At this point in time we’ve look at the pH of weak acids in solution. We’ve also looked at how salts can change the pH of a solution. But what happens when we have a weak acid and a related salt in solution? Consider:

\[
\text{HA (aq) + H}_2\text{O (l)} \rightleftharpoons \text{H}_3\text{O}^+ (aq) + \text{A}^- (aq)
\]

where \( K_a = \frac{[\text{A}^-][\text{H}_3\text{O}^+]}{[\text{HA}]} \). We know from our discussion on weak acids that if HA is a weak acid \( K_a < 1 \) and the equilibrium lies to the left side of the equation. However, a significant amount of \( \text{H}^+ \) will still form, so we will have an acidic solution. If we add the salt of the conjugate bases (say NaA for this example), we will be adding \( \text{Na}^+ \) ions and \( \text{A}^- \) ions to the solution. The \( \text{Na}^+ \) ions will not change the pH (cation of a strong base), but the \( \text{A}^- \) ion is a base. How will this change the equilibrium? If we think of this in terms of Le Chatlier, we know increasing the concentration of \( \text{A}^- \) will shift equilibrium to the left: less \( \text{H}^+ \) will be formed which results in a lower \% dissociation and a less acidic pH. We cause this shift in the equilibrium the common ion effect. The common ion effect takes place when we have a weak acid and it’s conjugate base, or a weak base and it’s conjugate acid.

The common ion effect will result in a change in the pH of the solution we have made (adding the conjugate base of a weak acid to the solution will increase the pH). But the common ion effect has another important consequence.

Consider the following standard situation again:

\[
\text{HA (aq) + H}_2\text{O (l)} \rightleftharpoons \text{H}_3\text{O}^+ (aq) + \text{A}^- (aq)
\]

If I add only HA to a beaker of water, I will end up with \( \text{H}^+ \), \( \text{A}^- \) and HA. \( \text{H}^+ \) and \( \text{A}^- \) will be present in a relatively small amount. If I add \( \text{OH}^- \) to the solution, the \( \text{OH}^- \) will react with \( \text{H}^+ \) and the equilibrium will be shifted quite dramatically. We can think about this as having a great deal of concentration on the reactants side, so it is very easy to shift equilibrium to the products (a bit of a stretch, but that may be a useful mental picture). When we have a large shift in the equilibrium we expect a relatively large change in the pH.

If I add HA and \( \text{A}^- \) to a beaker of water I have large amounts of HA and \( \text{A}^- \) in solution. If I add \( \text{OH}^- \) to the solution, \( \text{OH}^- \) will still react with the HA, but the product of that reaction will be \( \text{A}^- \) and water. We already have a relatively large amount of \( \text{A}^- \) in solution, so there is less of a driving force to form more \( \text{A}^- \). Equilibrium will still shift to the products, but the shift will be much smaller which should mean a smaller change in pH.

When the common ion effect is present, we have formed a buffered solution, a solution that will resist pH change.