How to determine the pH of a solution:
Use these steps to determine the pH of almost any solution (we will use this over and over...know this one!) Let’s use 1.0 M acetic acid for this example.

1. List the major species that are in solution:
   - Acetic acid is a weak acid, so will not fully dissociate (K < 1), acetic acid is major
   - Water is also present
   Our two major species are acetic acid and water.

2. Pick the strongest acid (or base), write out a chemical reaction that shows how H⁺ (or OH⁻) is formed
   - Acetic acid is a stronger acid than water. If will dissociate to give H⁺ in the following general fashion:
     \[ \text{HA (aq)} + \text{H}_2\text{O (l)} \rightleftharpoons \text{H}^+ \text{(aq)} + \text{A}^- \text{(aq)} \]

3. Determine how much H⁺ (or OH⁻) will be formed
   Using the reaction we have just described, we will use our old friend the ICE table:
   
   \[
   \begin{array}{cccc}
   & \text{HA (aq)} & \text{H}_2\text{O (l)} & \text{H}^+ \text{(aq)} & \text{A}^- \text{(aq)} \\
   I & 1.0 \text{ M} & ----- & 0 & 0 \\
   C & -x & ----- & +x & +x \\
   E & 1.0 \text{ M} - x & ----- & +x & +x \\
   \end{array}
   \]

   From this ICE table we can write a \( K_a \) expression of \( K_a = \frac{(x)(x)}{1.0-x} \). If we had a numerical value for \( K_a \) we could use the quadratic equation to solve for \( x \). \( K_a \) in this example is 1.8 x 10⁻⁵ which tells us that equilibrium is far to the left (very little of the acid will dissociate to \( H^+ \) and \( A^- \)). For this reason we are going to make an approximation: 1.0 M – \( x \) is essentially equal to 1.0 M. This simplifies the \( K_a \) expression to: \( K_a = \frac{(x)(x)}{1.0} \) which will allow us to simply solve for \( x = 0.00424 \). We must check this assumption to make sure it is valid, we will cover that more in depth during lecture. In our ICE table \( x \) is equal to the \([H^+]\).

4. Solve for pH (pH = -log[H⁺])
   \[ \text{pH} = - \log[0.00424] = 2.37 \]

**Solutions with more than one acid present**
Most of the time this situation actually isn’t too bad. The key is to look at the \( K_a \) values. If the \( K_a \) values are quite different (a few orders of magnitude at least) we can assume that the acidic protons will be removed one at a time. For example, lactic acid has a \( K_a \) of 1.4 x 10⁻⁴ while HOCl has a \( K_a \) of 3.5 x 10⁻⁸. The lactic acid proton will be completely removed before the HOCl proton will dissociate.

**Percent dissociation:**

\[
\% \text{ dissociation} = \frac{[H^+]}{[HA]} \times 100
\]

If we look at the concentration of the \( H^+ \) that has been formed during the dissociation process we can determine \% dissociation.