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*	1:1 mole ratio
weak base and its conjugate acid	B and BH^+ **or** B and *BHCl*	1:1 mole ratio
weak acid + strong base (titration)	HA + NaOH → H_2O + NaA	1 weak acid: 0.5 strong base mole ratio
weak base + strong acid (titration)	B + HCl → BH^+ + Cl^-	1 weak base: 0.5 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 ≈ 1:1)
- pH of buffer ≈ pKa (of acid form)
  → weak acid K_a with an exponent ≈ 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:

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

[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!

$[HA]$ $= [C_3H_5O_3^-] = [A^-]$

1. Calculate the pH of a solution containing 0.75 *M* lactic acid, $HC_3H_5O_3$ ($K_a = 1.4 \times 10^{-4}$) and 0.25 *M* sodium lactate, $NaC_3H_5O_3$.

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

★ note: pH < pKa since [HA] > [A⁻]!

2. Calculate the pH of a solution prepared by mixing 30.0 mL of 0.300 *M* acetic acid, $HC_2H_3O_2$, with 20.0 mL of 0.350 *M* $NaC_2H_3O_2$. The K_a for acetic acid is 1.80×10^{-5}.

$$mol_{HC_2H_3O_2} = 0.300 M \times 0.0300 L = 9.00E-3 \, mol$$

$$mol_{C_2H_3O_2^-} = 0.350 M \times 0.0200 L = 7.00E-3 \, mol$$

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

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 pH = 9 can best be made by using
   - a. $H_3PO_4 + H_2PO_4^-$
   - c. $H_2PO_4^- + PO_4^{3-}$
   - [✓ b.] $H_2PO_4^- + HPO_4^{2-}$
   - d. $HPO_4^{2-} + PO_4^{3-}$

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. $H_2C_2O_4$ $K_a = 5.9 \times 10^{-2}$
   - [✓ c.] $HC_2H_3O_2$ $K_a = 1.8 \times 10^{-5}$
   - 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:

molecular ~~complete~~ equation: $HNO_2(aq) + NaOH(aq) \rightarrow H_2O(l) + NaNO_2(aq)$

net ionic equation: $HNO_2(aq) + OH^-(aq) \rightarrow H_2O(l) + NO_2^-(aq)$

Let's Practice!

1. Write the balanced net ionic equation for the reaction that occurs between HCl and NH_3.

molecular : $HCl_{(aq)} + NH_{3(aq)} \rightarrow NH_4^+{}_{(aq)} + Cl^-{}_{(aq)}$

net ionic: $H^+{}_{(aq)} + NH_{3(aq)} \rightarrow NH_4^+{}_{(aq)}$

2. Write the balanced net ionic equation for the reaction that occurs between KOH and $HC_2H_3O_2$.

molecular: $KOH_{(aq)} + HC_2H_3O_{2(aq)} \rightarrow H_2O_{(l)} + KC_2H_3O_{2(aq)}$

net ionic: $OH^-{}_{(aq)} + HC_2H_3O_{2(aq)} \rightarrow H_2O_{(l)} + C_2H_3O_2^-{}_{(aq)}$

How Does a Buffer Work? Explaining with Reactions.

Example Buffer #1: $HBrO_3$ and $LiBrO_3$

1. 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:

$$H^+{}_{(aq)} + BrO_3^-{}_{(aq)} \rightarrow HBrO_{3(aq)}$$

2. 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:

$$OH^-{}_{(aq)} + HBrO_{3(aq)} \rightarrow H_2O_{(l)} + BrO_3^-{}_{(aq)}$$

Example Buffer #2: NH_4Cl and NH_3

1. 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:

$$H^+{}_{(aq)} + NH_{3(aq)} \rightarrow NH_4^+{}_{(aq)}$$

2. 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:

$$OH^-{}_{(aq)} + NH_4^+{}_{(aq)} \rightarrow H_2O_{(l)} + NH_{3(aq)}$$