CHM 152 Exam 2 Key Diebolt Fall 2017

**Part One: Multiple choice.** (33 pts: 3 pts each)

1. The pH of Budweiser beer is 4.30. The $[OH^-]$ is ___ and the solution would be _____.

   A. $5.0\times10^{-5}$ M, acidic  B. $5.0\times10^{-5}$ M, basic  **C. $2.0\times10^{-10}$ M, acidic**
   D. $2.0\times10^{-10}$ M, basic  E. $2.0\times10^{-11}$ M, acidic
   **pOH = 14 – 4.30 = 9.70**  **$[OH^-] = 10^{-9.70} = 2.0\times10^{-10}$ M, pH < 7 so the beer is acidic**

2. Select the conjugate acid for $HPO_4^{2-}$

   A. $H_3PO_4$  **B. $H_2PO_4^-$** = C.A. (1 more $H^+$)  C. $H^+$  D. $PO_4^{3-}$  E. $OH^-$

3. Select the solution with the **lowest pH**.

   A. 0.10 M $Ba(OH)_2$ **SB**  B. 0.10 M $HNO_2$ **WA**  C. 0.010 M $HClO_4$ **SA but lower [ ]**  **D. 0.10 M HCl** **SA w/ highest [ ]**  E. 0.10 M NaOH **SB**

4. According to the **Bronsted-Lowry theory**, a **base** is

   A. an electron pair donor = **Lewis base**  B. an electron pair acceptor = **Lewis acid**
   C. a proton donor **= BL acid**  **D. a proton acceptor = BL base**
   E. a hydroxide ion donor **= Arrhenius base**

5. Consider the $K_a$ values for the following acids:

   $NH_4^+$  $K_a = 5.6 \times 10^{-10}$
   $HC_3H_3O_2$  $K_a = 5.5 \times 10^{-5}$
   $HCO_2H$  $K_a = 1.8 \times 10^{-4}$
   $HC_6H_5O$  $K_a = 1.6 \times 10^{-10}$ **smallest $K_a$ = strongest acid; thus, strongest conj base**
   $HNO_2$  $K_a = 4.5 \times 10^{-4}$ **largest $K_a$ = strongest acid; thus, weakest conj base**

   Which of the following conjugate bases is the **weakest base**?

   A . $NH_3$  B. $C_3H_3O_2^-$  C. $HCO_2^-$  D. $C_6H_5O^-$  **E. $NO_2^-$**

6. When comparing two different 0.10 M acid solutions, which statement is **True**?

   A. The stronger acid has a lower percent ionization. **F, as strength ↑, % ionization ↑**
   B. The weaker acid has a lower $pK_a$ value. **F, $pK_a$ = - log $K_a$; $pK_a$↑ as $K_a$ ↓ for weaker acid**
   C. The weaker acid has a lower pH reading. **F, it has lower $[H_3O^+]$; thus, higher pH**
   **D. The stronger acid is a better proton donor. T, as strength ↑, acid loses $H^+$ more readily**
   E. The weaker acid has a larger $K_a$ value. **F, it has less $[H_3O^+]$ so smaller $K_a$**

7. What change will be caused by the addition of a small amount of $HI_{(aq)}$ to a buffer solution containing $LiHCO_{2(aq)}$ and $HCO_2H_{(aq)}$?

   **Add $H_3O^+$(HI = SA); neut: $H_3O^+ + HCO_2^-$↓ → $HCO_2H$↑ + $H_2O$**
   A. $[H_3O^+]$ will significantly increase
   B. $[OH^-]$ will significantly increase
   **C. $[HCO_2H]$ will increase and $[HCOO^-]$ will decrease** WA = $HCO_2H$↑, WB = $HCOO^-$↓
   D. $[HCO_2H]$ will decrease and $[HCOO^-]$ will increase
   E. $[HCO_2H]$ and $[HCOO^-]$ will both increase

8. Which of the following combinations could produce an effective **buffer solution**?

   A. $KClO_4$, $HClO_4$ **NS, SA**  **B. HBrO, KBrO** **WA, CB**  C. HI, NaOH **SA SB**  D. NaOH, NaCl **SB, NS**  E. KBr, HBr **NS, SA**

   **Buffers contain: a) weak acid/ salt with conj base; or 2) weak base/salt with conj acid**

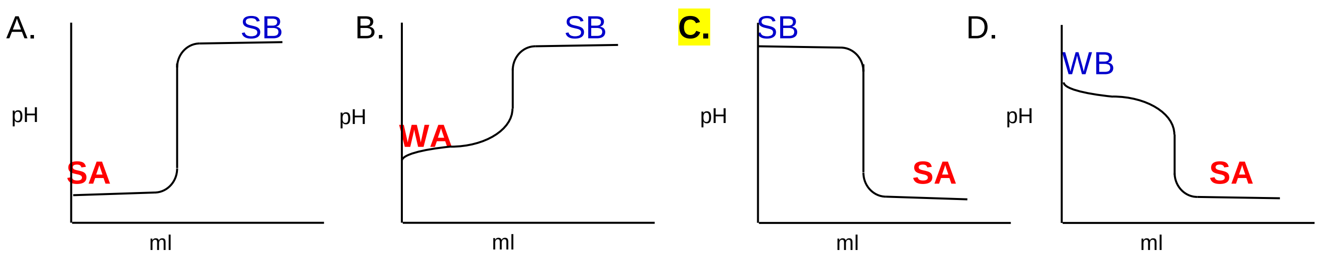
9. When **$HC_2H_3O_3$*(aq)*** (WA) **is titrated by RbOH*(aq)*** (SB), the pH at the equivalence point is $\geq$ 7 because a basic salt has formed. **WA** titrated by SB forms basic salt ($RbC_2H_3O_2$!

A. 7, neutral B. pH < 7, acidic C. pH < 7, basic D. pH > 7, acidic **E. pH > 7, basic**

10. What occurs when $NH_4NO_{3(aq)}$ is added to an $NH_{3(aq)}$ solution? **Added Conj. Acid from acidic salt so pH ↓; thus, $[OH^-]$↓ (adding common ion, $NH_4^+$: $NH_3 + H_2O \leftrightharpoons NH_4^+ + OH^-$)**

**A. pH decreases and $[OH^-]$ decreases**
B. pH decreases and $[OH^-]$ increases
C. pH increases and $[OH^-]$ decreases
D. pH increases and $[OH^-]$ increases
E. the pH and $[OH^-]$ do not change

11. Which plot shows the titration curve expected when **LiOH (SB)** is titrated by **$HClO_{4(aq)}$ (SA)**?




**Part Two. Short Answer.**

1. (6 pts) Predict whether each of the following salt solutions is **acidic, basic or neutral**.

A. $Li_2CO_3$ — N B = basic salt

B. $NH_4ClO_4$ — A N = acidic salt

C. $Sr(NO_3)_2$ — N N = neutral salt

**Part Three. Problems.** Please **SHOW YOUR WORK** for full credit. Use the correct number of **significant figures** for your answers and **circle your final answer** for each problem. (61 pts)

1. Ethylamine, $C_2H_5NH_2$, acts as a **weak base** in its reaction with $H_2O$.

a) Write the hydrolysis reaction that occurs for ethylamine ($C_2H_5NH_2$). (4 pts)

$$C_2H_5NH_2 + H_2O(l) \leftrightharpoons C_2H_5NH_3^+ + OH^-$$

b) A 0.015 M ethylamine solution has a pH of 11.42. Calculate the $K_b$ for ethylamine. (9 pts)

| $C_2H_5NH_2$ | + $H_2O$ $\leftrightharpoons$ | $C_2H_5NH_3^+$ + | $OH^-$ |
|---|---|---|---|
| 0.015 | | 0 | 0 |
| -x | | +x | +x |
| 0.015 - x | | x | x |

pH = 11.42 pOH = 14 – 11.42 = 2.58

$[OH^-] = 10^{-pOH}$ $[OH^-] = 10^{-2.58} = 2.63\times10^{-3}$ M x = $[OH^-] = [C_2H_5NH_3^+] = 2.63\times10^{-3}$ M

$[C_2H_5NH_2] = 0.015 - 2.63\times10^{-3} = 0.0124$M **Don't drop x term, we know x from the pH!**

$$K_b = \frac{[C_2H_5NH_3^+][OH^-]}{[C_2H_5NH_2]} \qquad K_b = \frac{x^2}{0.015 - x} = \frac{(2.63x10^{-3})^2}{0.0124} \qquad K_b = 5.6x10^{-4}$$

c) Calculate the % ionization for this 0.015 M ethylamine solution. (3 pts)

$$\%\text{ionization} = \frac{x}{[C_2H_5NH_2]}\times100\% = \frac{2.63x10^{-3}}{[0.015]}x100\% = 17.5\% = \underline{18\%\text{ ionized}}$$

22 pts

2. Calculate the pH, pOH, $[OH^-]$ and $[H_3O^+]$ for a $2.75\times10^{-2}$ M $Ba(OH)_{2(aq)}$ solution. (8 pts)

$Ba(OH)_2$ is SB Dissociation reaction: $Ba(OH)_2 \rightarrow Ba^{2+} + 2\ OH^-$

$$[OH^-] = \left(\frac{2.75\text{x}10^{-2}\text{ mol } Ba(OH)_2}{\text{L}}\right)\left(\frac{2\text{ mol } OH^-}{1\text{ mol } Ba(OH)_2}\right) = 5.50\times10^{-2}\text{ M } OH^-$$ **(3 sf)**

pOH = - log $5.50\times10^{-2}$ =1.2596 **pOH = 1.260 (3 dp)** pH = 14 – 1.260 **pH = 12.740 (3 dp)**

$[H_3O^+] = 10^{-pH} = 10^{-12.740}$ **$[H_3O^+] = 1.82\text{x}10^{-13}$ M (3 sf)** ☺ **pH is high for a SB**

3. You have 200.0 mL of a buffer solution containing 0.175 M $HCO_2H$ and 0.225 M $NaHCO_2$. What is the pH after 25.0 mL of 0.300 M KOH is added to this buffer solution? For $HCO_2H$, $K_a = 1.8 \times10^{-4}$. (13 pts)

using moles (or mmoles) and HH eq:

moles $OH^-$ = 0.025 L x 0.300 M = 0.0075 moles $OH^-$ (= 7.50 mmoles)

moles $HCO_2H$= 0.200 L x 0.175 M = 0.0350 moles HA (= 35.0 mmoles)

moles $HCO_2^-$ = 0.200 L x 0.225 M = 0.0450 moles $A^-$ (= 45.0 mmoles)

**$OH^-$ will react with the weak acid, $HNO_2$: $[HNO_2]\downarrow$ $[NO_2^-]\uparrow$**

| AB neut. Rxn: | $OH^-$ + | $HCO_2H \rightarrow$ | $HCO_2^-$ | + $H_2O$ |
|---|---|---|---|---|
| Initial | 0.0075 | 0.0350 | 0.0450 | |
| Change | -0.0075 | -0.0075 | +0.0075 | |
| final | 0 | 0.0275 | 0.0525 | |

mol HA (WA) = 0.0275 mol ( = 27.5 mmol) moles $A^-$ (WB) = 0.0525 mol (= 52.5 mmol)

$pH = pK_a + \log [A^-]/[HA]$ pH = - log $1.8\times10^{-4}$ + log (0.0525/0.0275)

pH = 3.745 + 0.281 **pH = 4.03**

- **pH = 4.02** if rounded pH's to 2 sf before adding

Note: Can also solve calculating diluted M's, finding final M's after neutralization, and plugging the final concentrations in to Ka to find x = $[H_3O^+]$ .

21 pts

4. A solution of perchloric acid, $HClO_4(aq)$, is being titrated with LiOH(aq). Calculate the pH after 15.0 mL of 0.200 M LiOH(aq) is added to 30.0 mL of 0.150 M $HClO_4(aq)$. (10 pts)

SA-SB neutralization: $HClO_4 + LiOH \rightarrow LiClO_4 + H_2O$

Net rxn: $H^+ + OH^- \rightarrow H_2O$ (or $H_3O^+ + OH^- \rightarrow 2H_2O$)

moles acid = 30.0 mL x 0.150 M $HClO_4$ = 4.50 mmol $H_3O^+$

moles base = 15.0 mL x 0.200 M LiOH = 3.00 mmol $OH^-$ **= LR**

moles acid > moles base

moles $H_3O^+$ remaining = 4.50 mmol – 3.00 mmol = 1.50 mmol $H_3O^+$

***Change (ICF table) also could have been used to find which is in excess**

M $H_3O^+$ = $\frac{1.50 \text{ mmoles}}{45.0 \text{ mL}}$ = 0.0333 M $H_3O^+$

pH = -log 0.0333 = **1.478** ☺ **pH is low since there is excess SA!**

4. A sample of benzoic acid, $HC_7H_5O_2$, is being titrated with KOH solution. What is the pH after 20.0 mL of 0.250 M KOH has been added to 25.0 mL of 0.200 M $HC_7H_5O_2$? For $HC_7H_5O_2$, $K_a = 6.4 \times 10^{-5}$. (14 pt)

moles $HC_7H_5O_2$ = $25.0 \text{ ml}\left(\frac{1\text{ L}}{1000 \text{ mL}}\right)\left(\frac{0.200 \text{ moles } HC_7H_5O_2}{\text{L}}\right)$ = $5.00 \times 10^{-3}$ moles $HC_7H_5O_2$

moles KOH = $20.0 \text{ ml}\left(\frac{1\text{ L}}{1000 \text{ mL}}\right)\left(\frac{0.250 \text{ moles KOH}}{\text{L}}\right)$ = $5.00 \times 10^{-3}$ moles $OH^-$ (Same # moles!)

| Neutralization rxn | $HC_7H_5O_2$ + | $OH^-$ → | $C_7H_5O_2^-$ | + $H_2O$ |
|---|---|---|---|---|
| Initial moles | 0.00500 | 0.00500 | 0 | |
| Change | -0.00500 | -0.00500 | +0.00500 | |
| Final moles | 0 | 0 | 0.00500 | |

At equivalence point, so only $C_7H_5O_2^-$, a weak base, remains:

***This is a weak base problem so we need to convert moles to M before we plug it into $K_b$. V terms don't cancel since $K_b = x^2/[B]$!**

Divide moles by total V: $[C_7H_5O_2^-]$ = $\frac{5.00 \times 10^{-3} \text{ moles}}{0.0450 \text{ L}}$ = 0.111 M

| WB hydrolysis | $C_7H_5O_2^-$ | + $H_2O$ ⇆ | $HC_7H_5O_2$ + | $OH^-$ |
|---|---|---|---|---|
| I | 0.111 | | 0 | 0 |
| C | -x | | x | x |
| E | 0.111 - x | | x | x |

$K_b$ for $C_7H_5O_2^-$ = $\frac{1 \times 10^{-14}}{6.4 \times 10^{-5}}$ = $1.56 \times 10^{-10}$ $K_b = \frac{[HC_7H_5O_2][OH^-]}{[C_7H_5O_2^-]}$

$1.56 \times 10^{-10} = \frac{x^2}{0.111}$ x = $[OH^-]$ = $\sqrt{0.111(1.56 \times 10^{-10})}$ = $4.16 \times 10^{-6}$

pOH = -log $4.16 \times 10^{-6}$ = 5.38 pH = 14 – 5.38 **pH = 8.62**

☺ **pH > 7 at the equivalence pt for a WA-SB titration**

24 pts