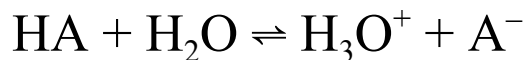


Common Ion Effect Acid + Conjugate Base

For moderately concentrated solutions of a weak acid, HA, and significant added amounts of its conjugate base, A⁻:



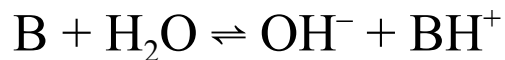
$$[\text{HA}] \approx C_{\text{HA}} \qquad [\text{A}^-] \approx C_{\text{A}^-}$$

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]} \approx \frac{[\text{H}_3\text{O}^+]C_{\text{A}^-}}{C_{\text{HA}}}$$

$$[\text{H}_3\text{O}^+] \approx K_a \times \left(\frac{C_{\text{HA}}}{C_{\text{A}^-}} \right)$$

Common Ion Effect Base + Conjugate Acid

For moderately concentrated solutions of a weak base, B, and significant added amounts of its conjugate acid, BH⁺:



$$[\text{B}] \approx C_{\text{B}} \qquad [\text{BH}^+] \approx C_{\text{BH}^+}$$

$$K_b = \frac{[\text{OH}^-][\text{BH}^+]}{[\text{B}]} \approx \frac{[\text{OH}^-]C_{\text{BH}^+}}{C_{\text{B}}}$$

$$[\text{OH}^-] \approx K_b \times \left(\frac{C_{\text{B}}}{C_{\text{BH}^+}} \right)$$

Henderson-Hasselbalch Equations

$$\text{pH} = \text{p}K_a + \log\left(\frac{C_{A^-}}{C_{\text{HA}}}\right)$$

$$\text{pOH} = \text{p}K_b + \log\left(\frac{C_{\text{BH}^+}}{C_{\text{B}}}\right)$$

Volume Is Irrelevant

- ☞ For a solution of an acid to which significant amounts of its conjugate base have been added so that *both* have moderate concentrations ($\gg 10^{-5}$ M), the volume of the solution does not affect the pH.

If V is the volume of a solution made by adding significant amounts of both HA and A^- :

$$C_{\text{HA}} = \text{mol HA}/V$$

$$C_{A^-} = \text{mol } A^-/V$$

Substituting into K_a :

$$K_a \approx \frac{[\text{H}_3\text{O}^+] \left(\frac{\text{mol } A^-}{V} \right)}{\frac{\text{mol HA}}{V}} = \frac{[\text{H}_3\text{O}^+] (\text{mol } A^-)}{\text{mol HA}}$$

$$[\text{H}_3\text{O}^+] \approx K_a \times \left(\frac{\text{mol HA}}{\text{mol } A^-} \right)$$

- ☞ This is also true for a solution of a base to which significant amounts of its conjugate acid have been added so that *both* have moderate concentrations ($\gg 10^{-5}$ M).

Equimolar Solutions of a Conjugate Pair

If $C_{\text{HA}} = C_{\text{A}^-}$, then

$$K_a = \frac{[\text{H}_3\text{O}^+]\text{C}_{\text{A}^-}}{C_{\text{HA}}} = [\text{H}_3\text{O}^+]$$

And $\text{pH} = \text{p}K_a$

Likewise, if $C_{\text{B}} = C_{\text{BH}^+}$, then

$$K_b = \frac{[\text{OH}^-]\text{C}_{\text{BH}^+}}{C_{\text{B}}} = [\text{OH}^-]$$

And $\text{pOH} = \text{p}K_b$

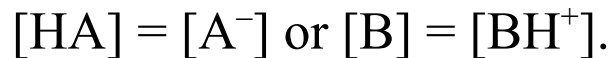
Buffer Solutions

- ☞ A buffer solution is an equilibrium mixture of a weak acid and its conjugate base, or a weak base and its conjugate acid, with both members of the conjugate pair present in moderate concentrations.

- ☞ The essential properties of a buffer solution are:
 1. pH does not change with moderate dilution.
 2. pH does not change significantly with small additions of acid or base.

The Ideal Buffer

① Make



Then,

$$\text{pH} = \text{p}K_a \text{ and } \text{pOH} = \text{p}K_b$$

✓ Response will be equally good to added H_3O^+ or OH^- .

② Make concentrations high.

✓ Then, the ratio $[\text{A}^-]/[\text{HA}]$ or $[\text{BH}^+]/[\text{B}]$ will not change *significantly* with small added amounts of H_3O^+ or OH^- .

Nearly Ideal Buffer

✓ To make a buffer with a desired pH:

① Try to choose a conjugate pair whose $pK_a \approx \text{pH}$.

② Adjust the ratio $[A^-]/[HA]$ or $[BH^+]/[B]$ to achieve the desired pH.

③ The practical limit on the choice of conjugate acid-base pair is

$$\text{pH} \approx pK_a \pm 1$$

✓ Otherwise, one member of the pair is present in too small an amount to make an effective buffer.