Acid Base Equilibrium Review

Proof of true understanding of acid base equilibrium culminates in the ability to find pH of any solution or combination of solutions. The ability to determine pH of a multitude of solutions encompasses many of the basic knowledge pieces that make up the solution equilibrium component of Big Idea 6 in the AP Chemistry course description. Also keep in mind that titration curves, which graph the relationship between pH and concentrations, are created by series of pH calculations. Titrations are a very important, yet often rushed topic in AP Chemistry, so let’s make sure you see the big picture.

You should be able to find the pH of:
1. A strong acid or base
2. A weak acid or base
3. A salt solution
4. A buffer solution
5. A mixture of a strong acid or base with any of the others (aka Invader Problem)

Let’s start by making sure you can identify which type of solution is present before we review the math. For each of these solutions, identify which of the solution types (#1-5 above) are present.

1. NH₃
2. HClO₄
3. HC₂H₃O₂
4. NH₄Cl
5. HNO₃ + NH₃
6. HNO₂ + NaNO₂
7. NaCHO₂
8. C₂H₅NH₂ + C₂H₅NH₃Cl
9. HF
10. KOH + HC₂H₅O₂

Now, let’s review the math to calculate the pH for each type of solution.

**STRONG ACID OR BASE**
To find the pH of a strong acid or base, take the –log of the acid or base concentration. The –log of a strong acid concentration will give the pH. The –log of a strong base concentration will give the pOH. The only reason we can use the concentration of the molarity directly is because these acids and bases are strong. Strong acids and bases dissociate 100% so for every one mole of acid or base there is one mole of hydronium ions (H⁺) and hydroxide ions (OH⁻) respectively.

Ex 1) Calculate the pH of 0.450M HCl

Ex 2) Calculate the pH of 0.710M KOH
**WEAK ACID OR BASE**
To find the pH of a weak acid or base, write a Lowry Brønsted reaction by adding the weak acid or weak base to water. Put that reaction in a RICE table. The given concentration is the “initial” concentration of the acid or base reactant. Solve for the RICE table “x” by writing the equilibrium expression, plugging in the values from the equilibrium line of the RICE table, and setting it equal to the given $K_a$ or $K_b$. To find the pH/pOH, take the $-\log[x]$ since “x” represents the hydronium ion concentration in acids and the hydroxide ion concentration in bases.

Ex 3) Calculate the pH of a 0.150M solution of ammonia. The $K_b$ of ammonia is $1.80 \times 10^{-5}$.

Ex 4) Calculate the pH of a 0.227M solution of hydrocyanic acid (HCN). The $K_a = 6.2 \times 10^{-10}$.

**SALT SOLUTION (HYDROLYSIS)**
To find the pH of a salt solution, first determine which ion within the salt will undergo hydrolysis. In other words, which ion from the salt will react with water to form a compound that won’t significantly dissociate? A quick way to figure that out is to look at each ion in the salt and ask yourself, “if this is paired with an $H^+$ or $OH^-$ will that substance be ‘strong’ or ‘weak’?” We’re looking for the one that makes a weak product because that product will not re-ionize and will thus change the $H^+/OH^-$ concentration in the solution. Once you’ve identified the ‘weak’ ion, write an equation for that ion splitting a water molecule and use that as the Reaction line in a RICE table. Proceed with solving for pH as you do with any other weak acid or base problem.

One other important component of hydrolysis problems is the conversion of $K_a$ to $K_b$ or vice versa. The question prompts given on the AP exam will provide the $K_a$ or $K_b$ of the *conjugate* acid or base, not the $K$ of the ion you are using in the hydrolysis equation. The $K_a$ or $K_b$ of the hydrolyzing ion needs to be determined using $K_w = K_a \times K_b$. *This is a necessary step for every hydrolysis problem!*

Ex 5) Calculate the pH of a 0.95M $KC_3H_5O_3$ solution (potassium lactate).

The $K_a$, for lactic acid, $HC_3H_5O_3$ is $7.1 \times 10^{-12}$.
**BUFFER SOLUTION**

*Method 1:* If a solution contains a weak acid and its conjugate base, or a weak base and its conjugate acid, then a buffer is present. Using the Henderson-Hasselbach (HH) equation is one way to determine pH of a buffer solution. There is a pH version and a pOH version of the equation. The acid version of this equation is provided to students on the AP Equation insert of the AP Chemistry exam. You are allowed access to that formula sheet for the duration of the test. You can also program this equation into your calculator!

\[
pH = pK_a + \log \frac{[A^-]}{[HA]}
\]
\[
pOH = pK_b + \log \frac{[HB^+]}{[B]}
\]

Using two versions of the HH equation prevents you from getting the concentrations of the acids/bases and their conjugates confused when plugging in to the log ratio. Always put the ION concentration in the numerator and the acid/base concentration in the denominator. Say “ION ON TOP!” over and over to help you remember.

\[
pH = pK_a + \log \frac{[ion]}{[acid]}
\]
\[
pOH = pK_b + \log \frac{[ion]}{[base]}
\]

Ex 6a) Calculate the pH of a solution which is 0.53M in HC₆H₄NO₂ and 0.50M NaC₆H₄NO₂. The Kₐ for nicotinic acid is 1.7x10⁻⁵.

Ex 7a) Calculate the pH of a solution which is 0.245M in NH₃ and 0.245M in NH₄Cl. The Kₐ for ammonia is 1.8x10⁻⁵.

*Method 2:* Another method for solving buffer problems using an equation similar to HH but without the logarithms. (Actually, when deriving the HH equation, this equation occurs just one step prior to taking the logarithms that generated the HH equation). In this approach we use a modified Kₐ expression that isolates the [H⁺], rather than the Kₐ:

\[
[H^+] = K_a \frac{[acid]}{[base]}
\]

So, in your buffer system, start by identifying the component that is easily recognizable as either a weak acid or base. That concentration is the one that will go in the [acid] or [base] brackets. Then, the conjugate of that weak acid/base is what will go in the remaining bracket. Always plug in the Kₐ and you will always be solving for the [H⁺]. If they give you a Kₐ, just convert it to a Ka with Kₐ= Kₐ x Kₐ and then plug in the Ka. There is no need to consider a “basic” version of this equation because we are almost always interested in the [H⁺] even if it’s a basic solution. Once you have the [H⁺], convert it to pH with the standard –log[H⁺] and you’re done.

Ex 6b) Calculate the pH of a solution which is 0.53M in HC₆H₄NO₂ and 0.50M NaC₆H₄NO₂. The Kₐ for nicotinic acid is 1.7x10⁻⁵.
Ex 7b) Calculate the pH of a solution which is 0.245M in NH₃ and 0.245M in NH₄Cl. The \( K_b \) for ammonia is \( 1.8 \times 10^{-5} \).

MIXTURE OF STRONG ACID/BASE WITH ANYTHING ELSE

Treat the addition of a strong acid or base to a system already at equilibrium as an “invader.” This invasion creates a war. The war is represented by the one-way \( \rightarrow \) stoichiometric reaction between the invader and the system at equilibrium. Once the war is over, the stoichiometric survivors remaining will then determine the pH of the solution. This is a multistep process that will be shown in the three examples provided for you below, with each example having a different outcome.

Ex 8) 125mL of a 0.525M HCl is mixed with 125mL of a buffer system containing 1M acetic acid and 1M sodium acetate. Find the pH of the solution. The \( K_a \) of acetic acid is \( 1.8 \times 10^{-5} \).

Before you start, since this is a buffer, make sure you can identify the components:
The acid is _________ and the base is _____________

Step 1: Write the invasion equation as the strong acid enters the buffer system. (Think: which component of the buffer will neutralize this invader? And note, this is NOT an equilibrium reaction \( \rightarrow \) reaction only)

Step 2: Figure out the stoichiometric survivors using a quick mole stoichiometry table such as:

<table>
<thead>
<tr>
<th>Initial moles</th>
<th>Change in moles</th>
<th>Final moles</th>
</tr>
</thead>
</table>

Step 3: Based on the survivors, make a plan. Are there both [acid] and [base] components of a buffer? Or, is there just one survivor? If it’s still a buffer, plug the final moles of each piece into either HH or buffer equation and solve for pH. If there is only one survivor left, find its pH appropriately as an acid, base or salt.
Ex 9) 125mL of a 0.500M NaOH is mixed with 125mL of 0.500M acetic acid. Find the pH of the solution. The $K_a$ of acetic acid is $1.8 \times 10^{-5}$.

**Step 1:** Write the invasion equation as the strong acid enters the acid system. (Note, this is NOT an equilibrium reaction $\rightarrow$ reaction only)

**Step 2:** Figure out the stoichiometric survivors using a quick mole stoichiometry table such as:

<table>
<thead>
<tr>
<th>Initial moles</th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>Change in moles</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Final moles</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Step 3:** Based on the survivors, make a plan. Are there both [acid] and [base] components of a buffer? Or, is there just one survivor? If it’s still a buffer, plug the final moles of each piece into either HH or buffer equation and solve for pH. If there is only one survivor left, find its pH appropriately as an acid, base or salt.
Ex 10) 200mL of a 0.500M NaOH is mixed with 125mL of 0.500M acetic acid. Find the pH of the solution. The $K_a$ of acetic acid is $1.8 \times 10^{-5}$.

*Step 1:* Write the invasion equation as the strong acid enters the acid system. (Note, this is NOT an equilibrium reaction $\rightarrow$ reaction only)

*Step 2:* Figure out the stoichiometric survivors using a quick mole stoichiometry table such as:

<table>
<thead>
<tr>
<th>Initial moles</th>
<th>Change in moles</th>
<th>Final moles</th>
</tr>
</thead>
</table>

*Step 3:* Based on the survivors, make a plan. Are there both [acid] and [base] components of a buffer? Or, is there just one survivor? If it’s still a buffer, plug the final moles of each piece into either HH or buffer equation and solve for pH. If there is only one survivor left, find its pH appropriately as an acid, base or salt.
Acid Base Equilibrium Review

Proof of true understanding of acid base equilibrium culminates in the ability to find pH of any solution or combination of solutions. The ability to determine pH of a multitude of solutions encompasses many of the basic knowledge pieces that make up the solution equilibrium component of Big Idea 6 in the AP Chemistry course description. Also keep in mind that titration curves, which graph the relationship between pH and concentrations, are created by series of pH calculations. Titrations are a very important, yet often rushed topic in AP Chemistry, so let’s make sure you see the big picture.

You should be able to find the pH of:
   1. A strong acid or base
   2. A weak acid or base
   3. A salt solution
   4. A buffer solution
   5. A mixture of a strong acid or base with any of the others (aka Invader Problem)

Let’s start by making sure you can identify which type of solution is present before we review the math. For each of these solutions, identify which of the solution types (#1-5 above) are present.

1. NH₃ weak base
2. HClO₄ strong acid
3. HC₂H₃O₂ weak acid
4. NH₄Cl salt
5. HNO₃ + NH₃ mixture of strong acid and weak base
6. HNO₂ + NaNO₂ buffer
7. NaCHO₂ salt
8. C₂H₃NH₂ + C₂H₅NH₃Cl buffer
9. HF weak acid
10. KOH + HC₂H₅O₂ mixture of strong base and weak acid

Now, let’s review the math to calculate the pH for each type of solution.

**STRONG ACID OR BASE**

To find the pH of a strong acid or base, take the –log of the acid or base concentration. The –log of a strong acid concentration will give the pH. The –log of a strong base concentration will give the pOH. The only reason we can use the concentration of the molarity directly is because these acids and bases are strong. Strong acids and bases dissociate 100% so for every one mole of acid or base there is one mole of hydronium ions (H⁺) and hydroxide ions (OH⁻) respectively.

Ex 1) Calculate the pH of 0.450M HCl

Since this is a strong acid and dissociates 100% assume that the concentration of the acid equals the concentration of the hydronium ion in solution.

\[ \text{pH} = -\log[H^+] \]

\[ \text{pH} = -\log[0.450] \]

\[ \text{pH} = 0.347 \]
Ex 2) Calculate the pH of 0.710M KOH

Since this a strong base and dissociates 100% assume that the concentration of the acid equals the concentration of the hydroxide ions in solution

\[ \text{pOH} = -\log[\text{OH}^-] \]

\[ \text{pOH} = -\log[0.710] \]

\[ \text{pOH} = 0.347 \]

\[ \text{pH} = 14 - 0.347 \]

\[ \text{pH} = 13.651 \]

**Weak Acid or Base**

To find the pH of a weak acid or base, write a Lowry Brönsted reaction by adding the weak acid or weak base to water. Put that reaction in a RICE table. The given concentration is the “initial” concentration of the acid or base reactant. Solve for the RICE table “x” by writing the equilibrium expression, plugging in the values from the equilibrium line of the RICE table, and setting it equal to the given \( K_a \) or \( K_b \). To find the pH/pOH, take the \(-\log[x]\) since “x” represents the hydronium ion concentration in acids and the hydroxide ion concentration in bases.

Ex 3) Calculate the pH of a 0.150M solution of ammonia. The \( K_b \) of ammonia is \( 1.80 \times 10^{-5} \).

\[
K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} \]

The use of a RICE table will lead to the below substitutions into the equilibrium expression. The x in the denominator is neglected due to mathematical insignificance.

\[ K_b = \frac{x^2}{M - x} \text{ where } x = [\text{OH}^-] \]

\[ 1.80 \times 10^{-5} = \frac{x^2}{0.150} \]

\[ x = 1.64 \times 10^{-3} = [\text{OH}^-] \]

\[ \text{pOH} = -\log(1.64 \times 10^{-3}) = 2.784 \]

\[ \text{pH} = 11.216 \]

Ex 4) Calculate the pH of a 0.227M solution of hydrocyanic acid (HCN). The \( K_a = 6.2 \times 10^{-10} \).

The use of a RICE table will lead to the below substitutions into the equilibrium expression. The x in the denominator is neglected due to mathematical insignificance.

\[ K_a = \frac{x^2}{M - x} \text{ where } x = [\text{H}^+] \]

\[ 6.20 \times 10^{-10} = \frac{x^2}{0.227} \]

\[ x = 1.19 \times 10^{-5} = [\text{H}^+] \]

\[ \text{pH} = -\log(1.19 \times 10^{-5}) = 4.926 \]
SALT SOLUTION (HYDROLYSIS)

To find the pH of a salt solution, first determine which ion within the salt will undergo hydrolysis. In other words, which ion from the salt will react with water to form a compound that won’t significantly dissociate? A quick way to figure that out look at each ion in the salt and ask yourself, “if this is paired with an H⁺ or OH⁻ will that substance be ‘strong’ or ‘weak’?” We’re looking for the one that makes a weak product because that product will not re-ionize and will thus change the H⁺/OH⁻ concentration in the solution. Once you’ve identified the ‘weak’ ion, write an equation for that ion splitting a water molecule and use that as the Reaction line in a RICE table. Proceed with solving for pH as you do with any other weak acid or base problem.

One other important component of hydrolysis problems is the conversion of Kₐ to Kₐ or vice versa. The question prompts given on the AP exam will provide the Kₐ or Kₐ of the conjugate acid or base, not the K of the ion you are using in the hydrolysis equation. The Kₐ or Kₐ of the hydrolyzing ion needs to be determined using Kₐ = Kₐ x Kₐ. This is a necessary step for every hydrolysis problem!

Ex 5) Calculate the pH of a 0.95M KC₃H₅O₃ solution (potassium lactate).

The dissociated ion from the salt that undergoes hydrolysis is C₃H₅O₃⁻. Add that ion to water and put it in a RICE table along with the following equation. C₃H₅O₃⁻ + H₂O ⇌ HC₃H₅O₃ + OH⁻

Kₐ is needed in the equation instead of Kₐ.

\[ K_b = \frac{K_a}{K_w} = \frac{1 \times 10^{-14}}{7.1 \times 10^{-12}} = 1.4 \times 10^{-3} \]

Equilibrium expression from reaction:

\[ K_b = \frac{[HC₃H₅O₃][OH^-]}{[C₃H₅O₃^-]} \]

The use of a RICE table will lead to the below substitutions into the equilibrium expression.

\[ K_b = \frac{x^2}{0.95} = 1.4 \times 10^{-3} \]

\[ x = [OH^-] = 3.6 \times 10^{-2} \]

\[ pOH = - \log(3.6 \times 10^{-2}) = 1.47 \]

\[ pH = 14 - 1.47 = 12.53 \]
**BUFFER SOLUTION**

**Method 1:** If a solution contains a weak acid and its conjugate base, or a weak base and its conjugate acid, then a buffer is present. Using the Henderson-Hasselbach (HH) equation is one way to determine pH of a buffer solution. There is a pH version and a pOH version of the equation. The acid version of this equation is provided to students on the AP Equation insert of the AP Chemistry exam. You are allowed access to that formula sheet for the duration of the test. You can also program this equation into your calculator!

\[
pH = pK_a + \log \left( \frac{[A^-]}{[HA]} \right) \quad pOH = pK_b + \log \left( \frac{[HB^-]}{[B]} \right)
\]

Using two versions of the HH equation prevents you from getting the concentrations of the acids/bases and their conjugates confused when plugging in to the log ratio. Always put the ION concentration in the numerator and the acid/base concentration in the denominator. Say “ION ON TOP!” over and over to help you remember.

**Ex 6a)** Calculate the pH of a solution which is 0.53M in HCsH4NO2 and 0.50M NaC6H4NO2. The $K_a$ for nicotinic acid is $1.7 \times 10^{-5}$.

**Solution:**

Since this is a solution which contains a weak acid and its conjugate base we will plug numbers in the acidic version of Henderson Hasselbach (HH) and solve.

\[
pH = pK_a + \log \left( \frac{[ion]}{[acid]} \right)
\]

\[
pH = -\log(1.7 \times 10^{-5}) + \log \left( \frac{0.50}{0.53} \right) = 4.74
\]

**Ex 7a)** Calculate the pH of a solution which is 0.245M in NH3 and 0.245M in NH4Cl. The $K_b$ for ammonia is $1.8 \times 10^{-5}$.

**Solution:** Since this is a solution that contains a weak acid and its conjugate base we will plug numbers into the basic version of HH and solve.

\[
pOH = pK_b + \log \left( \frac{[ion]}{[base]} \right)
\]

\[
pOH = -\log(1.8 \times 10^{-5}) + \log \left( \frac{0.245}{0.245} \right) = 4.74
\]

\[
pH = 14 - 4.74 = 9.26
\]

You can see above that the ratio of base to conjugate acid is 1. When this is the case pOH=pKb. This is a very important concept in titrations because it’s considered to be the titration’s midpoint or halfway point.

**Method 2:** Another method for solving buffer problems using an equation similar to HH but without the logarithms. (Actually, when deriving the HH equation, this equation occurs just one step prior to taking the logarithms that generated the HH equation). In this approach we use a modified $K_a$ expression that isolates the $[H^+]$, rather than the $K_a$:

\[
[H^+] = K_a \frac{[acid]}{[base]}
\]
So, in your buffer system, start by identifying the component that is easily recognizable as either a weak acid or base. That concentration is the one that will go in the [acid] or [base] brackets. Then, the conjugate of that weak acid/base is what will go in the remaining bracket. Always plug in the $K_a$ and you will always be solving for the $[H^+]$. If they give you a $K_b$, just convert it to a $K_a$ with $K_w = K_a \times K_b$ and then plug in the $K_a$. There is no need to consider a “basic” version of this equation because we are almost always interested in the $[H^+]$ even if it’s a basic solution. Once you have the $[H^+]$, convert it to pH with the standard $-\log[H^+]$ and you’re done.

Ex 6b) Calculate the pH of a solution which is 0.53M in HC$_6$H$_4$NO$_2$ and 0.50M NaC$_6$H$_4$NO$_2$. The $K_a$ for nicotinic acid is 1.7x10$^{-5}$.

In this solution the HC$_6$H$_4$NO$_2$ is a recognizable [acid] and thus the NaC$_6$H$_4$NO$_2$ must be the [base].

\[
[H^+] = 1.7 \times 10^{-5} \frac{[0.53M]}{[0.50M]} = 1.8 \times 10^{-5}
\]

\[
pH = -\log(1.8 \times 10^{-5}) = 4.74
\]

Ex 7b) Calculate the pH of a solution which is 0.245M in NH$_3$ and 0.245M in NH$_4$Cl. The $K_b$ for ammonia is 1.8x10$^{-5}$.

In this solution the NH$_3$ is a recognizable [base] and thus the NH$_4$Cl must be the [acid]. Since we will always use the $K_a$ in this equation, you’ll need to convert $K_b$ to $K_a$ using $K_w = K_a \times K_b$. Save yourself a step and just use the $K_w/K_b$ fraction in the $K_a$ spot of the formula.

\[
[H^+] = \left( \frac{1 \times 10^{-14}}{1.8 \times 10^{-5}} \right) \frac{[0.245M]}{[0.245M]} = 5.6 \times 10^{-10}
\]

\[
pH = -\log(5.6 \times 10^{-10}) = 9.25
\]

And, did you notice what happened when the [acid]/[base] ratio is equal to 1? Yep, the $[H^+] = K_a$, or if we take the $-\log$ of both, the $pH = pK_a$. This represents the “best” buffer because it contains equivalent numbers of acid and base particles that could deal with an invader from either side.
MIXTURE OF STRONG ACID/BASE WITH ANYTHING ELSE
Treat the addition of a strong acid or base to a system already at equilibrium as an “invader.” This invasion creates a war. The war is represented by the one-way (→) stoichiometric reaction between the invader and the system at equilibrium. Once the war is over, the stoichiometric survivors remaining will then determine the pH of the solution. This is a multistep process that will be shown in the three examples provided for you below, with each example having a different outcome.

Ex 8) 125mL of a 0.525M HCl is mixed with 125mL of a buffer system containing 1M acetic acid and 1M sodium acetate. Find the pH of the solution. The $K_a$ of acetic acid is $1.8 \times 10^{-5}$.

Before you start, since this is a buffer, make sure you can identify the components:
The acid is acetic acid and the base is acetate ion

**Step 1:** Write the invasion equation as the strong acid enters the buffer system. (Think: which component of the buffer will neutralize this invader? And note, this is NOT an equilibrium reaction → reaction only)
The first step is to write a reaction representing the “invasion” of the strong acid into the buffer system. The strong acid $H^+$ or the strong base $OH^-$ will always be a reactant. In this case the invader $H^+$ will react with only one component of the buffer but will produce the other. This is NOT an equilibrium reaction. This reaction will go to completion. After the invasion the war will eventually come to an end.
Writing these first equations is usually the sticking point for most students.

RXN: $H^+ + C_2H_3O_2^- \rightarrow HC_2H_3O_2$

**Step 2:** Figure out the stoichiometric survivors using a quick mole stoichiometry table such as:
The next step is to make a “stoichiometry table” which somewhat resembles a RICE table but instead of concentrations or pressures, it’s a comparison of moles. From the table it can be determined how many moles of each reactant react and how much of what is left over.

<table>
<thead>
<tr>
<th></th>
<th>$H^+$</th>
<th>$C_2H_3O_2^-$</th>
<th>$HC_2H_3O_2$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial moles</td>
<td>0.0656</td>
<td>0.125</td>
<td>0.125</td>
</tr>
<tr>
<td>Change in moles</td>
<td>-0.0656</td>
<td>-0.0656</td>
<td>+0.0656</td>
</tr>
<tr>
<td>Final moles</td>
<td>0</td>
<td>0.0594</td>
<td>0.1906</td>
</tr>
</tbody>
</table>

**Step 3:** Based on the survivors, make a plan. Are there both [acid] and [base] components of a buffer? Or, is there just one survivor? If it’s still a buffer, plug the final moles of each piece into either HH or buffer equation and solve for pH. If there is only one survivor left, find its pH appropriately as an acid, base or salt.

Now that the war is over, study what’s remaining. The invader lost. There is none left. Since there is still moles of acid and conjugate base left over in the solution we now solve this as an ordinary buffer problem. The only thing to watch out for is that the table was calculated in MOLES. We need to convert these moles to concentration by dividing by the total final volume in the majority of cases.

However, since both of our approaches rely on a ratio of the two components, dividing by the total volume will just cancel out anyway, so we can save a step and just plug in the moles rather than convert to molarity.

**Method 1:**
Method 2:

\[
[H^+] = (1.8 \times 10^{-5}) \frac{[0.1906]}{[0.0594]} = 5.8 \times 10^{-5}
\]

pH = −log(5.8 \times 10^{-5}) = 4.23

Ex 9) 125mL of a 0.500M NaOH is mixed with 125mL of 0.500M acetic acid. Find the pH of the solution.

The \( K_a \) of acetic acid is 1.8 \times 10^{-5}.

Step 1: Write the invasion equation as the strong acid enters the acid system. (Note, this is NOT an equilibrium reaction → reaction only)

The first step is to write a reaction representing the “invasion” of the strong base into a weak acid. In this case the invader OH\(^-\) will react with the only thing there is to react with. This is an easier equation to write than the last.

RXN: \( \text{OH}^- + \text{HC}_2\text{H}_3\text{O}_2 \rightarrow \text{C}_2\text{H}_3\text{O}_2^- + \text{H}_2\text{O} \)

Step 2: Figure out the stoichiometric survivors using a quick mole stoichiometry table such as:

<table>
<thead>
<tr>
<th></th>
<th>OH(^-)</th>
<th>HC(_2)H(_3)O(_2)</th>
<th>C(_2)H(_3)O(_2)(^-)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial moles</td>
<td>0.0625</td>
<td>0.0625</td>
<td>0</td>
</tr>
<tr>
<td>Change in moles</td>
<td>-0.0625</td>
<td>-0.0625</td>
<td>+0.0625</td>
</tr>
<tr>
<td>Final moles</td>
<td>0</td>
<td>0</td>
<td>0.0625</td>
</tr>
</tbody>
</table>

Step 3: Based on the survivors, make a plan. Are there both [acid] and [base] components of a buffer? Or, is there just one survivor? If it’s still a buffer, plug the final moles of each piece into either HH or buffer equation and solve for pH. If there is only one survivor left, find its pH appropriately as an acid, base or salt.

In this scenario the only thing left over is an ion that undergoes salt hydrolysis. Which means we’re going to treat this just like finding the pH of a salt solution. First, take the moles remaining and convert it to molarity so that it looks like the other salt solution problems you’ve worked.

\[
0.0625\,\text{mol} / 0.250\,\text{L} = 0.250\,\text{M} \text{C}_2\text{H}_3\text{O}_2^-
\]

Remember, this is just a regular salt hydrolysis problem. (Don’t forget to convert \( K_a \) to \( K_b \))

\[
K_b = \frac{K_w}{K_a} = \frac{1 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.56 \times 10^{-10}
\]

The dissociated ion from the salt that undergoes hydrolysis is \( \text{C}_2\text{H}_3\text{O}_2^- \). Add that ion to water and this is what happens: \( \text{C}_2\text{H}_3\text{O}_2^- + \text{H}_2\text{O} \rightleftharpoons \text{HC}_2\text{H}_3\text{O}_2 + \text{OH}^- \)

This is the R equation for your RICE table, if you need it. And, it makes the equilibrium expression:

\[
K_b = \frac{[\text{HC}_2\text{H}_3\text{O}_2][\text{OH}^-]}{[\text{C}_2\text{H}_3\text{O}_2^-]}\]
From a rice table we substitute to get (neglecting $-x$):

$$5.56 \times 10^{-10} = \frac{x^2}{0.250}$$

$$x = [OH^-] = 1.18 \times 10^{-5}$$

$$pOH = -\log(1.18 \times 10^{-5}) = 4.92$$

$$pH = 14 - 4.92 = 9.08$$

Note: This is a significant scenario in an acid base titration because this represents the equivalence point of a titration and shows why the pH in a weak/strong situation will not have an equivalence point of 7.

Ex 10) 200mL of a 0.500M NaOH is mixed with 125mL of 0.500M acetic acid. Find the pH of the solution. The $K_a$ of acetic acid is $1.8 \times 10^{-5}$.

**Step 1:** Write the invasion equation as the strong acid enters the acid system. (Note, this is NOT an equilibrium reaction $\rightarrow$ reaction only)

The first step is to write a reaction representing the “invasion” of the strong base into a weak acid. In this case the invader $OH^-$ will react with the only thing there is to react with. This is an easier equation to write than the last.RXN: $OH^- + HC_2H_3O_2 \rightarrow C_2H_3O_2^- + H_2O$

**Step 2:** Figure out the stoichiometric survivors using a quick mole stoichiometry table such as:

<table>
<thead>
<tr>
<th></th>
<th>OH$^-$</th>
<th>HC$_2$H$_3$O$_2$</th>
<th>C$_2$H$_3$O$_2$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial moles</td>
<td>0.100</td>
<td>0.0625</td>
<td>0</td>
</tr>
<tr>
<td>Change in moles</td>
<td>$-0.0625$</td>
<td>$-0.0625$</td>
<td>$+0.0625$</td>
</tr>
<tr>
<td>Final moles</td>
<td>0.0375</td>
<td>0</td>
<td>0.0625</td>
</tr>
</tbody>
</table>

**Step 3:** Based on the survivors, make a plan. Are there both [acid] and [base] components of a buffer? Or, is there just one survivor? If it’s still a buffer, plug the final moles of each piece into either HH or buffer equation and solve for pH. If there is only one survivor left, find its pH appropriately as an acid, base or salt.

In this scenario there is some strong base left over. The invader wins!! Any time there is a strong base or strong acid left over, you just use those left over moles to find pH. Any other moles of any other substance will not significantly change the pH if there are strong ions present. DON’T FORGET to convert the moles to concentration by dividing by total volume before finding pH. 0.0375mol/0.325L = 0.115M OH$^-$

$pOH= -\log[OH^-]$

$pOH= -\log(0.115)$

$pOH= 0.94$

$pH = 14 - 0.94$

$pH= 13.06$