- **How Strong is a Weak Acid?**
  - The equation for dissociation: HA (*aq*) + $H_2O(l) \rightleftharpoons H_3O^+(aq) + A^-(aq)$
  - The equilibrium constant for this reaction is called the **acid-dissociation constant**, K_a: $K_a = [H_3O^+][A^-] / [HA]$

- **Calculating K_a from the pH**
  - The pH of a 0.10 *M* solution of formic acid, HCOOH, at 25 °C is 2.38. Calculate K_a for formic acid at this temperature.
    - $K_a = [H_3O^+][HCOO^-]/[HCOOH]$
    - Assume that $[H_3O^+] = [HCOO^-]$, $[H_3O^+] = = 10^{-2.38} = 4.2 \times 10^{-3}$ M

	[HCOOH], *M*	$[H_3O^+]$, *M*	$[HCOO^-]$, *M*
I	0.10	0	0
C	-4.2×10^{-3}	$+4.2 \times 10^{-3}$	$+4.2 \times 10^{-3}$
E	$0.10 - 4.2 \times 10^{-3} = 0.0958$	4.2×10^{-3}	4.2×10^{-3}

    - $K_a = [4.2 \times 10^{-3}][4.2 \times 10^{-3}]/[0.0958] = 1.8 \times 10^{-4}$

- **Calculating Percent Ionization**
  - Percent ionization = $([H_3O^+]_{eq} / [HA]_{initial}) \times 100$
  - e.g. 1
    - $[H_3O^+]_{eq} = 4.2 \times 10^{-3}$ *M*
    - $[HCOOH]_{initial} = 0.10$ *M*
    - Percent ionization = $(4.2 \times 10^{-3} / 0.10) \times 100 = 4.2\%$
  - e.g. 2
    - In 0.100 M HF, the percent ionization is 8.1%. Calculate Ka.
      - $HF (aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + F^-(aq)$
      - $Ka = [H_3O^+][F^-] / [HF]$
      - Amount dissociated: $(8.1/100) * 0.100M = 8.1*10^{-3}M = [H_3O^+][F^-]$
      - $Ka = (8.1*10^{-3})^2 / (0.100-8.1*10^{-3}) = 7.14*10^{-4}$

- **Calculating pH Using K_a**
  - Procedure:
    - Write the equilibrium constant expression from the ionization equilibrium.
    - Set up an ICE table to determine equilibrium concentrations as a function of *x*.
    - Substitute equilibrium concentrations into the expression for K_a and solve for *x*.
  - e.g. What is the pH of 0.00100M acetic acid? $K_a = 1.8 * 10^{-5}$
    - $HA(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + A^-(aq)$
      $K_a = [H_3O^+][A^-] / [HA]$
      $K_a = x^2 / 0.001\text{-}x$ (ICE table)
        - option 1: quadratic function
        - option 2: assume $x <<< 0.00100$, $x^2/0.001 = 1.8*10^{-5}$
          - $x = 1.34 \times 10^{-4}$, so pH = 3.873
            - Handy rule of thumb: If $[HA]_o >= 100Ka$, assume $[HA]_{eq} = [HA]_o$
            - 5% rule: $x / [HA]_o * 100 < 5\%$
              $x / [HA] * 100 = 13\% > 5\%$ **(NO)**
        - option 3: successive approximation
          - First approx.: $x = 1.34 \times 10^{-4}$ $(x^2/0.001)$
          - Second approx.: $x = 1.25 \times 10^{-4}$ $(x^2/ (0.001\text{-}1.34 \times 10^{-4})$
          - Third approx.: $x = 1.26 \times 10^{-4}$ $(x^2/(0.001\text{-}1.25 \times 10^{-4})$
          - **$x = 1.26 \times 10^{-4}$, pH = 3.900**

- **Polyprotic Acids**
  - Polyprotic acids have more than one acidic proton.
  - Generally, $K_{a1} >> K_{a2} >> K_{a3}$ so you can ignore the 2nd (and 3rd) ionizations.