

CH 302 Unit 2 Day 3

ACID, BASE EQUILIBRIUM

BUFFERS

Acid Base Equilibria Roadmap

1. Define and identify acids and bases
2. Solve for pH and pOH for strong and weak acids/bases
3. Identify and analyze the products of a full neutralization reaction (adding acid to base and vice versa)
4. Identify and analyze the products of a partial neutralization
5. Understand neutralization reactions in the context of titrations and indicators
6. Understand the role of pH in regulating the dominant species of a polyprotic acid

} Last week

} Today

Acid and Base Question Types (Simplified)

- Strong acid, strong base questions
 - Simple relationships converting $[H_3O^+]$ and $[OH^-]$ to pH, pOH
- Weak acid, weak base questions
 - Approximations or quadratic formula (if necessary) with your general formula:
 - $HA \rightleftharpoons H^+ + A^-$ OR $B \rightleftharpoons BH^+ + OH^-$
 - Solve for $[H^+]$ using K_a Solve for $[OH^-]$ using K_b
- **Buffer questions**
 - Partial neutralization of a weak acid and its salt (conjugate base) ; weak base and its salt (conjugate acid)
 - Solve for pH, pOH using Henderson-Hasselbalch equation
- **Neutralization reactions and titration experiments**

Neutralization Reactions: Salts

- The product of a neutralization reaction is a salt. In acid/base chemistry, your salt can be neutral, acidic, or basic depending on the reaction.

1. **GENERIC REACTION** (very helpful):



2. **Strong Acid, Strong Base: results in a neutral salt**



3. **Strong Acid to weak base: results in an acidic salt**



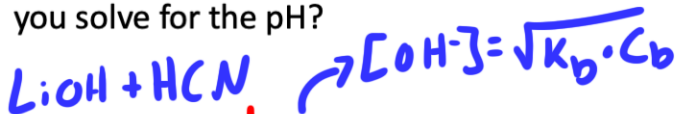
4. **Strong base to weak acid: results in a basic salt**



$\text{H}_3\text{O}^+ \text{Cl}^-$
↑

Neutralization Reactions: Salts

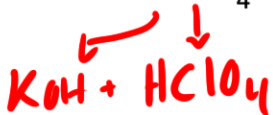
Identify whether the following solutions will be acidic, basic, or neutral. How would you solve for the pH?



0.3 M LiCN **basic**



0.08 M KClO_4 **neutral**



0.05 M $\text{CH}_3\text{CH}_2\text{NH}_3^+\text{Cl}^-$



$[\text{H}^+] = \sqrt{K_a \cdot C_b}$

0.1 M KF \rightarrow **weak base**



$K_b = \text{CH}_3\text{CH}_2\text{NH}_2$
 $K_a = \frac{K_w}{K_b}$

Type II RICE table
moles

$$[H^+] = \sqrt{\frac{K_w}{K_b} \cdot 0.5M}$$

$$K_b \text{ NH}_3 = 1.8 \times 10^{-5}$$

One step harder...

An ammonium perchlorate solution is made by combining 200 mL 1.0 M perchloric acid (HClO₄) and 200 mL 1.0 M ammonia (NH₃). What is the pH?



R	HClO ₄ + NH ₃ \rightleftharpoons NH ₄ ClO ₄ + H ₂ O		
I	0.2	0.2	∅
C	-0.2	-0.2	+0.2
E	∅	∅	0.2

* moles

$$[H^+] = \sqrt{K_a \cdot C_A}$$

-log ↓
 $C_A = \frac{0.2}{0.4} \rightarrow 0.5M$

pH = 4.78
 $K_A = \frac{K_w}{K_B} \checkmark$

Future note: this question is identical to calculating the buffer zone of a titration experiment

And even harder...

An ammonium perchlorate solution is made by combining 100 mL 1.1 M perchloric acid (HClO₄) and 100 mL 2.0 M ammonia (NH₃). What is the pH?

$$-\log K_a = pK_a$$

$$K_b \text{ NH}_3 = 1.8 \times 10^{-5}$$

$$K_a = \frac{K_w}{K_b}$$

R	$\text{HClO}_4 + \text{NH}_3 \rightleftharpoons \text{NH}_4^+ \text{ClO}_4^- + \text{H}_2\text{O}$		
I	.11	0.20	\emptyset
C	-.11	-.11	+.11
E	\emptyset	.09	.11

$$\text{pH} = \text{p}K_a + \log \frac{A^-}{HA}$$

$$9.17 = 9.255 + \log \frac{0.09}{0.11}$$

Future note: this question is identical to calculating the pH at the equivalence point of a titration

A^- = Conj base, HA = weak acid
= base = conj. acid

Buffers: Conceptual, Calculation

$\frac{\text{moles } A^-}{\text{moles HA}}$

- The purpose of a buffer is to resist changes in pH.
- Here's the idea:
 - If you add 0.1 mole of NaOH to pure water, you are adding 0.1 mole of OH^- . This results in a pretty big change in pH
 - If you add 0.1 mole of NaOH to a solution of acetic acid and sodium acetate, you are just creating 0.1 mole more of sodium acetate. **This barely increases the pH.**

✓ $pH = pK_a + \log\left(\frac{[A^-]}{[HA]}\right)$ ~~$pOH = pK_b + \log\left(\frac{[BH^+]}{[B]}\right)$~~



What is a buffer and what is not:

- A buffer is made of:
 - A weak acid and its salt (conjugate base)
 - A weak base and its salt (conjugate acid)
- A buffer is made by:
 - Mixing a weak acid and a strong base
 - Mixing a weak base and a strong acid

Buffer:

100 mL 0.5 M HNO_2 + 100 mL 0.5 M NaNO_2 ✓

100 mL 0.25 M LiOH + 100 mL 0.5 M HNO_2 ✓

Not a buffer

100 mL 0.5 M LiOH + 100 mL 0.5 M HNO_2
= 100% LiNO_2



Choosing a Buffer

Buffer works +/- 1 pH pt from pKa
1:10 } HA:A⁻
10:1 }

- A buffer only functions in the “Buffer Zone,” which is +/- 1 pH point of the pK_a (for a weak acid buffer) or pK_b (generally the standard for a weak base buffer).
- Buffers are used in the real world (reaction chemistry, physiology, pharmacology, etc.) to maintain a stable pH environment. You choose a buffer with a pK_a closest to the pH environment you want to hold constant.

Example Question: The human bloodstream is held constant by a buffer system at pH = 7.3.

Which of the following buffer systems is likely found in the bloodstream?

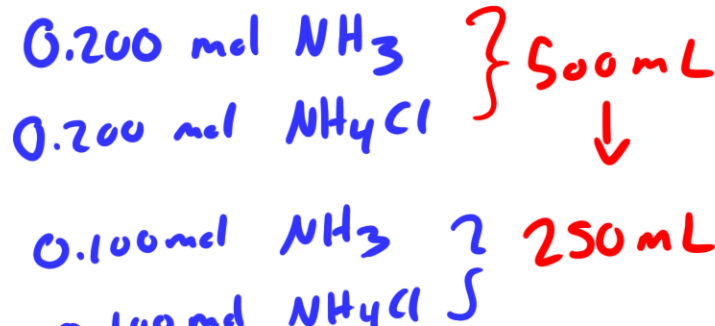
- a. Carbonic Acid, $pK_a = 6.37$
- b. Acetic Acid, $pK_a = 4.75$
- c. Hydrofluoric Acid, $pK_a = 3.14$
- d. Ammonium, $pK_a = 9.21$

Exam Question

A buffer was prepared by mixing 0.200 mole of ammonia ($pK_b = 4.76$) and 0.200 mole of ammonium chloride to form an aqueous solution with a total volume of 500 mL.

50.0 mL of 1.00 M HCl was added to 250 mL of this solution.

What is the pH of this solution?



Type II
RICE
in moles

0.100

R	$\text{HCl} + \text{NH}_3 \rightleftharpoons \text{NH}_4\text{Cl} + \text{H}_2\text{O}$		
I	.05	0.100	0.100
C	-.05	-0.05	+0.05
E	\emptyset	.05	.15

$$\text{pH} = \text{pK}_a + \log \frac{.05}{.15} = 8.76$$

Exam Question

$\rightarrow pK_a = 9.24$

A buffer was prepared by mixing 0.200 mole of ammonia ($pK_b = 4.76$) and 0.200 mole of ammonium chloride to form an aqueous solution with a total volume of 500 mL.

50.0 mL of 1.00 M HCl was added to 250 mL of this solution.

What is the pH of this solution?

$$pH = pK_a + \log\left(\frac{0.100 - 0.05}{0.100 + 0.05}\right)$$
$$9.24 + \log\frac{.05}{.15} = \boxed{8.76}$$

Too-Hard-For-The-Exam Question

A potassium acetate buffer solution is prepared by mixing 100 mL 0.200 M CH_3COOH and 100 mL 0.200 M KCH_3COO .

What volume of 0.125 M KOH is necessary to raise the pH to 5.440? $\text{pK}_a = 4.74$

$$\text{pH} = \text{pK}_a + \log \frac{\text{A}^-}{\text{HA}}$$

↓ ↓

$$5.440 = 4.74 + \log \left(\frac{0.02 + x}{0.02 - x} \right)$$

Algebra

$$5.440 - 4.74 = \frac{0.02 + x}{0.02 - x}$$

10

$$\begin{aligned} & \rightarrow x = 0.0134 \\ & \cdot 0.0134 \text{ moles} \\ & \times \frac{\text{L}}{0.125 \text{ mol}} \\ & \boxed{= 0.107 \text{ L}} \end{aligned}$$