Question 5

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

(a) Lactic acid, HC\(_3\text{H}_5\text{O}_3\), reacts with water to produce an acidic solution. Shown below are the complete Lewis structures of the reactants.

\[
\begin{align*}
\text{H} & \quad \text{:O:} & \quad \text{:O:} \\
\text{H} & \quad \text{C} & \quad \text{C} & \quad \text{C} & \quad \text{C} & \quad \text{O:} & \quad \text{H} & \quad \text{H} \\
\text{H} & \quad & \quad & \quad & \quad & \quad & \quad & \quad \\
\end{align*}
\]

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

\[
\begin{align*}
\left[ \begin{array}{c}
\text{H} \\
\text{H} & \quad \text{:O:} & \quad \text{:O:} \\
\text{H} & \quad \text{C} & \quad \text{C} & \quad \text{C} & \quad \text{C} & \quad \text{O:} & \quad \text{H} & \quad \text{H} \\
\end{array} \right]^- & \quad + \quad \left[ \begin{array}{c}
\text{H} & \quad \text{O:} \\
\text{H} \\
\end{array} \right]^+ \\
\end{align*}
\]

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\(_4\)Cl (molar mass 53.5 g mol\(^{-1}\)). Include specific amounts and equipment where appropriate.

NH\(_4\)Cl\((s)\) 50 mL buret 100 mL graduated cylinder 100 mL pipet
Distilled water 100 mL beaker 100 mL volumetric flask Balance

\[
\text{mass of NH}_4\text{Cl} = (0.100 \text{ L})(1.00 \text{ mol L}^{-1})(53.5 \text{ g mol}^{-1}) = 5.35 \text{ g NH}_4\text{Cl}
\]

1. Measure out 5.35 g NH\(_4\)Cl 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\(_4\)Cl to the 100 mL volumetric flask.
4. Continue to add distilled water to the volumetric flask while swirling the flask to dissolve the NH\(_4\)Cl and remove all NH\(_4\)Cl 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.
(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.

<table>
<thead>
<tr>
<th></th>
<th>pH Before HCl Added</th>
<th>pH After HCl Added</th>
</tr>
</thead>
<tbody>
<tr>
<td>Distilled Water</td>
<td>7.0</td>
<td>1.0</td>
</tr>
<tr>
<td>Buffer 1</td>
<td>4.7</td>
<td>2.7</td>
</tr>
<tr>
<td>Buffer 2</td>
<td>4.7</td>
<td>4.3</td>
</tr>
</tbody>
</table>

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

\[ \text{C}_2\text{H}_3\text{O}_2^- + \text{H}_3\text{O}^+ \rightarrow \text{HC}_2\text{H}_3\text{O}_2 + \text{H}_2\text{O} \]

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 \([\text{H}^+]\). Because \(K_a\) for acetic acid is a constant, the ratio of \([\text{H}^+]\) to \(K_a\) must also be constant; this means that the ratio of \([\text{HC}_2\text{H}_3\text{O}_2]\) to \([\text{C}_2\text{H}_3\text{O}_2^-]\) is the same for both buffers, as shown by the following equation, derived from the equilibrium-constant expression for the dissociation of acetic acid.

\[ \frac{[\text{HC}_2\text{H}_3\text{O}_2]}{[\text{C}_2\text{H}_3\text{O}_2^-]} = \frac{[\text{H}^+]}{K_a} \]

After the addition of the \(\text{H}^+\), 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 \(\text{C}_2\text{H}_3\text{O}_2^-\) in buffer 1 must have reacted with the added \(\text{H}^+\) compared to the proportion that reacted in buffer 2. The difference between these proportions means that the original concentrations of \(\text{HC}_2\text{H}_3\text{O}_2\) and \(\text{C}_2\text{H}_3\text{O}_2^-\) 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).
(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.

<table>
<thead>
<tr>
<th>Both buffer solutions have the same acid to conjugate-base mole ratio in the formula below.</th>
<th>One point is earned for the correct answer involving ratio of acid to base in the buffer.</th>
</tr>
</thead>
<tbody>
<tr>
<td>[ [H^+] = K_a \frac{[HC_2H_3O_2]}{[C_2H_3O_2^-]} ]</td>
<td>Therefore, the buffers have the same [H⁺] and pH.</td>
</tr>
</tbody>
</table>
5. Answer the following questions about laboratory situations involving acids, bases, and buffer solutions.

(a) Lactic acid, \( \text{H}_3\text{C}_2\text{H}_5\text{O}_3 \), reacts with water to produce an acidic solution. Shown below are the complete Lewis structures of the reactants.

\[
\begin{align*}
\text{H} & \quad \text{H} \\
:O & \quad :O \\
\text{H} - \text{C} - \text{C} - \text{C} - \text{O} - \text{H} & \quad + \quad \text{H} - \text{O} - \text{H} \\
\text{H} & \quad \text{H}
\end{align*}
\]

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 \( \text{NH}_4\text{Cl} \) (molar mass 53.5 g mol\(^{-1}\)). Include specific amounts and equipment where appropriate.

- \( \text{NH}_4\text{Cl(s)} \) 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.

<table>
<thead>
<tr>
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</thead>
<tbody>
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</tr>
<tr>
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<td>2.7</td>
</tr>
<tr>
<td>Buffer 2</td>
<td>4.7</td>
<td>4.3</td>
</tr>
</tbody>
</table>

(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:

\[
0.1 \text{ L}(1.00\text{ M}) = 0.1 \text{ moles}
\]

(2) Determine the mass of \( \text{NH}_4\text{Cl} \) needed:

\[
\text{mass} = \frac{18\text{ g}}{1\text{ mol}} = 5.35 \text{ g} \text{ NH}_4\text{Cl}
\]

GO ON TO THE NEXT PAGE.
5A

ADDITIONAL PAGE FOR ANSWERING QUESTION 5.

3. Mass out the desired NH₄Cl with the Balance (5.35 grams)

4. Place the massed NH₄Cl into a 100mL beaker, and add enough water until dissolved completely. (But don't add exactly all 100mL distilled water, just enough for dissolving.)

5. Pour the dissolved NH₄Cl into a 100mL volumetric flask, and rinse the 100mL beaker to make sure all dissolved NH₄Cl is successfully transferred 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) \[ \text{H}^+ + \text{C}_2\text{H}_5\text{O}_2^- \rightarrow \text{HC}_2\text{H}_5\text{O}_2^- \]

ii. The difference in pH values of Buffer 1 & 2 is most likely due to the difference in initial concentrations of \([\text{HC}_2\text{H}_5\text{O}_2^-]\) and \([\text{C}_2\text{H}_5\text{O}_2^-]\). Most likely, the magnitude of concentrations in buffer solution 2 was greater than in buffer solution 1. We can see this in the Henderson-Hasselbach equation: \[ \text{pH} = \text{pK}_a + \log \left( \frac{[\text{A}^-]}{[\text{HA}]} \right) \]. If the magnitudes of \([\text{A}^-]\) or \([\text{HA}^-]\) are large, additional acid or base added would not have a significant effect on the final pH.

iii. The pH of buffersolution 1 is the same as pH of buffersolution 2 before any HCl or NaOH 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:

\[ \text{pH} = \text{pK}_a + \log \left( \frac{[\text{A}^-]}{[\text{HA}^-]} \right) \]

Since the concentrations of [A^-] = [HA^-], then \( \log \left( \frac{[\text{A}^-]}{[\text{HA}^-]} \right) \) would equal to 0. Therefore, pH would equal to \( \text{pK}_a \). Since the buffers of 1 & 2 both involve the same constituents (H₃C₂H₅O₂ & C₂H₅O₂⁻), then the values of pH would also be the same. Thus, pH are also the same.
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, \( \text{HC}_3\text{H}_5\text{O}_3 \), reacts with water to produce an acidic solution. Shown below are the complete Lewis structures of the reactants.

\[
\begin{align*}
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{O} \quad \text{O} \\
\text{H} & \quad \text{C} \quad \text{C} \quad \text{C} \quad \text{O} \quad \text{H} \\
\text{H} & \quad \text{H}
\end{align*}
\]

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 \( \text{NH}_4\text{Cl} \) (molar mass 53.5 g mol\(^{-1}\)). Include specific amounts and equipment where appropriate.

\[
\begin{array}{llll}
\text{NH}_4\text{Cl(s)} & 50 \text{ mL buret} & 100 \text{ mL graduated cylinder} & 100 \text{ mL pipet} \\
\text{Distilled water} & 100 \text{ mL beaker} & 100 \text{ mL volumetric flask} & \text{Balance}
\end{array}
\]

(c) Two buffer solutions, each containing acetic acid and sodium acetate, are prepared. A student adds 0.10 mol of \( \text{HCl} \) 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 \( \text{HCl} \) is added.

<table>
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<tr>
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<tr>
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<td>4.3</td>
</tr>
</tbody>
</table>

(i) Write the balanced net-ionic equation for the reaction that takes place when the \( \text{HCl} \) 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 \( \text{HCl} \) is added.

(iii) Explain why the pH of buffer 1 is the same as the pH of buffer 2 before 0.10 mol of \( \text{HCl} \) is added.
b) Since \( M = 1.00 \text{ mol/L} \) and \( V = 100 \text{ ml} = 0.100 \text{ L} \)
\[
\text{n = M \cdot V} = (1.00 \text{ mol/L}) (0.100 \text{ L})
\]
\[
\text{n = 0.100 \text{ mol} = \# of moles of NH}_4\text{Cl(aq)}
\]
\[
\text{m}_{\text{NH}_4\text{Cl(s)}} \text{ to be added} = \text{n}_{\text{NH}_4\text{Cl(aq)}} \cdot \text{Molar Mass}_{\text{NH}_4\text{Cl}}
\]
\[
= (0.100 \text{ mol}) \cdot (53.5 \text{ g/mol})
\]
\[
= 5.35 \text{ g}
\]

Since 5.35 g of NH\(_4\)Cl(s) is to be added, this specific amount of the solid NH\(_4\)Cl(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\(_4\)Cl(s) in head. The solvent and solute are mixed well to prepare the 1.00 mL 100 mL aqueous solution of NH\(_4\)Cl(aq).

e) (i) \( \text{CH}_3\text{COO}^-\text{(aq)} + \text{H}^+\text{(aq)} \rightarrow \text{CH}_3\text{COOH (aq)} + \text{e}^-\text{(aq)} \)

(ii) The reason why the pH values are different is that the amount of \( \text{CH}_3\text{COO}^-\text{(aq)} \) ions present in the two buffer solutions are different. There are less moles of \( \text{CH}_3\text{COO}^-\text{(aq)} \) present in buffer I to neutralize the \( \text{H}^+\text{(aq)} \) 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.
• According to the Henderson - Hasselbach equation
  \[ \text{pH} = pK_a + \log \frac{[\text{Base}]}{[\text{Acid}]} \]

• In full case
  \[ \text{pH} = pK_{CH_3COOH} + \log \frac{[CH_3COO^-]}{[CH_3COOH]} \]

• The fact that \( \text{pH}_1 = \text{pH}_2 \) mean that
  \[ \frac{[CH_3COO^-]}{[CH_3COOH]} \] ratio is the same in both buffer solutions.
5. Answer the following questions about laboratory situations involving acids, bases, and buffer solutions.

(a) Lactic acid, H₃C₅H₂O₄, reacts with water to produce an acidic solution. Shown below are the complete Lewis structures of the reactants.

\[
\begin{align*}
\text{H} & \quad \text{H} & \quad \text{H} \\
\text{H} & \quad \text{O} & \quad \text{O} \\
\text{H} & \quad \text{C} & \quad \text{C} & \quad \text{O} & \quad \text{H} \\
\text{H} & \quad & \quad & \quad & \quad \\
\text{H} & \quad & \quad & \quad & \quad \\
\end{align*}
\]

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 NH₄Cl (molar mass 53.5 g mol⁻¹). Include specific amounts and equipment where appropriate.

<table>
<thead>
<tr>
<th>Chemical</th>
<th>Amount</th>
</tr>
</thead>
<tbody>
<tr>
<td>NH₄Cl(s)</td>
<td>50 mL buret</td>
</tr>
<tr>
<td>Distilled water</td>
<td>100 mL beaker</td>
</tr>
<tr>
<td></td>
<td>100 mL graduated cylinder</td>
</tr>
<tr>
<td></td>
<td>100 mL volumetric flask</td>
</tr>
<tr>
<td></td>
<td>100 mL pipet</td>
</tr>
<tr>
<td></td>
<td>100 mL balancing</td>
</tr>
</tbody>
</table>

(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.

<table>
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<tr>
<td>Buffer 2</td>
<td>4.7</td>
</tr>
</tbody>
</table>

(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) 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 gms of HCl to the 100 ml of water

c) i) \( \text{CH}_3\text{COO}^- + \text{H}^+ \rightarrow \text{CH}_3\text{COOH} \)

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, resulting in excess of \( \text{H}^+ \) ions. Hence pH decreases.

iii) Before the HCl is added, the pH of both the buffers is the same because the \( \text{H}^+ \) concentration in both buffers is the same.
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 NH₄Cl). 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.