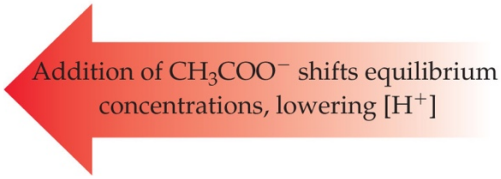
- **Common Ion Effect**
  - **LeChâtelier's Principle**: The shift in equilibrium that occurs by adding a compound that has an ion in common with the dissolved substance.
  - The presence of a common ion suppresses the ionization of a weak acid or base.
    - E.g. 1 Assume we have a solution of acetic acid happily at equilibrium $CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$. What happens if we add some solid sodium acetate? $NaCH_3COO(aq) \rightarrow Na^+(aq) + CH_3COO^-(aq)$



    - E.g. 2 Assume we have a solution of ammonia happily at equilibrium. $NH_3(aq) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$ What happens if we add some solid ammonium chloride? $NH_4Cl(aq) \rightarrow NH_4^+(aq) + Cl^-(aq)$
      - Addition of NH_4^+ shifts reaction left, and the solution is LESS BASIC.
  - **Calculation**
    - What is the pH of 1.0 L of a 0.30 M acetic acid solution *before* and *after* adding 0.30 mol sodium acetate?
      - Before: $K_a = 1.8 \times 10^{-5} = \frac{[H^+][CH_3COO^-]}{[CH_3COOH]} = \frac{x^2}{0.30 - x}$

        $x = 2.3 \times 10^{-3}$ M = $[H^+]$, pH = 2.63 (Ignoring *x* in denominator)
      - After:

	$[CH_3COOH]$ (*M*)	**$[H^+]$ (*M*)**	**$[CH_3COO^-]$ (*M*)**
Initial (*M*)	**0.30**	**0**	**0.30**
Change (*M*)	**–*x***	**+*x***	**+*x***
Equilibrium (*M*)	**0.30 – *x***	***x***	**0.30 + *x***

      - $K_a = 1.8 \times 10^{-5} = \frac{[H^+][CH_3COO^-]}{[CH_3COOH]} = \frac{x\,(0.30 + x)}{0.30 - x}$

        $x = 1.8 \times 10^{-5}$ = $[H^+]$, pH = 4.74

- **Acid-Base Buffers**
  - A buffer is a mixture of a weak acid and its conjugate base.
  - Adding a small amount of H^+ or OH^- reacts only slightly with one component of the buffer, so the pH doesn't change much.
  - $HA \rightleftharpoons H^+(aq) + A^-(aq)$
    1. Add OH^-, react with HA
       $HA + OH^- \rightleftharpoons H_2O + A^-$ (neutralization)
    2. Add H^+, react with A^-
       $H^+ + A^- \rightleftharpoons HA$ (remove H^+ from the solution)

- **Henderson–Hasselbalch equation**
  - $$-\log[H^+] = -\log K_a + \log \frac{[A^-]}{[HA]}$$
    - Take –log of both sides: $K_a = [H^+][A^-] / [HA]$
  - 1. To calculate the pH of a solution when [HA], $[A^-]$, and pKa are known.
  - 2. To calculate the pKa of any weak acid when [HA], $[A^-]$, and pH are known.
  - 3. To calculate the ratio of HA to A^- needed to construct a buffer system at a specific pH.