an acid; anion acts as a base</td>
<td>Acidic if $K_a &gt; K_b$, basic if $K_b &gt; K_a$, neutral if $K_a = K_b$</td>
</tr>
<tr>
<td>Cation is highly charged metal ion; anion is from strong acid</td>
<td>Al(ClO₄)₃, FeCl₃</td>
<td>Hydrated cation acts as an acid; anion has no effect on pH</td>
<td>Acidic</td>
</tr>
</tbody>
</table>
TABLE 14.8 Several Series of Oxyacids and Their $K_a$ Values

<table>
<thead>
<tr>
<th>Oxyacid</th>
<th>Structure</th>
<th>$K_a$ Value</th>
</tr>
</thead>
<tbody>
<tr>
<td>HClO₄</td>
<td>$\text{H-O-Cl-O}$</td>
<td>Large (~$10^3$)</td>
</tr>
<tr>
<td>HClO₃</td>
<td>$\text{H-O-Cl}$</td>
<td>~1</td>
</tr>
<tr>
<td>HClO₂</td>
<td>$\text{H-O-Cl-O}$</td>
<td>$1.2 \times 10^{-2}$</td>
</tr>
<tr>
<td>HClO</td>
<td>$\text{H-O-Cl}$</td>
<td>$3.5 \times 10^{-8}$</td>
</tr>
<tr>
<td>H₂SO₄</td>
<td>$\text{H-O-S=O}$</td>
<td>Large</td>
</tr>
<tr>
<td>H₂SO₃</td>
<td>$\text{H-O-S=O}$</td>
<td>$1.5 \times 10^{-2}$</td>
</tr>
<tr>
<td>HNO₃</td>
<td>$\text{H-O=N-O}$</td>
<td>Large</td>
</tr>
<tr>
<td>HNO₂</td>
<td>$\text{H-O-N-O}$</td>
<td>$4.0 \times 10^{-4}$</td>
</tr>
</tbody>
</table>

ACID-BASE PROPERTIES OF OXIDES

1. Molecules containing the grouping $\text{H-O-X}$ can behave as acids and the acid strength depends on the electron withdrawing ability of $X$.

2. Substances in this group can also behave as bases if the hydroxide ion instead of a proton is produced.

3. If $X$ has a relatively high electronegativity, the $\text{O-X}$ bond will be covalent and strong. When $\text{H-O-X}$ compound is dissolved in water, the $\text{O-X}$ bond will remain intact. The polar, and relatively weaker, $\text{O-H}$ bond will break and release a proton.

4. If $X$ has a relatively low electronegativity, the $\text{O-X}$ bond will be ionic and subject to being broken in polar water.

5. When a covalent oxide dissolves in water, an acidic solution forms. These oxides are called **ACIDIC OXIDES**.  

   \[
   \text{SO}_2 (g) + \text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{SO}_3 (aq)
   \]

6. When an ionic oxide dissolves in water, a basic solution results. The oxides are called **BASIC OXIDES**.  

   \[
   \text{CaO}(s) + \text{H}_2\text{O}(l) \rightarrow \text{Ca(OH)}_2 (aq)
   \]
THE LEWIS ACID-BASE MODEL

1. This is an even more general model for acid-base behavior.

2. A Lewis acid is an electron-pair acceptor and a Lewis base is an electron-pair donor.

3. A Lewis acid has an empty atomic orbital that it can use to accept (share) an electron pair from a molecule that has a lone pair of electrons.

4. The Lewis model encompasses the Bronsted-Lowry model, but the reverse is not true.

5. A Bronsted-Lowry acid-base reaction (H+ + O- \rightarrow H2O) are encompassed by the Lewis model:

\[
\begin{align*}
\text{H}^+ &+ \text{N} &\rightarrow\left[ \begin{array}{c}
\text{H}\\
\text{N}
\end{array} \right] \\
\text{LEWIS} &\text{ACID} &\text{LEWIS} &\text{BASE} \\
\text{E-PAIR} &\text{ACCEPTOR} &\text{E-PAIR} &\text{DONOR}
\end{align*}
\]

6. An example of the Lewis model not covered by the Bronsted-Lowry model is:

\[
\text{Fe}^{3+} + 6 \left[ \text{C} \equiv \text{N} \right]^\text{-} \rightarrow \left[ \text{Fe} \left( \text{C} \equiv \text{N} \right)_6 \right]^3\text{-}
\]

<table>
<thead>
<tr>
<th>Model</th>
<th>Definition of Acid</th>
<th>Definition of Base</th>
</tr>
</thead>
<tbody>
<tr>
<td>Arrhenius</td>
<td>H⁺ producer</td>
<td>OH⁻ producer</td>
</tr>
<tr>
<td>Bronsted–Lowry</td>
<td>H⁺ donor</td>
<td>H⁺ acceptor</td>
</tr>
<tr>
<td>Lewis</td>
<td>Electron-pair acceptor</td>
<td>Electron-pair donor</td>
</tr>
</tbody>
</table>

TABLE 14.10 Three Models for Acids and Bases
7. The hydration of a metal ion can be viewed as a Lewis acid-base reaction:

\[ \text{Al}^{3+} + 6 \text{L} \rightarrow [\text{AlL}_6]^{3+} \]

*Identify the Lewis acid and the Lewis base in each of the following reactions:

a. \( \text{Fe}^{3+} (aq) + 6 \text{H}_2\text{O} (l) \leftrightarrow \text{Fe(H}_2\text{O)}_6^{3+} (aq) \)

\[ \text{LA} \text{ L} \text{B} \]

b. \( \text{H}_2\text{O} (l) + \text{CN}^- (aq) \leftrightarrow \text{HCN} (aq) + \text{OH}^- (aq) \)

\[ \text{L} \text{ A} \text{ L} \text{ B} \]

c. \( \text{HgI} (s) + 2 \text{I}^- (aq) \leftrightarrow \text{HgI}_4^{2+} (aq) \)

\[ \text{L} \text{ A} \text{ L} \text{ B} \]

---

**TABLE 14.1 Various Ways to Describe Acid Strength**

<table>
<thead>
<tr>
<th>Property</th>
<th>Strong Acid</th>
<th>Weak Acid</th>
</tr>
</thead>
<tbody>
<tr>
<td>( K_a ) value</td>
<td>( K_a ) is large</td>
<td>( K_a ) is small</td>
</tr>
<tr>
<td>Position of the dissociation (ionization) equilibrium</td>
<td>Far to the right</td>
<td>Far to the left</td>
</tr>
<tr>
<td>Equilibrium concentration of H(^+) compared with original concentration of HA</td>
<td>( [\text{H}^+] \approx [\text{HA}]_0 )</td>
<td>( [\text{H}^+] &lt; [\text{HA}]_0 )</td>
</tr>
<tr>
<td>Strength of conjugate base compared with that of water</td>
<td>( A^- ) much weaker base than ( \text{H}_2\text{O} )</td>
<td>( A^- ) much stronger base than ( \text{H}_2\text{O} )</td>
</tr>
</tbody>
</table>
### TABLE 14.2 Values of $K_a$ for Some Common Monoprotic Acids

<table>
<thead>
<tr>
<th>Formula</th>
<th>Name</th>
<th>Value of $K_a$</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\text{HSO}_4^-$</td>
<td>Hydrogen sulfate ion</td>
<td>$1.2 \times 10^{-2}$</td>
</tr>
<tr>
<td>$\text{HClO}_2$</td>
<td>Chlorous acid</td>
<td>$1.2 \times 10^{-2}$</td>
</tr>
<tr>
<td>$\text{HC}_2\text{H}_3\text{O}_2^+$</td>
<td>Monochloracetic acid</td>
<td>$1.35 \times 10^{-3}$</td>
</tr>
<tr>
<td>$\text{HF}$</td>
<td>Hydrofluoric acid</td>
<td>$7.2 \times 10^{-4}$</td>
</tr>
<tr>
<td>$\text{HNO}_2$</td>
<td>Nitrous acid</td>
<td>$4.0 \times 10^{-4}$</td>
</tr>
<tr>
<td>$\text{HC}_2\text{H}_3\text{O}_3^+$</td>
<td>Acetic acid</td>
<td>$1.8 \times 10^{-5}$</td>
</tr>
<tr>
<td>$[\text{Al(\text{H}_2\text{O})_6}]^3^+$</td>
<td>Hydrated aluminum(III) ion</td>
<td>$1.4 \times 10^{-5}$</td>
</tr>
<tr>
<td>$\text{HOCI}$</td>
<td>Hypochlorous acid</td>
<td>$3.5 \times 10^{-8}$</td>
</tr>
<tr>
<td>$\text{HCN}$</td>
<td>Hydrocyanic acid</td>
<td>$6.2 \times 10^{-10}$</td>
</tr>
<tr>
<td>$\text{NH}_4^+$</td>
<td>Ammonium ion</td>
<td>$5.6 \times 10^{-10}$</td>
</tr>
<tr>
<td>$\text{HOC}_6\text{H}_5$</td>
<td>Phenol</td>
<td>$1.6 \times 10^{-10}$</td>
</tr>
</tbody>
</table>

*The units of $K_a$ are customarily omitted.

### TABLE 14.3 Values of $K_b$ for Some Common Weak Bases

<table>
<thead>
<tr>
<th>Name</th>
<th>Formula</th>
<th>Conjugate Acid</th>
<th>$K_b$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ammonia</td>
<td>$\text{NH}_3$</td>
<td>$\text{NH}_4^+$</td>
<td>$1.8 \times 10^{-5}$</td>
</tr>
<tr>
<td>Methylamine</td>
<td>$\text{CH}_3\text{NH}_2$</td>
<td>$\text{CH}_3\text{NH}_3^+$</td>
<td>$4.38 \times 10^{-4}$</td>
</tr>
<tr>
<td>Ethylamine</td>
<td>$\text{C}_2\text{H}_5\text{NH}_2$</td>
<td>$\text{C}_2\text{H}_5\text{NH}_3^+$</td>
<td>$5.6 \times 10^{-4}$</td>
</tr>
<tr>
<td>Aniline</td>
<td>$\text{C}_6\text{H}_5\text{NH}_2$</td>
<td>$\text{C}_6\text{H}_5\text{NH}_3^+$</td>
<td>$3.8 \times 10^{-10}$</td>
</tr>
<tr>
<td>Pyridine</td>
<td>$\text{C}_5\text{H}_5\text{N}$</td>
<td>$\text{C}_5\text{H}_5\text{NH}_3^+$</td>
<td>$1.7 \times 10^{-9}$</td>
</tr>
</tbody>
</table>

### TABLE 14.4 Stepwise Dissociation Constants for Several Common Polyprotic Acids

<table>
<thead>
<tr>
<th>Name</th>
<th>Formula</th>
<th>$K_{s1}$</th>
<th>$K_{s2}$</th>
<th>$K_{s3}$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Phosphoric acid</td>
<td>$\text{H}_3\text{PO}_4$</td>
<td>$7.5 \times 10^{-3}$</td>
<td>$6.2 \times 10^{-8}$</td>
<td>$4.8 \times 10^{-13}$</td>
</tr>
<tr>
<td>Arsenic acid</td>
<td>$\text{H}_3\text{AsO}_4$</td>
<td>$5 \times 10^{-3}$</td>
<td>$8 \times 10^{-8}$</td>
<td>$6 \times 10^{-10}$</td>
</tr>
<tr>
<td>Carbonic acid</td>
<td>$\text{H}_2\text{CO}_3$</td>
<td>$4.3 \times 10^{-7}$</td>
<td>$5.6 \times 10^{-11}$</td>
<td>( \approx 10^{-10} )</td>
</tr>
<tr>
<td>Sulfuric acid</td>
<td>$\text{H}_2\text{SO}_4$</td>
<td>Large</td>
<td>$1.2 \times 10^{-2}$</td>
<td>( \approx 10^{-7} )</td>
</tr>
<tr>
<td>Sulfurous acid</td>
<td>$\text{H}_2\text{SO}_3$</td>
<td>$1.5 \times 10^{-2}$</td>
<td>$1.0 \times 10^{-7}$</td>
<td>( \approx 10^{-10} )</td>
</tr>
<tr>
<td>Hydrosulfuric acid</td>
<td>$\text{H}_2\text{S}$</td>
<td>$1.0 \times 10^{-7}$</td>
<td>( \approx 10^{-10} )</td>
<td>( \approx 10^{-10} )</td>
</tr>
<tr>
<td>Oxalite acid</td>
<td>$\text{C}_2\text{H}_4\text{O}_4$</td>
<td>$6.5 \times 10^{-2}$</td>
<td>$6.1 \times 10^{-5}$</td>
<td>( \approx 10^{-10} )</td>
</tr>
<tr>
<td>Ascorbic acid</td>
<td>(vitamin C)</td>
<td>$7.9 \times 10^{-5}$</td>
<td>$1.6 \times 10^{-12}$</td>
<td>( \approx 10^{-10} )</td>
</tr>
</tbody>
</table>

The $K_{s2}$ value for $\text{H}_2\text{S}$ is very uncertain. Because it is so small, the $K_{s2}$ value is very difficult to measure accurately.