

## 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<sub>3</sub>; for H<sub>2</sub>SO<sub>4</sub>,  $[H_3O^+] = 2 \times [H_2SO_4]$ , 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:

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

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

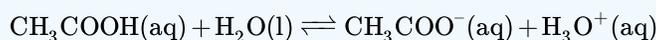
We can test our assumption by substituting for  $x$ ;  $(0.50 - 0.0030) = 0.497$ , which rounds to 0.50 to two significant figures. Because the concentration of hydronium ion is very small for a weak acid, for most typical solutions, the concentration of hydronium ion can be estimated simply as:

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

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

### ? Exercise 10.6.1

- Nitrous acid (HNO<sub>2</sub>) is a weak acid with a  $K_a$  of  $4.3 \times 10^{-4}$ . 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  $4.2 \times 10^{-3}$  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  $1.8 \times 10^{-5}$ ; 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?

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