During Class Invention #

Name(s) with Lab section in Group

Arrhenius Acids, [H⁺], [OH⁻] pH and pOH

1a. Define the terms Arrhenius acid and Arrhenius base.

An acid is a substance which, when dissolved in water, increases the concentration of hydrogen ion, H⁺(aq). For example,

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

A base is a substance which, when added to water, increases the concentration of hydroxide ion, OH⁻(aq). For example,

\[ \text{NaOH}(aq) \rightarrow \text{Na}^+(aq) + \text{OH}^-(aq) \]

b. Write a chemical equation that describes the behavior of an Arrhenius acid in water.

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

c. Write a chemical equation that describes the behavior of an Arrhenius base in water.

\[ \text{NaOH}(aq) \rightarrow \text{Na}^+(aq) + \text{OH}^-(aq) \]

2. In the space below, list some examples of Arrhenius acids and Arrhenius bases.

<table>
<thead>
<tr>
<th>Arrhenius Acids</th>
<th>Arrhenius Bases</th>
</tr>
</thead>
<tbody>
<tr>
<td>HCl</td>
<td>NaOH</td>
</tr>
<tr>
<td>H₂SO₄</td>
<td>KOH</td>
</tr>
<tr>
<td>HC₂H₃O₂</td>
<td>Ba(OH)₂</td>
</tr>
<tr>
<td>HBr</td>
<td>Ca(OH)₂</td>
</tr>
<tr>
<td>HClO₄</td>
<td></td>
</tr>
</tbody>
</table>

2a. Write the autoionization reaction for water and the equilibrium expression for the autoionization reaction.

\[ \text{H}_2\text{O}(l) \rightleftharpoons \text{H}^+(aq) + \text{OH}^- (aq) \]

\[ K_w = [\text{H}^+][\text{OH}^-] \]

or another way

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

\[ K_w = [\text{H}_3\text{O}^+][\text{OH}^-] \]
b. What is the magnitude of the equilibrium constant at 25 °C for the autoionization reaction of water?

\[ K_w = 1.0 \times 10^{-14} \]

c. What are the concentrations of H\(^+\) and OH\(^-\) at 25 °C in pure water?

\[ \text{H}_2\text{O}(l) \rightleftharpoons \text{H}^+(aq) + \text{OH}^-(aq) \]

\[
\begin{array}{c|c|c}
\text{Initial} & 0 & 0 \\
\hline
\text{Change} & +x & +x \\
\hline
\text{Equilibrium} & 0+x & 0+x \\
\end{array}
\]

\[ x = [\text{H}_2\text{O}]_{\text{reacted}} \]

\[ K_w = [\text{H}^+][\text{OH}^-] \]

\[ 1.0 \times 10^{-14} = (x)(x) \]

\[ 1.0 \times 10^{-7} \, \text{M} = x \]

d. The [H\(^+\)] in a particular aqueous solution is 1.0 \times 10^{-4} \, \text{M}. Calculate the [OH\(^-\)] for this solution.

\[ K_w = [\text{H}^+][\text{OH}^-] \]

\[ 1.0 \times 10^{-14} = (1.0 \times 10^{-4} \, \text{M})(\text{[OH}^-]) \]

\[ \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-4} \, \text{M}} = [\text{OH}^-] \]

\[ 1.0 \times 10^{-10} \, \text{M} = [\text{OH}^-] \]

e. The [OH\(^-\)] in a particular aqueous solution is 1.0 \times 10^{-5} \, \text{M}. Calculate the [H\(^+\)] of this solution.

\[ K_w = [\text{H}^+][\text{OH}^-] \]

\[ 1.0 \times 10^{-14} = [\text{H}^+](1.0 \times 10^{-5} \, \text{M}) \]

\[ \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-5} \, \text{M}} = [\text{H}^+] \]

\[ 1.0 \times 10^{-9} \, \text{M} = [\text{H}^+] \]

f. The [H\(^+\)] in a particular aqueous solution is 6.0 \, \text{M}. Calculate the [OH\(^-\)] of this solution.

\[ K_w = [\text{H}^+][\text{OH}^-] \]

\[ 1.0 \times 10^{-14} = (6.0 \, \text{M})(\text{[OH}^-]) \]

\[ \frac{1.0 \times 10^{-14}}{6.0 \, \text{M}} = [\text{OH}^-] \]

\[ 1.7 \times 10^{-15} \, \text{M} = [\text{OH}^-] \text{ note it is possible for the [H}^+] or the [OH}^-] to be smaller than 1 \times 10^{-14} \, \text{M}. \]
3a. Define $pH$ and $pOH$ for aqueous solutions of acids or bases.

\[
\begin{align*}
\text{pH} &= -\log[H^+] \\
pOH &= -\log[OH^-] \\
pH + pOH &= 14 \\
\text{For neutral aqueous solutions, the pH = 7.}
\end{align*}
\]

b. Indicate the range of pH which characterizes an acidic solution and the range which characterizes a basic solution.

<table>
<thead>
<tr>
<th>pH Range</th>
<th>Description</th>
</tr>
</thead>
<tbody>
<tr>
<td>0 - 6.99</td>
<td>acidic</td>
</tr>
<tr>
<td>7</td>
<td>neutral</td>
</tr>
<tr>
<td>7.01 - 14</td>
<td>basic</td>
</tr>
</tbody>
</table>

c. Calculate the pH and pOH of a solution with a $[H^+] = 3.68 \times 10^{-8}$ M.

\[
\begin{align*}
\text{pH} &= -\log[H^+] \\
pH &= -\log[3.68 \times 10^{-8}] \\
pH &= -(-7.43) \\
pH &= 7.43 \\
pH + pOH &= 14 \\
7.43 + pOH &= 14 \\
pOH &= 14 - 7.43 \\
pOH &= 6.57
\end{align*}
\]

d. Calculate the $[H^+]$ and $[OH^-]$ of a solution with a pH = 4.22.

\[
\begin{align*}
\text{pH} &= -\log[H^+] \\
4.22 &= -\log[H^+] \\
-4.22 &= \log[H^+] \\
10^{-4.22} &= 10^{\log[H^+]} \\
6.03 \times 10^{-5} \text{ M} &= [H^+] \\
K_w &= [H^+][OH^-] \\
1.0 \times 10^{-14} &= (6.03 \times 10^{-5})[OH^-] \\
1.0 \times 10^{-14} &= 6.03 \times 10^{-5} \times [OH^-] \\
1.66 \times 10^{-10} \text{ M} &= [OH^-]
\end{align*}
\]