Formation of Buffers
Buffer formed using a weak acid and a salt containing its conjugate base

A buffer can be formed by the addition of a salt containing the conjugate base to a solution of a weak acid.

**Weak acid solution**

\[ HF + H_2O \leftrightarrow H_3O^+ + F^- \]

**Weak acid and conjugate base**

Add NaF to HF
Example #1

Calculate the mass of NaNO₂, in grams, which must be added to 25 mL of 0.050 M HNO₂ (Kₐ = 4.5 x 10⁻⁴) solution to prepare a buffer with a pH of 3.50. Assume volume change by addition of the solid is negligible.

- The resultant solution is a buffer, the best approach is to apply the Henderson-Hasselbalch equation.

  \[ \text{pH} = \text{pK}_a + \log \left( \frac{[A^-]}{[HA]} \right) \]

  \[ \text{pK}_a = -\log (K_a) = -\log (4.5 \times 10^{-4}) = 3.35 \]

  \[ 3.50 = 3.35 + \log \left( \frac{[A^-]}{0.050} \right) \]

  \[ 0.15 = \log \left( \frac{[A^-]}{0.050} \right) \quad 1.4 = \frac{[A^-]}{0.050} \]

  \[ [A^-] = 0.070 \text{M} \]

- Mass of A⁻ can be calculated.

  \[ 0.070 \text{ mol/L} \times 0.025 \text{ L} \times \frac{69.00 \text{ g NaNO}_2}{1 \text{ mol}} = 0.12 \text{ g NaNO}_2 \]
Buffer formed using a weak acid and a strong base during partial neutralization

A buffer can be formed by neutralization because conjugate base is produced when the neutralization reaction occurs via the reaction below.

\[ \text{HA} + \text{OH}^- \rightarrow \text{A}^- + \text{H}_2\text{O} \]

Weak Acid Solution
\[ \text{HF} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{F}^- \]

Weak Acid and Conjugate Base
\[ \text{HF} + \text{OH}^- \rightarrow \text{F}^- + \text{H}_2\text{O} \]
Example #2

Calculate the volume, in mL, of 0.050 M NaOH solution which must be added to 30. mL of 0.040 M \( \text{HC}_2\text{H}_3\text{O}_2 \) \((K_a = 1.8 \times 10^{-5})\) solution to produce a buffer with a pH of 4.80. Assume volumes are additive.

- The resultant solution is a buffer, the best approach is to apply the Henderson-Hasselbalch equation.
  \[\text{pH} = pK_a + \log \left( \frac{[A^-]}{[HA]} \right)\]
- Volumes are additive so working with moles as opposed to molarity will provide simplicity
- Moles of \( \text{HC}_2\text{H}_3\text{O}_2 \) = 0.030 L \( \times \frac{0.040 \text{ moles}}{1 \text{ Liter}} \) = 0.0012 moles \( \text{HC}_2\text{H}_3\text{O}_2 \)
- \( pK_a = -\log (K_a) = -\log (1.8 \times 10^{-5}) = 4.74 \)
- Moles of acid decrease as moles of base are added according to the balanced equation for the reaction \( \text{HC}_2\text{H}_3\text{O}_2 + \text{OH}^- \rightarrow \text{C}_2\text{H}_3\text{O}_2^- + \text{H}_2\text{O} \). Moles of acid in the buffer = initial moles of acid – moles of base added.
  \[4.80 = 4.74 + \log \left( \frac{A^-}{0.0012-A^-} \right)\]
  \[0.06 = \log \left( \frac{A^-}{0.0012-A^-} \right)\]
  \[1.1 = \left( \frac{A^-}{0.0012-A^-} \right)\]
  \[1.1 (0.0012 - A^-) = A^-\]
  \[A^- = 0.00062 \text{ moles A}^- = \text{moles OH}^-\]
- 0.00062 moles \( \text{OH}^- \times \frac{1 \text{ L}}{.050 \text{ moles}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 12 \text{ mL} \]
Buffer formed using a weak base and a salt containing its conjugate acid

A buffer can be formed when a weak base solution has a salt containing its conjugate acid added.

Weak Base Solution
\[ \text{NH}_3 + \text{H}_2\text{O} \leftrightarrow \text{NH}_4^+ + \text{OH}^- \]

Weak Base and Conjugate Acid
Add \text{NH}_4\text{Cl} to \text{NH}_3
Example #3

Calculate the mass of CH₃NH₃Cl, in grams, which must be added to 535 mL of 0.15 M CH₃NH₂ (Kₐ = 4.4 x 10⁻⁴) solution to produce a pH of 11.10. Assume volume change by addition of the solid is negligible.

- The resultant solution is a buffer, the best approach is to apply the Henderson-Hasselbalch equation.

- pH = pKₐ + log \( \frac{[B]}{[HB^+]} \)

- Kₐ can be found by using Kₐ. Kₐ x Kₐ = Kₐ

- pKₐ = -log (Kₐ) = -log (2.3 x 10⁻¹¹) = 10.64

- pH = pKₐ + log \( \frac{[B]}{[HB^+]} \)

- 0.052 M x 0.535 L x \( \frac{67.53 \text{ g}}{1 \text{ mol}} \) = 1.9 g CH₃NH₃Cl
Buffer formed using a weak base and a strong acid during partial neutralization

A buffer can be formed by neutralization because conjugate acid is produced when the neutralization reaction occurs via the reaction below.

\[ B + H^+ \rightarrow HB^+ \]

Weak Base Solution
\[ \text{NH}_3 + \text{H}_2\text{O} \leftrightarrow \text{NH}_4^+ + \text{OH}^- \]

Weak Base and Conjugate acid
\[ \text{NH}_3 + \text{H}^+ \leftrightarrow \text{NH}_4^+ \]
Example #4

Determine the volume, in mL, of 0.060 M HCl which must be added to 40. mL of 0.50 M NH₃ solution to produce a buffer with a pH of 9.00. Assume the volumes are additive. (Kᵦ = 1.8 x 10⁻⁵)

The resultant solution is a buffer, the best approach is to apply the Henderson-Hasselbalch equation.

pH = pKᵦ + log \( \frac{[B]}{[HB^+]\} \)

Volumes are additive so working with moles as opposed to molarity will provide simplicity

Moles of NH₃ = 0.040 L x \( \frac{0.050 \text{ moles}}{1 \text{ L}} \) = 0.0020 moles NH₃

Kᵦ can be found by using Kᵦ. \( Kᵦ \times Kᵦ = K_w \) \( Kᵦ = \frac{K_w}{Kᵦ} \) \( Kᵦ = \frac{1 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10} \)

pKᵦ = -log (Kᵦ) = -log (5.6 x 10⁻¹⁰) = 9.25

Moles of base decrease as moles of acid are added according to the balanced equation for the reaction

\( \text{NH}_3 + \text{H}^+ \rightarrow \text{NH}_4^+ \) moles of acid in the base = initial moles of base – moles of acid added

9.00 = 9.26 – log \( \frac{0.0020-\text{HB}^+}{\text{HB}^+} \) -0.26 = log \( \frac{0.0020-\text{HB}^+}{\text{HB}^+} \) 0.55 = \( \frac{0.0020-\text{HB}^+}{\text{HB}^+} \)

\( \text{HB}^+ = 0.0013 \text{ moles of HB}^+ = 0.0013 \text{ moles of HCl} \times \frac{1 \text{ L}}{0.060 \text{ mol HCl}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 22 \text{ mL HCl solution.} \)