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):**

   Acid + Base ⇌ Salt + H_2O

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

   $HCl + NaOH \rightleftharpoons NaCl + H_2O$ neutral salt

3. **Strong Acid to weak base: results in an acidic salt** HCl, NH_3

   $H_3O^+ Cl^-$

   * $HCl + NH_3 \rightleftharpoons NH_4Cl + H_2O$ weak acid salt

4. **Strong base to weak acid: results in a basic salt** NaOH, CH_3COOH

   $CH_3COOH + NaOH \rightleftharpoons NaCH_3COO + H_2O$ weak base salt

Neutralization Reactions: Salts

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

$LiOH + HCN$ → $[OH^-] = \sqrt{K_b \cdot C_b}$

0.3 M LiCN — basic

$Li^+ CN^-$ — CN^-: conj base

$HCl + CH_3CH_2NH_2$ → $K_b = CH_3CH_2NH_2$; $K_a = \frac{K_w}{K_b}$

0.05 M $CH_3CH_2NH_3^+Cl^-$ — conj. acid

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

0.08 M $KClO_4$ — neutral

$KOH + HClO_4$

0.1 M KF → weak base

KOH HF

Type II RICE table

moles

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

K_b NH_3 = 1.8 x 10^{-5}

One step harder…

An ammonium perchlorate solution is made by combining 200 mL 1.0 M perchloric acid ($HClO_4$) and 200 mL 1.0 M ammonia (NH_3) . What is the pH?

Acid + Base ⇌ Salt + H_2O

R	$HClO_4$ +	NH_3 ⇌	NH_4ClO_4 +	H_2O
I	0.2	0.2	Ø	
C	−0.2	−0.2	+0.2	
E	Ø	Ø	0.2	

* moles

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

$C_A = \frac{0.2}{0.4}$ → 0.5M

$K_A = \frac{K_w}{K_B}$ ✓

−log ↓

pH = 4.78

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

And even harder…

$K_b\ NH_3 = 1.8 \times 10^{-5}$

$K_a = \frac{K_w}{K_b}$ → $-\log K_a = pK_a$

An ammonium perchlorate solution is made by combining 100 mL 1.1 M perchloric acid ($HClO_4$) and 100 mL 2.0 M ammonia (NH_3) . What is the pH?

	$HClO_4$ +	NH_3 ⇌	$NH_4^+ClO_4$ + H_2O
R			
I	.11	0.20	∅
C	−.11	−.11	+.11
E	∅	.09	.11

$pH = pK_a + \log\frac{A^-}{HA}$

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

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

Buffers: Conceptual, Calculation

- **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.**

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

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

moles A^- / moles HA

$\frac{[A^-]}{[HA]}$

Common ion effect → $CH_3COOH \rightleftharpoons CH_3COO^- + H^+$

1 mol 1 mol

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% $LiNO_2$

Buffer works +/- 1 pH pt from pKa

1:10 } HA : A^-
10:1

Choosing a Buffer

- **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, pKa = 6.37
b. Acetic Acid, pKa = 4.75
c. Hydrofluoric Acid, pKa = 3.14
d. Ammonium, pKa = 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?

0.200 mol NH_3
0.200 mol NH_4Cl } 500 mL
↓
0.100 mol NH_3
0.100 mol NH_4Cl } 250 mL

Type II
RICE
in moles

0.100 ...

R	$HCl + NH_3 \rightleftarrows NH_4Cl + H_2O$		
I	.05	0.100	0.100
C	-.05	-0.05	+0.05
E	∅	.05	.15

$$pH = pK_a + \log\frac{.05}{.15} = 8.76$$

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?

$pK_a = 9.24$

$$pH = pK_a + \log\left(\frac{0.100 - 0.05}{0.100 + 0.05}\right)$$

$$9.24 + \log\frac{.05}{.15} = 8.76$$

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? pKa = 4.74

$$pH = pKa + \log \frac{A^-}{HA}$$

$$5.440 = 4.74 + \log\left(\frac{.02 + x}{.02 - x}\right)$$

$$10^{5.440 - 4.74} = \frac{0.2 + x}{0.2 - x}$$

Algebra $\rightarrow x = .0134$

$$.0134 \text{ moles} \times \frac{L}{0.125 \text{ mol}} = 0.107 L$$