

10.6: The pH of Weak Acid Solutions

For a solution of a strong acid, calculating the $[H_3O^+]$ concentration is simple; because the acid is 100% dissociated, the concentration of hydronium ions is equal to the molar concentration of the strong acid (this is, of course, only true for a monoprotic acid such as HCl or HNO_3 ; for H_2SO_4 , $[H_3O^+] = 2 \times [H_2SO_4]$, etc.). For a weak acid, however, the hydronium ion concentration will be much, much less than the molar concentration of the acid and $[H_3O^+]$ must be calculated using the value of K_a . We can approach this using an *ICE* table, like we did for previous equilibrium problems. If we prepared a solution of acetic acid that was exactly 0.50 M, then *initially* $[CH_3COOH]$ is 0.50 M and both $[CH_3COO^-]$ and $[H_3O^+]$ are zero. A small amount of CH_3COOH will ionize; let's call this x , making the *change* for $[CH_3COOH]$ " $-x$ ", increasing both $[CH_3COO^-]$ and $[H_3O^+]$ by the amount " $+x$ ". Finally, the *equilibrium* concentration of $[CH_3COOH]$ will be $(0.50\text{ M} - x)$ and both $[CH_3COO^-]$ and $[H_3O^+]$ will be x . The completed table is shown below.

	$[CH_3COOH]$	$[CH_3COO^-]$	$[H_3O^+]$
=== Initial ===	0.50 M	0	0
<i>Change</i>	$-x$	$+x$	$+x$
<i>Equilibrium</i>	$0.50\text{ M} - x$	x	x

The expression for K_a for acetic acid is given in [equation in section 10.5](#). Substituting for our equilibrium values:

$$K_a = 1.8 \times 10^{-5} = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]} = \frac{x^2}{0.50 - x}$$

$$x^2 + 9.0 \times 10^{-6}x - 1.8 \times 10^{-5} = 0$$

The above equation is a quadratic equation and we could solve it using the standard quadratic formula. This is not necessary, however, because acetic acid is a *weak acid* and by definition, very little of the dissociated form will exist in solution, making the quantity x very, very small. If x is much, much less than 0.50 M (our initial concentration of acetic acid), then $(0.50\text{ M} - x) \approx 0.50\text{ M}$ and the equation simplifies to:

$$K_a = 1.8 \times 10^{-5} = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]} = \frac{x^2}{0.50}$$

$$x = [H_3O^+] = \sqrt{(1.8 \times 10^{-5}) \times 0.50} = 3.0 \times 10^{-3}\text{ M}$$

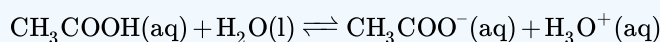
We can test our assumption by substituting for x ; $(0.50 - 0.0030) = 0.497$, which rounds to 0.50 to two significant figures. Because the concentration of hydronium ion is very small for a weak acid, for most typical solutions, the concentration of hydronium ion can be estimated simply as:

$$[H_3O^+] = \sqrt{(K_a \times C_0)}$$

where C_0 is the initial molar concentration of the weak acid.

? Exercise 10.6.1

- Nitrous acid (HNO_2) is a weak acid with a K_a of 4.3×10^{-4} . Estimate the hydronium ion concentration and the pH for a 0.50 M solution of nitrous acid in distilled water.
- Acetic acid is a weak acid with $K_a = 1.8 \times 10^{-5}$. For a solution of acetic acid in water, the $[H_3O^+]$ is found to be $4.2 \times 10^{-3}\text{ M}$. What is the concentration of *unionized* acetic acid in this solution?



- A solution is prepared in which acetic acid is 0.700 M and its conjugate base, acetate anion is 0.600 M. As shown above, the K_a of acetic acid is 1.8×10^{-5} ; what will the pH of this solution be?
- What concentration of the weak acid, acetic acid ($K_a = 1.8 \times 10^{-5}$) must you have in pure water in order for the final pH to be 2.38?

This page titled [10.6: The pH of Weak Acid Solutions](#) is shared under a [CC BY-SA 4.0](#) license and was authored, remixed, and/or curated by [Paul R. Young \(ChemistryOnline.com\)](#) via [source content](#) that was edited to the style and standards of the LibreTexts platform.