For any Acid:

\[ HA + H_2O \rightarrow H_3O^+ (aq) + A^- (aq) \]

Acid + Water \(\rightarrow\) Hydronium + Acid anion

The hydrogen of the acid is essentially a single proton that can bond through a coordination bond to water forming \(H_3O^+\), called the hydronium ion.

\[ H^+ + \overset{\text{O}}{\text{H}} - \rightarrow \overset{\text{O}}{\text{H}} - \overset{\text{O}}{\text{H}}^+ \]

Examples:

1. \( \text{HCl}(g) + H_2O(l) \rightarrow H_3O^+ (aq) + Cl^- (aq) \)

However, just because an acid is a strong acid, that does not mean it will only completely dissociate (ionize) with water:

\[ \overset{\text{N}}{\text{H}} - \overset{\text{O}}{\text{H}} + \overset{\text{N}}{\text{H}} \rightarrow \overset{\text{N}}{\text{H}} - \overset{\text{N}}{\text{H}}^+ + \overset{\text{O}}{\text{H}} - \overset{\text{O}}{\text{H}} \]

Strong Acids

- You will recall that the seven strong acids are HCl, HBr, HI, HNO₃, H₂SO₄, HClO₃, and HClO₄.
- These are, by definition, strong electrolytes and exist totally as ions in aqueous solution.
- For the monoprotic strong acids,

\[ [H_3O^+] = [\text{acid}] \]

For any Base:

\[ \text{MOH} + H_2O \leftrightarrow M^+ + OH^- + H_2O \]

Water acts as a solvent and does not form a complex for Arrhenius bases as it does for Arrhenius acids

Therefore,

\[ \text{NaOH} \overset{\text{H}_2\text{O}}{\rightarrow} \text{Na}^{+}(aq) + \text{OH}^- (aq) \]

\[ \text{Ca(OH)}_2 \overset{\text{H}_2\text{O}}{\rightarrow} \text{Ca}^{2+}(aq) + 2\text{OH}^- (aq) \]
**Strong Bases**

- Strong bases are the soluble hydroxides, which are the alkali metal and heavier alkaline earth metal hydroxides (Ca$^{2+}$, Sr$^{2+}$, and Ba$^{2+}$).
- Again, these substances dissociate completely in aqueous solution.

**Autoionization of Water**

*Autoionization* = self ionization of a molecule into ions by the intermolecular forces of the substance

In water, a very small fraction of molecules are ionized by the effects of hydrogen bonding.

\[
\text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{OH}^-
\]

Notice, this is an equilibrium reaction where:

\[
K = \frac{[\text{H}_3\text{O}^+][\text{OH}^-]}{[\text{H}_2\text{O}]}
\]

Remember, H$_2$O is a pure liquid, therefore the equilibrium expression reads:

\[
K = [\text{H}_3\text{O}^+][\text{OH}^-] = K_w
\]

$K_w$ is called the ion-product constant of water. *(AP equation Sheet)*

\[K_w = 1.0 \times 10^{-14} \text{ ( @ 25.0°C)}\]

You should memorize this number.

In neutral solutions, the concentration of hydronium ions must equal the concentration of hydroxide ions.

Since water is neutral

\[ [\text{H}_3\text{O}^+] = [\text{OH}^-] = 1.0 \times 10^{-7} \]

Using the ion-product constant of water, it is possible to calculate the H$^+$ concentration or OH$^-$ concentration in any acidic or basic solution if either is known.

This is true because the product of H$^+$ and OH$^-$ in any aqueous solution must always equal the ion-product constant (1.0 x 10$^{-14}$)

Because the ion product of H$^+$ and OH$^-$ is constant if:

- \([\text{H}^+] > 1.0 \times 10^{-7}\) the solution is acidic
- \([\text{H}^+] < 1.0 \times 10^{-7}\) the solution is basic
- Or -
- \([\text{H}^+] > [\text{OH}^-]\) acidic solution
- \([\text{H}^+] < [\text{OH}^-]\) basic solution
1. A 0.010 M NaOH solution was made in the laboratory. What is the hydronium ion concentration of the solution?

Because acid and base concentrations are usually extremely low, a log scale is generally used to make the numbers easier to work with.

\[ \text{pH} = -\log [\text{H}_3\text{O}^+] \]  

This is how we get the pH of water equal to 7

\[ [\text{H}_3\text{O}^+] = [\text{OH}^-] = 1.0 \times 10^{-7} \]

\[ \text{pH} = -\log (1.0 \times 10^{-7}) = 7 \]

### Acid Concentration

In the laboratory, we have worked with extremely strong acid solutions (0.1, 1.0, 3.0, 6.0 M, ...)

In the natural and biological world, acid concentrations are much much lower.

For example, stomach acid is only 1.00 x 10\(^{-2}\) M or 0.010 mol/L HCl

### pH scale

<table>
<thead>
<tr>
<th>acid</th>
<th>neutral</th>
<th>base</th>
</tr>
</thead>
<tbody>
<tr>
<td>0</td>
<td>7</td>
<td>14</td>
</tr>
</tbody>
</table>

highest \([\text{H}_3\text{O}^+]\) on left
lowest \([\text{H}_3\text{O}^+]\) on right

These are the pH values for several common substances:
2. For the following HCl solutions, list the [H+], [OH-] and pH using the information given.
   a. 0.050 M
   b. pH of 2.73
   c. 1.00 M
   d. pH of 10.4

In basic solutions, the H$_3$O$^+$ ion concentration is very low and we may want to describe the solution in terms of OH$^-$ ion concentration.

Using the same log scale,

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

3. Calculate the pOH for the solutions in sample problem 3 above

Using the log scale for ion concentration, we should recognize that for the expression:

$$K_w = [H_3O^+][OH^-]$$

$$-\log K_w = -\log [H_3O^+] + -\log [OH^-]$$

$$pK_w = pH + pOH = 14.00$$

Therefore, if we know pH, pOH, [H$_3$O$^+$], or [OH$^-$] all other values can be obtained.
4. For the following solutions, calculate pH, pOH, \([H_3O^+], [OH^-]\) and state whether the solution is neutral, basic or acidic.

a. 0.00003 M HCl
b. 1.0 \times 10^{-7} M NaOH
c. Sulfuric acid solution with a pH of 4.8
d. pH 8.92 calcium hydroxide solution
e. 0.00387 M perchloric acid solution