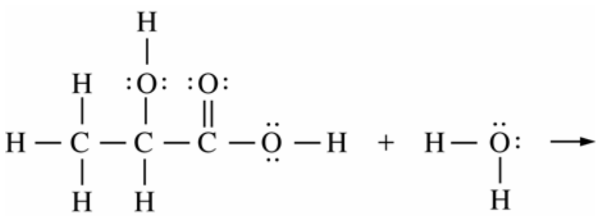
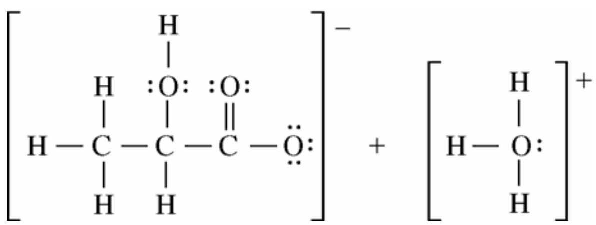
Question 5

Answer the following questions about laboratory solutions involving acids, bases, and buffer solutions.

(a) Lactic acid, $HC_3H_5O_3$, reacts with water to produce an acidic solution. Shown below are the complete Lewis structures of the reactants.



In the space provided above, complete the equation by drawing the complete Lewis structures of the reaction products.

	One point is earned for each correct structure.

(b) Choosing from the chemicals and equipment listed below, describe how to prepare 100.00 mL of a 1.00 *M* aqueous solution of NH_4Cl (molar mass 53.5 g mol^{-1}). Include specific amounts and equipment where appropriate.

$NH_4Cl(s)$	50 mL buret	100 mL graduated cylinder	100 mL pipet
Distilled water	100 mL beaker	100 mL volumetric flask	Balance

mass of NH_4Cl = (0.100 L)(1.00 mol L^{-1})(53.5 g mol^{-1}) = 5.35 g NH_4Cl 1. Measure out 5.35 g NH_4Cl using the balance. 2. Use the 100 mL graduated cylinder to transfer approximately 25 mL of distilled water to the 100 mL volumetric flask. 3. Transfer the 5.35 g NH_4Cl to the 100 mL volumetric flask. 4. Continue to add distilled water to the volumetric flask while swirling the flask to dissolve the NH_4Cl and remove all NH_4Cl particles adhered to the walls. 5. Carefully add distilled water to the 100 mL volumetric flask until the bottom of the meniscus of the solution reaches the etched mark on the flask.	One point is earned for the mass. One point is earned for using a volumetric flask. One point is earned for diluting to the mark.

Question 5 (continued)

(c) Two buffer solutions, each containing acetic acid and sodium acetate, are prepared. A student adds 0.10 mol of HCl to 1.0 L of each of these buffer solutions and to 1.0 L of distilled water. The table below shows the pH measurements made before and after the 0.10 mol of HCl is added.

	pH Before HCl Added	pH After HCl Added
Distilled Water	7.0	1.0
Buffer 1	4.7	2.7
Buffer 2	4.7	4.3

(i) Write the balanced net-ionic equation for the reaction that takes place when the HCl is added to buffer 1 or buffer 2.

$C_2H_3O_2^- + H_3O^+ \rightarrow HC_2H_3O_2 + H_2O$	One point is earned for the equation.

(ii) Explain why the pH of buffer 1 is different from the pH of buffer 2 after 0.10 mol of HCl is added.

Before the HCl was added, each buffer had the same pH and thus had the same $[H^+]$. Because K_a for acetic acid is a constant, the ratio of $[H^+]$ to K_a must also be constant; this means that the ratio of $[HC_2H_3O_2]$ to $[C_2H_3O_2^-]$ is the same for both buffers, as shown by the following equation, derived from the equilibrium-constant expression for the dissociation of acetic acid. $$\frac{[HC_2H_3O_2]}{[C_2H_3O_2^-]} = \frac{[H^+]}{K_a}$$ After the addition of the H^+, the ratio in buffer 1 must have been greater than the corresponding ratio in buffer 2, as evidenced by their respective pH values. Thus a greater proportion of the $C_2H_3O_2^-$ in buffer 1 must have reacted with the added H^+ compared to the proportion that reacted in buffer 2. The difference between these proportions means that the original concentrations of $HC_2H_3O_2$ and $C_2H_3O_2^-$ had to be smaller in buffer 1 than in buffer 2.	One point is earned for a correct answer involving better buffering capacity or relative amount of base (acetate ion).

Question 5 (continued)

(iii) Explain why the pH of buffer 1 is the same as the pH of buffer 2 before 0.10 mol of HCl is added.

Both buffer solutions have the same acid to conjugate-base mole ratio in the formula below. $$[H^+] = K_a \frac{[HC_2H_3O_2]}{[C_2H_3O_2^-]}$$ Therefore, the buffers have the same $[H^+]$ and pH.	One point is earned for the correct answer involving ratio of acid to base in the buffer.

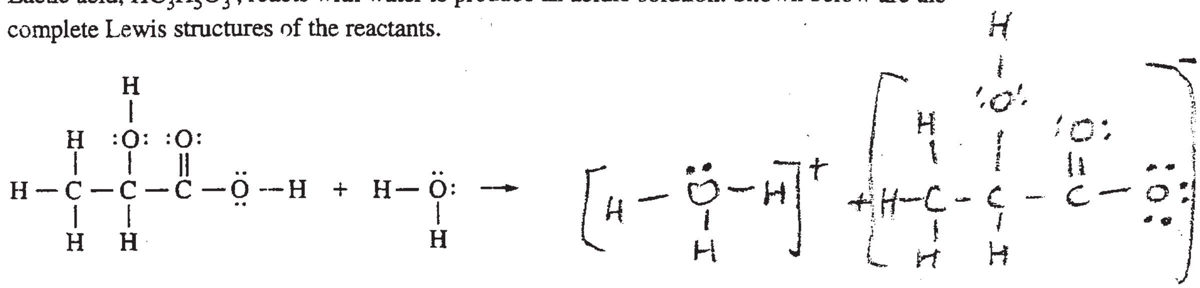
 5A, 1

Answer Question 5 and Question 6. The Section II score weighting for these questions is 15 percent each.

Your responses to these questions will be graded on the basis of the accuracy and relevance of the information cited. Explanations should be clear and well organized. Examples and equations may be included in your responses where appropriate. Specific answers are preferable to broad, diffuse responses.

5. Answer the following questions about laboratory situations involving acids, bases, and buffer solutions.

(a) Lactic acid, $HC_3H_5O_3$, reacts with water to produce an acidic solution. Shown below are the complete Lewis structures of the reactants.



In the space provided above, complete the equation by drawing the complete Lewis structures of the reaction products.

(b) Choosing from the chemicals and equipment listed below, describe how to prepare 100.00 mL of a 1.00 *M* aqueous solution of NH_4Cl (molar mass 53.5 g mol^{-1}). Include specific amounts and equipment where appropriate.

$NH_4Cl(s)$	50 mL buret	100 mL graduated cylinder	100 mL pipet
Distilled water	100 mL beaker	100 mL volumetric flask	Balance

(c) Two buffer solutions, each containing acetic acid and sodium acetate, are prepared. A student adds 0.10 mol of HCl to 1.0 L of each of these buffer solutions and to 1.0 L of distilled water. The table below shows the pH measurements made before and after the 0.10 mol of HCl is added.

	pH Before HCl Added	pH After HCl Added
Distilled water	7.0	1.0
Buffer 1	4.7	2.7
Buffer 2	4.7	4.3

(i) Write the balanced net-ionic equation for the reaction that takes place when the HCl is added to buffer 1 or buffer 2.

(ii) Explain why the pH of buffer 1 is different from the pH of buffer 2 after 0.10 mol of HCl is added.

(iii) Explain why the pH of buffer 1 is the same as the pH of buffer 2 before 0.10 mol of HCl is added.

5a.) ABOVE

b.) (1) Determine the amount of moles present in the desired solution:

(0.1 L)(1.00 M) = 0.1 moles

(2) Determine mass of NH_4Cl needed:

mass || 0.1 moles | 53.5 g / 1 mol NH_4Cl = 5.35 g NH_4Cl

GO ON TO THE NEXT PAGE.

(3) Mass out the desired $NH_4Cl_{(s)}$ with the Balance (5.35 grams)

(4) Place the massed NH_4Cl into a 100mL beaker, and add Enough water until dissolved completely (But dont add exactly all 100mL distilled water, just enough for dissociation!)

(5) Pour the dissolved NH_4Cl into a 100mL volumetric flask, and rince the 100mL beaker to make sure all dissolved NH_4Cl is successfully transfered into the volumetric flask

(6) With the 100mL graduated cylinder, pour distilled water into the 100mL volumetric flask until the liquid level reaches the 100mL mark!

c.) i. $\boxed{H^+_{(aq)} + C_2H_3O_2^-{}_{(aq)} \rightleftharpoons HC_2H_3O_{2\,(aq)}}$

ii. The difference in pH values of Buffer 1 & 2 is most likely due to the differences in initial concentrations of $[HC_2H_3O_2]$ and $[C_2H_3O_2^-]$. Most likely, the magnitudes of concentrations in buffer solution 2 was greater than is buffer solution 1. We can see this in the Henderson Hasselbach equation: $pH = pKa + \log\frac{[A^-]}{[HA]}$. If the magnitudes of $[A^-]$ & $[HA]$ are large, additional acid or base added would not have a significant effect on the final pH.

iii. The pH of buffer solution 1 is the same as pH of buffer solution 2 before any $HCl_{(aq)}$ is added because in order to make a proper buffer solution, the concentrations of $[HA]$ and $[A^-]$ should be the same. Thus, in the Henderson-Hasselbach equation, we can see that:

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

Since the concentrations of $[A^-] = [HA]$, then $\log\frac{[A^-]}{[HA]}$ would equal to 0. Therefore, pH would equal to pKa ($-\log Ka$). Since the buffers of 1 & 2 both involve the same constituents ($HC_2H_3O_2$ & $C_2H_3O_2^-$), then the values of pKa would also be the same. Thus, pH are also the same.

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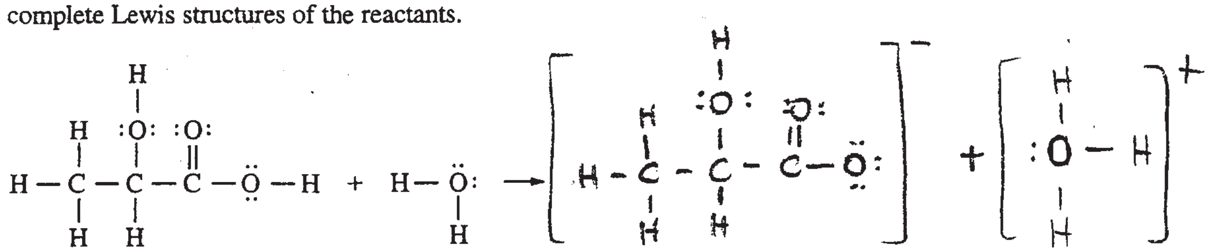

   In the space provided above, complete the equation by drawing the complete Lewis structures of the reaction products.

   (b) Choosing from the chemicals and equipment listed below, describe how to prepare 100.00 mL of a 1.00 *M* aqueous solution of NH_4Cl (molar mass 53.5 g mol^{-1}). Include specific amounts and equipment where appropriate.

$NH_4Cl(s)$	50 mL buret	100 mL graduated cylinder	100 mL pipet
Distilled water	100 mL beaker	100 mL volumetric flask	Balance

   (c) Two buffer solutions, each containing acetic acid and sodium acetate, are prepared. A student adds 0.10 mol of HCl to 1.0 L of each of these buffer solutions and to 1.0 L of distilled water. The table below shows the pH measurements made before and after the 0.10 mol of HCl is added.

	pH Before HCl Added	pH After HCl Added
Distilled water	7.0	1.0
Buffer 1	4.7	2.7
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   (i) Write the balanced net-ionic equation for the reaction that takes place when the HCl is added to buffer 1 or buffer 2.

   (ii) Explain why the pH of buffer 1 is different from the pH of buffer 2 after 0.10 mol of HCl is added.

   (iii) Explain why the pH of buffer 1 is the same as the pH of buffer 2 before 0.10 mol of HCl is added.

b) since $M = 1.00 mol/l$ and $V = 100 ml = 0.100 l$

$n = M \cdot V = (1.00 mol/l)(0.100 l)$

$n = 0.100 mol$ = # of moles of $NH_4Cl(aq)$

$m_{NH_4Cl(s)}$ to be added $= n_{NH_4Cl(aq)} \cdot \text{Molar Mass}_{NH_4Cl}$

$= (0.100 mol) \cdot (53.5 g/mol)$

$= 5.35 g$

since 5.35 g of $NHCl(s)$ is to be added, this specific amount of the solid $NH_4Cl(s)$ is measured with the balance provided.

- As to prepare the solvent with the predetermined volume of 100 ml, the 100 ml pipet is used to measure out 100 ml of water that was filled in a beaker of again 100 ml beforehand. The reason why a beaker is used to carry out the measurement is that it is the most accurate and precise apparatus provided.
- After the 100 ml beaker is cleaned out, the water in the 100 ml pipet is refilled in the beaker to be followed by the 5.35 g of $NH_4Cl(s)$ in hand. The solvent and solute are mixed well to prepare the 1.00 M, 100 ml aqueous solution of $NH_4Cl(aq)$.

c) (i) $CH_3COO^-(aq) + HCl(aq) \rightarrow CH_3COOH(aq) + Cl^-(aq)$

(ii) The reason why these pH values are different is that the amount of $CH_3COO^-(aq)$ ions present in the two buffer solutions are different. There are less moles of $CH_3COO^-(aq)$ present in buffer 1 to neutralize the $HCl(aq)$ added to minimize its influence on the pH.

(iii) The reason is that the ratio of the base concentration to acid concentration in the buffer solutions are the same.

 5B$_3$

ADDITIONAL PAGE FOR ANSWERING QUESTION 5.

- According to the Handerson – Hasselbach equation,

$$pH = pK_a + \log \frac{[Base]}{[Acid]}$$

- In this case,

$$pH = pK_{CH_3COOH} + \log \frac{[CH_3COO^-]}{[CH_3COOH]},$$

- The fact that $pH_1 = pH_2$ means that,

the $\frac{[CH_3COO^-]}{[CH_3COOH]}$ ratio is the same in both buffer solutions.

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 5c,

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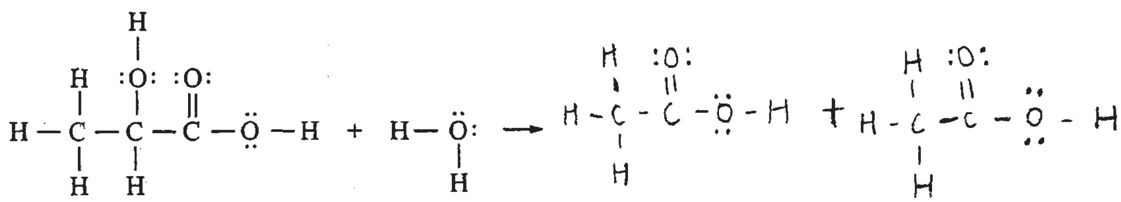

   In the space provided above, complete the equation by drawing the complete Lewis structures of the reaction products.

   (b) Choosing from the chemicals and equipment listed below, describe how to prepare 100.00 mL of a 1.00 *M* aqueous solution of NH_4Cl (molar mass 53.5 g mol^{-1}). Include specific amounts and equipment where appropriate.

$NH_4Cl(s)$	50 mL buret	100 mL graduated cylinder	100 mL pipet
Distilled water	100 mL beaker	100 mL volumetric flask	Balance

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   (ii) Explain why the pH of buffer 1 is different from the pH of buffer 2 after 0.10 mol of HCl is added.

   (iii) Explain why the pH of buffer 1 is the same as the pH of buffer 2 before 0.10 mol of HCl is added.

ADDITIONAL PAGE FOR ANSWERING QUESTION 5.

b) moles of solution = 0.1

Use the balance to get 5.35 gms of HCl
Use the 100 ml beaker to get 100 ml of water
Add the 5.35 gm of HCl to the 100 ml of water

c) i) $CH_3COO^- + H^+ \rightarrow CH_3COOH$

ii) In buffer 1, the amount of sodium acetate is relatively lower than that in buffer 2. Thus as HCl keeps getting added, the acetate ions in buffer 1 get used up more quickly, finally resulting in excess of H^+ ions. Hence pH decreases

iii) Before the HCl is added, the pH of both the buffers is the same because the initial H^+ concentration in both buffers is the same.

Question 5

Sample: 5A
Score: 8

This response earned all 8 points: 2 for part (a), 3 for part (b), 1 for part (c)(i), 1 for part (c)(ii), and 1 for part (c)(iii).

Sample: 5B
Score: 5

Both points were earned in part (a). The first point was earned in part (b) for the correct mass calculation; the other 2 points were not earned because the volumetric flask is not used, and the water is not filled to the mark that indicates a total volume of 100.00 mL. The point was not earned in part (c)(i) because the equation is not a net-ionic equation. The points were earned in parts (c)(ii) and (c)(ii).

Sample: 5C
Score: 3

The points were not earned in part (a). The first point was earned in part (b) for the correct mass of solute (even though the student identifies the solute as HCl instead of NH_4Cl). The other 2 points were not earned because the volumetric flask is not used, and the water is not filled to the mark that indicates a total volume of 100.00 mL. The points were earned in parts (c)(i) and (c)(ii). In part (c)(iii) the point was not earned because the student does not write about the importance of the ratio of acetate to acetic acid but simply states the obvious relationship between H^+ concentration and pH.