

Totally Epic AP Chem Review: Buffers and Titrations in a Day!

Buffer: solutions that resist change in pH

Whenever a weak acid or base is present with its conjugate salt – YOU HAVE A BUFFER!!!

4 Ways to Make a Buffer		
Components	Generic Form	Ideal Buffer Ratio
weak acid and its conjugate base	HA and A ⁻ or HA and NaA	<u>1:1</u> mole ratio
weak base and its conjugate acid	B and BH ⁺ or B and BHCl	<u>1:1</u> mole ratio
weak acid + strong base (titration)	HA + NaOH → H ₂ O + NaA	<u>1</u> weak acid: <u>0.5</u> strong base mole ratio
weak base + strong acid (titration)	B + HCl → BH ⁺ + Cl ⁻	<u>1</u> weak base: <u>0.5</u> strong acid mole ratio

How does a buffer work?

The acidic species donates protons to resist increases in pH, and the basic species will accept protons to resist decreases in pH.

When preparing a buffer solution, you want:

- $[HA] = [A^-]$ (acid/base ratio \approx 1:1)
- pH of buffer \approx pK_a (of acid form)
→ weak acid K_a with an exponent \approx buffer pH.
- High capacity (lots of acid and base)

Calculating the pH of a Buffer Solution: Henderson-Hasselbach Equation

We can derive an equation that relates the pH of a buffer solution to the initial concentration of the buffer components by rearranging the acid ionization constant expression. This can be written in two different formats:

$$[H_3O^+] = K_a \frac{[HA]}{[A^-]} \quad \text{or} \quad \text{pH} = \text{pK}_a + \log \left(\frac{[A^-]}{[HA]} \right)$$

[HA] = Weak acid or salt of conjugate base

[A⁻] = Weak base or salt of conjugate acid

- **WARNING:** If concentrations of separate solutions are given with volumes and then the two are added together, you must recalculate the "new concentrations" due to dilution.
- **Shortcut!!!** Since $\frac{[Acid]}{[Base]}$ is a ratio in the equations, the amount of moles may be substituted in place of concentration because the final volumes will be the same, and thus cancel out.

Let's Practice!

1. Calculate the pH of a solution containing 0.75 M lactic acid, $\text{HC}_3\text{H}_5\text{O}_3$ ($K_a = 1.4 \times 10^{-4}$) and 0.25 M sodium lactate, $\text{NaC}_3\text{H}_5\text{O}_3$.

$$\text{pH} = \text{p}K_a + \log \frac{[\text{A}^-]}{[\text{HA}]} = \underbrace{-\log(1.4 \times 10^{-4})}_{3.8539} + \underbrace{\log\left(\frac{0.25}{0.75}\right)}_{-0.47712} = \boxed{3.38}$$

* note: $\text{pH} < \text{p}K_a$
since $[\text{HA}] > [\text{A}^-]$!

2. Calculate the pH of a solution prepared by mixing 30.0 mL of 0.300 M acetic acid, $\text{HC}_2\text{H}_3\text{O}_2$, with 20.0 mL of 0.350 M $\text{NaC}_2\text{H}_3\text{O}_2$. The K_a for acetic acid is 1.80×10^{-5} .

$$\text{mol}_{\text{HC}_2\text{H}_3\text{O}_2} = 0.300 \text{ M} \times 0.0300 \text{ L} = 9.00 \times 10^{-3} \text{ mol}$$

$$\text{mol}_{\text{C}_2\text{H}_3\text{O}_2^-} = 0.350 \text{ M} \times 0.0200 \text{ L} = 7.00 \times 10^{-3} \text{ mol}$$

$$\text{pH} = \text{p}K_a + \log \frac{[\text{A}^-]}{[\text{HA}]} = \underbrace{-\log(1.80 \times 10^{-5})}_{4.7447} + \underbrace{\log\left(\frac{7.00 \times 10^{-3}}{9.00 \times 10^{-3}}\right)}_{-0.1091} = \boxed{4.636}$$

Acid	Acid Dissociation Constant, K_a
H_3PO_4	7×10^{-3}
H_2PO_4^-	8×10^{-8} ✓
HPO_4^{2-}	5×10^{-13}

1. On the basis of the information above, a buffer with a $\text{pH} = 9$ can best be made by using

- a. $\text{H}_3\text{PO}_4 + \text{H}_2\text{PO}_4^-$ c. $\text{H}_2\text{PO}_4^- + \text{PO}_4^{3-}$
 (b) $\text{H}_2\text{PO}_4^- + \text{HPO}_4^{2-}$ d. $\text{HPO}_4^{2-} + \text{PO}_4^{3-}$

2. Which of the following changes would affect the pH of a buffer solution?

- X I. Doubling the amount of acid and conjugate base used.
 X II. Doubling the amount of water in the solution.
 ✓ III. Adding a small amount of strong acid or strong base.

no correct answer!

III only!
~~a. II only~~

- b. I and II only c. II and III only d. I, II, and III

3. A buffer solution can be formed by dissolving equal moles of:

- b. HF and NaOH c. CH_3COOH and NaCl
 c. KBr and Na_3PO_4 (d) HF and NaF

4. Which of the following acids would be the best choice to create a buffered solution with a pH of 5?

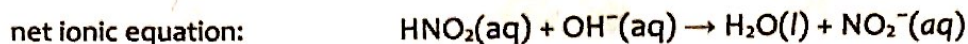
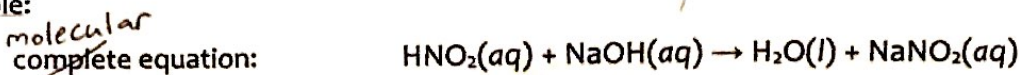
- a. $\text{H}_2\text{C}_2\text{O}_4$ $K_a = 5.9 \times 10^{-2}$ (c) $\text{HC}_2\text{H}_3\text{O}_2$ $K_a = 1.8 \times 10^{-5}$
 b. H_3AsO_4 $K_a = 5.6 \times 10^{-3}$ d. HOCl $K_a = 3.0 \times 10^{-8}$

Net Ionic Equations for Weak + Strong Acid/Base Reactions

Remember, strong acids and bases dissociate 100%, but weak acids and bases do not!

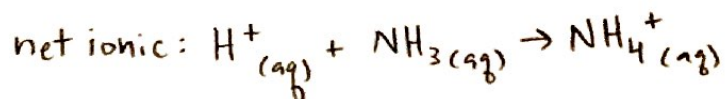
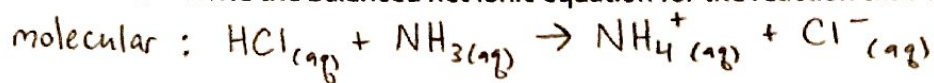
- In a strong/weak net ionic, the only spectator ion that will be removed is the conjugate of the Strong acid or base!

Example:

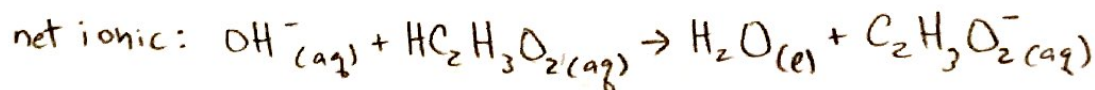
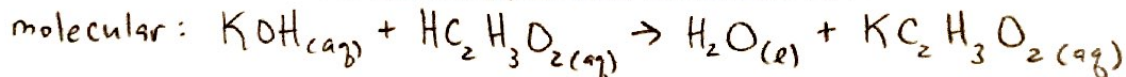


Let's Practice!

- Write the balanced net ionic equation for the reaction that occurs between HCl and NH₃.



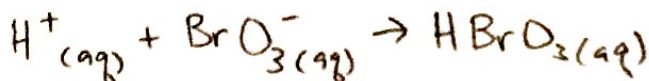
- Write the balanced net ionic equation for the reaction that occurs between KOH and HC₂H₃O₂.



How Does a Buffer Work? Explaining with Reactions.

Example Buffer #1: HBrO₃ and LiBrO₃

- Write the ^{net ionic} equation that represents that reaction that explains why adding a few drops of HBr will not significantly change the pH of the buffer solution:

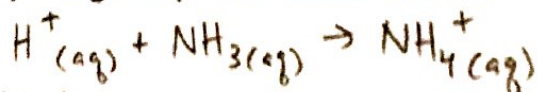


- Write the ^{net ionic} equation that represents that reaction that explains why adding a few drops of NaOH will not significantly change the pH of the buffer solution:



Example Buffer #2: NH₄Cl and NH₃

- Write the ^{net ionic} equation that represents that reaction that explains why adding a few drops of HBr will not significantly change the pH of the buffer solution:



- Write the ^{net ionic} equation that represents that reaction that explains why adding a few drops of KOH will not significantly change the pH of the buffer solution:

