2-23-12 Follow-up Wkst – Acid and Base pH Calculations

For each of the following solutions: Write a chemical equation, identify the limiting reactant (if there is one), and calculate the pH.
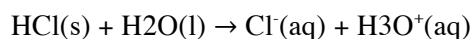
We will calculate the pH of the solutions using the following 3 steps for each problem.
    Step 1: What is left in solution?
    Step 2: What are the equilibrium concentrations of the species in solution?
    Step 3: What is the pH of the solution?

1. **0.1 M HCl**

$HCl(s) + H2O(l) \rightarrow Cl^-(aq) + H3O^+(aq)$

This reaction goes to completion because HCl is a strong acid. So, all the HCl disassociates into Cl^- and $H3O^+$ ions.

Calculate the pH

Step 1: What is left in solution?

In RICE tables, we need to convert all concentrations into moles. To make the calculations of concentrations easier later in the problem, we assume a volume of 1 L of 0.1 M HCl. This way, the number of moles and molarity is the same value (they just have different units).

R	$HCl(s) +$	$H2O(l) \rightarrow$	$Cl^-(aq) +$	$H3O^+(aq)$
I	0.1	-	0	0
C	-0.1	-	+0.1	+0.1
E	0	-	0.1	0.1

## Step 2: What are the equilibrium concentrations of the species in solution?

The only equilibrium concentrations we are concerned with when calculating the pH of a solution are the concentrations of $H3O^+$ ions, OH^- ions, weak acids and weak bases. So here we have $[H3O^+] = 0.1$ M.

Step 3: What is the pH of the solution?

$pH = -log[H3O^+] = -log[0.1M] = 1$

The pH of the solution is 1.

2. 0.1 M NaOH

$NaOH(s) \rightarrow Na^+(aq) + OH^-(aq)$
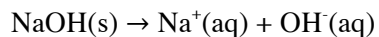
This reaction goes to completion because NaOH is a strong base. So, all the NaOH disassociates into OH⁻ and Na⁺ ions.

Calculate the pH

Step 1: What is left in solution?

In RICE tables, we need to convert all concentrations into moles. To make the calculations of concentrations easier later in the problem, we assume a volume of 1 L of 0.1 M HCl. This way, the number of moles and molarity is the same value (they just have different units).

R	$NaOH(s) \rightarrow$	$Na^+(aq) +$	$OH^-(aq)$
I	0.1	0	0
C	-0.1	+0.1	+0.1
E	0	0.1	0.1

### Step 2: What are the equilibrium concentrations of the species in solution?

The only equilibrium concentrations we are concerned with when calculating the pH of a solution are the concentrations of H3O⁺ ions, OH⁻ ions, weak acids and weak bases. So here we have [OH⁻] = 0.1 M.

Step 3: What is the pH of the solution?

pH =14 – pOH = 14- (-log[OH⁻] = 14- (-log[0.1] = 14 – 1 = 13

The pH of the solution is 13.

3. 100 mL 0.1 M HCl + 100 mL 0.1 M NaOH
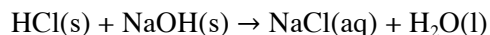
$HCl(s) + NaOH(s) \rightarrow NaCl(aq) + H_2O(l)$

This reaction goes to completion because HCl and NaOH are strong acids and bases. So, all the NaOH disassociates into OH^- and Na^+ ions and all the HCl disassociates into Cl^- and $H3O^+$ ions.

There is no limiting reactant because the HCl and NaOH are added in stoichiometric amounts.

Calculate the pH

Step 1: What is left in solution?

In RICE tables, we need to convert all concentrations into moles.

$n(HCl) = c*V = 0.1 \text{ mol/L} * 0.1 \text{ L} = 0.01$ moles
$n(NaOH) = c*V = 0.1 \text{ mol/L} * 0.1 \text{ L} = 0.01$ moles

R	HCl(s) +	NaOH(s) →	NaCl(aq) +	H20(l)
I	0.01	0.01	0	-
C	-0.01	-0.01	+0.01	-
E	0	0	0.01	-

### Step 2: What are the equilibrium concentrations of the species in solution?

The only equilibrium concentrations we are concerned with when calculating the pH of a solution are the concentrations of $H3O^+$ ions, OH^- ions, weak acids and weak bases. Here we don't have any of these species in solution. We just have water which autoionizes, $[OH^-] = 1*10^{-7}$ M and $[H3O^+] = 1*10^{-7}$ M.

Step 3: What is the pH of the solution?

$pH = -\log[H3O^+] = -\log[1*10^{-7} \text{ M}] = 7$

The pH of the solution is 7. This makes sense because this is a neutralization reaction where the acid and base are added in the same amounts and cancel each other out. You could arrive at this result without doing the actual pH calculation above because we know that pure water is neutral and has a pH of 7.

4. 200 mL 0.1 M HCl + 100 mL 0.1 M NaOH
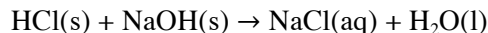
$HCl(s) + NaOH(s) \rightarrow NaCl(aq) + H_2O(l)$

This reaction goes to completion because HCl and NaOH are strong acids and bases. So, all the NaOH disassociates into OH^- and Na^+ ions and all the HCl disassociates into Cl^- and $H3O^+$ ions.

There is a limiting reactant in this problem and it is NaOH.

Calculate the pH

Step 1: What is left in solution?

In RICE tables, we need to convert all concentrations into moles.

$n(HCl) = c*V = 0.1$ mol/L $* 0.2$ L $= 0.02$ moles
$n(NaOH) = c*V = 0.1$ mol/L $* 0.1$ L $= 0.01$ moles

R	HCl(s) +	NaOH(s) →	NaCl(aq) +	H20(l)
I	0.02	0.01	0	-
C	-0.01	-0.01	+0.01	-
E	0.01	0	0.01	-

So then we have,

R	HCl(s) +	H2O(l) →	Cl⁻(aq) +	H3O⁺(aq)
I	0.01	-	0	0
C	-0.01	-	+0.01	+0.01
E	0	-	0.01	0.01

Step 2: What are the equilibrium concentrations of the species in solution?

The only equilibrium concentrations we are concerned with when calculating the pH of a solution are the concentrations of $H3O^+$ ions, OH^- ions, weak acids and weak bases. So here we have $[H3O^+] = n/V = 0.01$ moles / 0.3 L = 0.033 M.
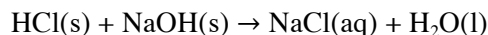
Note we had to convert back to concentrations from moles. The volume of the solution is 0.3 L because 200 mL was added to 100 mL.

Step 3: What is the pH of the solution?

$pH = -log[H3O^+] = -log[0.033$ M$] = 1.477$

The pH of the solution is 1.477.

5. 100 mL 0.1 M HCl + 200 mL of 0.1 M NaOH

$HCl(s) + NaOH(s) \rightarrow NaCl(aq) + H_2O(l)$

This reaction goes to completion because HCl and NaOH are strong acids and bases. So, all the NaOH disassociates into OH^- and Na^+ ions and all the HCl disassociates into Cl^- and $H3O^+$ ions.

There is a limiting reactant in this problem and it is HCl.

Calculate the pH

Step 1: What is left in solution?

In RICE tables, we need to convert all concentrations into moles.

$n(HCl) = c*V = 0.1$ mol/L $* 0.1$ L $= 0.01$ moles
$n(NaOH) = c*V = 0.1$ mol/L $* 0.2$ L $= 0.02$ moles

R	HCl(s) +	NaOH(s) →	NaCl(aq) +	H20(l)
I	0.01	0.02	0	-
C	-0.01	-0.01	+0.01	-
E	0	0.01	0.01	-

So then we have,

R	NaOH(s) →	Na^+(aq) +	OH^-(aq)
I	0.01	0	0
C	-0.01	+0.01	+0.01
E	0	0.01	0.01

Step 2: What are the equilibrium concentrations of the species in solution?

The only equilibrium concentrations we are concerned with when calculating the pH of a solution are the concentrations of $H3O^+$ ions, OH^- ions, weak acids and weak bases. So here we have $[OH^-] = n/V = 0.01$ moles / 0.3 L = 0.033 M.
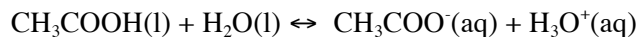
Note we had to convert back to concentrations from moles. The volume of the solution is 0.3 L because 200 mL was added to 100 mL.

Step 3: What is the pH of the solution?

$pH = 14 - pOH = 14 - (-log[OH^-]) = 14 - (-log[0.033$ M$]) = 14 - 1.477 = 12.523$

The pH of the solution is 12.523.

6. 0.1 M CH$_3$COOH

CH$_3$COOH(l) + H$_2$O(l) $\leftrightarrow$ CH$_3$COO$^-$(aq) + H$_3$O$^+$(aq)

This reaction does not go to completion because acetic acid is a weak acid. So this reaction will reach an equilibrium state associated with K$_a$ of acetic acid. K$_a$(CH$_3$COOH) = 1.8 * 10^{-5}.

Because this reaction does not go to completion, we do not have a limiting reactant.

Calculate the pH

Step 1: What is left in solution?

In RICE tables, we need to convert all concentrations into moles. To make the calculations of concentrations easier later in the problem, we assume a volume of 1 L of 0.1 M CH$_3$COOH. This way, the number of moles and molarity is the same value (they just have different units).

R	CH$_3$COOH(l)+	H$_2$O(l) $\leftrightarrow$	CH$_3$COO$^-$(aq) +	H$_3$O$^+$(aq)
I	0.1	-	0	0
C	-x	-	+x	+x
E	0.1-x(~0.1)	-	x	x

Step 2: What are the equilibrium concentrations of the species in solution?

We calculate the equilibrium concentrations using K$_a$(CH$_3$COOH). The equilibrium concentration of CH$_3$COOH is about equal to 0.1. We can ignore the –x because it is so small which we know due to the very small K$_a$ value.

$$K_a(CH_3COOH) = 1.8 * 10^{-5} = \frac{x^2}{0.1} \Rightarrow x = \sqrt{(1.8 * 10^{-5}) * (0.1)} = 1.34 * 10^{-3}$$

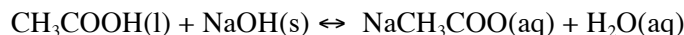So here we have [H$_3$O$^+$] = 1.34*10^{-3} M

Step 3: What is the pH of the solution?

pH = -log[H3O$^+$] = -log[1.34*10^{-3} M] = 2.87

The pH of the solution is 2.87.

7. 100 mL 0.1 M CH_3COOH + 100 mL 0.1 M NaOH

$CH_3COOH(l) + NaOH(s) \leftrightarrow NaCH_3COO(aq) + H_2O(aq)$

This reaction goes to completion because NaOH is a strong base and dissociates fully. The OH^- pulls all the H^+ off of the CH_3COOH that it can get.

There is no limiting reactant in this problem because the two reactants are added in stoichiometric quantities.

Then we have,

$NaCH_3COO(s) \rightarrow Na^+(aq) + CH_3COO^-(aq)$
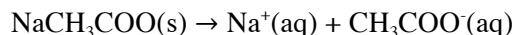
This reaction goes to completion because $NaCH_3COO$ is soluble in water.

After this reaction we are left with CH_3COO^- ions in solution. CH_3COO^- is a weak base and thus will reach an equilibrium state associated with K_b of acetate and that has the reaction equation stated below.

$CH_3COO^-(aq) + H_2O(l) \leftrightarrow CH_3COOH(aq) + OH^-(aq)$
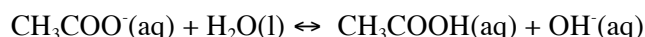
Calculate the pH

Step 1: What is left in solution?

In RICE tables, we need to convert all concentrations into moles.

$n(CH_3COOH) = n(NaOH) = c*V = 0.1$ mol/L $* 0.1$ L $= 0.01$ moles

R	$CH_3COOH(l)+$	$NaOH(s) \leftrightarrow$	$NaCH_3COO(aq) +$	$H_2O(aq)$
I	0.01	0.01	0	-
C	-0.01	-0.01	+0.01	-
E	0	0	0.01	-

Then,

R	$NaCH_3COO(s) \rightarrow$	$Na^+(aq) +$	$CH_3COO^-(aq)$
I	0.01	0	0
C	-0.01	+0.01	+0.01
E	0	0.01	0.01

Finally,

R	$CH_3COO^-(aq) + H_2O(l) \leftrightarrow$		$CH_3COOH(aq)+$	$OH^-($ aq)
I	0.01	-	0	0
C	-x	-	+x	+x
E	0.01-x(~0.01)	-	x	x

Step 2: What are the equilibrium concentrations of the species in solution?

We only need to concern ourselves with the last RICE table since that is what is left in solution. We calculate the equilibrium concentrations using $K_b(CH_3COO^-)$. Because acetate is the conjugate base of acetic acid, $K_b(CH_3COO^-)=K_w/K_a(CH_3COOH)=1*10^{-14}/1.8*10^{-5}= 5.56*10^{-10}$. The equilibrium number of moles of CH_3COO^- is about equal to 0.01. We can ignore the $-x$ because it is so small which we know

due to the very small K_b value. So the equilibrium concentration of CH_3COO^- is $[CH_3COO^-] = n/V = 0.01$ moles/0.2 L = 0.05 M.

$$K_b(CH_3COO^-) = 5.56 * 10^{-10} = \frac{x^2}{0.05} \Rightarrow x = \sqrt{(5.56 * 10^{-10}) * (0.05)} = 5.27 * 10^{-6}$$

So here we have $[OH^-] = 5.27*10^{-6}$ M
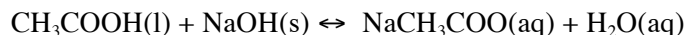
Step 3: What is the pH of the solution?

pH = 14 – pOH = 14- (-log$[OH^-]$) = 14- (-log[5.27*10^{-6} M]) = 14 – 5.28 = 8.72

The pH of the solution is 8.72.

8. 100 mL 0.1 M CH₃COOH + 200 mL 0.1 M NaOH

$CH_3COOH(l) + NaOH(s) \leftrightarrow NaCH_3COO(aq) + H_2O(aq)$

This reaction goes to completion because NaOH is a strong base and dissociates fully. The OH^- pulls all the H^+ off of the CH_3COOH that it can get.

Here the limiting reactant. It is CH_3COOH. So there is left over NaOH.

We then have two things going on
- NaCH₃COO is dissociating and forming CH_3COO^- which is a weak base. The CH_3COO^- reaches equilibrium. However, the amount of OH^- produced is negligible compared to the amount of OH^- produced by the left over NaOH. So we ignore the effect of this on the pH of the solution.
    - $NaCH_3COO(s) \rightarrow Na^+(aq) + CH_3COO^-(aq)$
    - $CH_3COO^-(aq) + H_2O(l) \leftrightarrow CH_3COOH(aq) + OH^-(aq)$
- The leftover NaOH dissociated 100% forming producing OH^- ions and making the solution strongly basic.
    - $NaOH(s) \rightarrow Na^+(aq) + OH^-(aq)$

Calculate the pH

Step 1: What is left in solution?

In RICE tables, we need to convert all concentrations into moles.

$n(CH_3COOH) = c*V = 0.1$ mol/L $* 0.1$ L $= 0.01$ moles
$n(NaOH) = c*V = 0.1$ mol/L $* 0.2$ L $= 0.02$ moles

R	$CH_3COOH(l)+$	$NaOH(s) \leftrightarrow$	$NaCH_3COO(aq) +$	$H_2O(aq)$
I	0.01	0.02	0	-
C	-0.01	-0.01	+0.01	-
E	0	0.01	0.01	-

Then,

R	$NaOH(s) \rightarrow$	$Na^+(aq) +$	$OH^-(aq)$
I	0.01	0	0
C	-0.01	+0.01	+0.01
E	0	0.01	0.01

Step 2: What are the equilibrium concentrations of the species in solution?

The only equilibrium concentrations we are concerned with when calculating the pH of a solution are the concentrations of $H3O^+$ ions, OH^- ions, weak acids and weak bases. Again, here we can ignore the presence of the weak base in solution because the amounts of OH^- it produces is negligible compared to those produced by the 0.01 mole of NaOH. So here we have $[OH^-] = n/V = 0.01$ moles/0.3 L $= 0.033$ M.
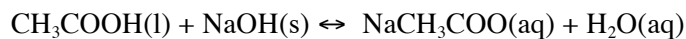
Note we had to convert back to concentrations from moles. The volume of the solution is 0.3 L because 200 mL was added to 100 mL.

Step 3: What is the pH of the solution?

pH $=14 - pOH = 14- (-log[OH^-]) = 14- (-log[0.033$ M$]) = 14 - 1.477 = 12.523$

The pH of the solution is 12.523.

9. 200 mL 0.1 M CH₃COOH + 100 mL 0.1 M NaOH

$CH_3COOH(l) + NaOH(s) \leftrightarrow NaCH_3COO(aq) + H_2O(aq)$

This reaction goes to completion because NaOH is a strong base and dissociates fully. The OH^- pulls all the H^+ off of the CH_3COOH that it can get. $NaCH_3COO$ is formed and dissociates completely forming CH_3COO^- which is a weak base. The limiting reactant is NaOH. So there is left over CH_3COOH which is a weak acid.

Both the CH_3COO^- and CH_3COOH are weak acids and bases, so they establish an equilibrium corresponding to their K_b or K_a respectively.

Calculate the pH

Step 1: What is left in solution?

In RICE tables, we need to convert all concentrations into moles.

$n(CH_3COOH) = c*V = 0.1$ mol/L $* 0.2$ L $= 0.02$ moles
$n(NaOH) = c*V = 0.1$ mol/L $* 0.1$ L $= 0.01$ moles

R	$CH_3COOH(l)+$	$NaOH(s) \leftrightarrow$	$NaCH_3COO(aq) +$	$H_2O(aq)$
I	0.02	0.01	0	-
C	-0.01	-0.01	+0.01	-
E	0,01	0	0.01	-

Then,

R	$NaCH_3COO(s) \rightarrow$	$Na^+(aq) +$	$CH_3COO^-(aq)$
I	0.01	0	0
C	-0.01	+0.01	+0.01
E	0	0.01	0.01

After these two reactions go to completion, the following two equilibria are established.

R	$CH_3COO^-(aq) + H_2O(l) \leftrightarrow$		$CH_3COOH(aq)+$	$OH^-(aq)$
I	0.01	-	0	0
C	-x	-	+x	+x
E	0.01-x(~0.01)	-	x	x

R	$CH_3COOH(l)+$	$H_2O(l) \leftrightarrow$	$CH_3COO^-(aq) + H_3O^+(aq)$	
I	0.01	-	0	0
C	-x	-	+x	+x
E	0.01-x(~0.01)	-	x	x

Step 2: What are the equilibrium concentrations of the species in solution?

We can calculate the equilibrium concentrations using either $K_a(CH_3COOH)$ or $K_b(CH_3COO^-)$. The equilibrium number of moles of CH_3COOH and CH_3COO^- are about equal to 0.01. We can ignore the $-x$ because it is so small which we know due to the very small K_a and K_b value. So,
$[CH_3COO^-] = n/V = 0.01$ moles/0.3 L $= 0.033$ M
$[CH_3COOH] = n/V = 0.01$ moles/0.3 L $= 0.033$ M.

The thing that is different about this problem is that we know the concentrations of both CH_3COOH and CH_3COO^- and are solving for only the concentration of either H_3O^+ or OH^- ions. We choose to use $K_a(CH_3COOH)$ because then we can solve for $[H_3O^+]$ and then be able to calculate pH more directly.

$$K_a(CH_3COOH) = 1.8 * 10^{-5} = \frac{[H_3O^+] * [CH_3COO^-]}{[CH_3COOH]} = \frac{[H_3O^+] * (0.033)}{(0.033)}$$

$$\Rightarrow [H_3O^+] = \frac{(1.8 * 10^{-5}) * (0.033)}{(0.033)} = 1.8 * 10^{-5}$$

So here we have $[H_3O^+] = 1.8*10^{-5}$ M

Step 3: What is the pH of the solution?

$pH = -\log[H3O^+] = -\log[1.8*10^{-5} M] = 4.74$

The pH of the solution is 4.74.

10. 0.1 M NaCH₃COO

$NaCH_3COO(s) \rightarrow Na^+(aq) + CH_3COO^-(aq)$
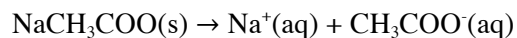
This reaction goes to completion because $NaCH_3COO$ is soluble salt in water.

After this reaction we are left with CH_3COO^- ions in solution. CH_3COO^- is a weak base and thus will reach an equilibrium state associated with K_b of acetate and that has the reaction equation stated below.

$CH_3COO^-(aq) + H_2O(l) \leftrightarrow CH_3COOH(aq) + OH^-(aq)$
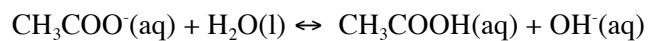
Calculate the pH

Step 1: What is left in solution?

In RICE tables, we need to convert all concentrations into moles. To make the calculations of concentrations easier later in the problem, we assume a volume of 1 L of 0.1 M $NaCH_3COO$. This way, the number of moles and molarity is the same value (they just have different units).

R	$NaCH_3COO(s) \rightarrow$	$Na^+(aq) +$	$CH_3COO^-(aq)$
I	0.1	0	0
C	-0.1	+0.1	+0.1
E	0	0.1	0.1

Then,

R	$CH_3COO^-(aq) + H_2O(l) \leftrightarrow$		$CH_3COOH(aq)+$	$OH^-(\ aq)$
I	0.1	-	0	0
C	-x	-	+x	+x
E	0.1-x(~0.1)	-	x	x

Step 2: What are the equilibrium concentrations of the species in solution?

We only need to concern ourselves with the last RICE table since that is what is left in solution. We calculate the equilibrium concentrations using $K_b(CH_3COO^-)$. Because acetate is the conjugate base of acetic acid, $K_b(CH_3COO^-)=K_w/K_a(CH_3COOH)=1*10^{-14}/1.8*10^{-5}= 5.56*10^{-10}$. The equilibrium number of moles of CH_3COO^- is about equal to 0.1. We can ignore the –x because it is so small which we know due to the very small K_b value. So the equilibrium concentration of CH_3COO^- is $[CH_3COO^-] = 0.1M$.

$$K_b(CH_3COO^-) = \ 5.56 \ * \ 10^{-10} = \frac{x^2}{0.1} \Rightarrow x = \ \sqrt{(5.56 \ * \ 10^{-10}) * (0.1)} = 7.46 * 10^{-6}$$

So here we have $[OH^-] = 7.46*10^{-6}$ M

Step 3: What is the pH of the solution?

pH $=14 – pOH = 14- (-log[OH^-]) = 14- (-log[7.46*10^{-6}$ M]$) = 14 – 5.13 = 8.87$

The pH of the solution is 8.87.

11. 0.1 M NH₄Cl

$NH_4Cl(s) \rightarrow \ NH_4^+(aq) + Cl^-(aq)$

This reaction goes to completion because NH_4Cl is soluble salt in water.

After this reaction we are left with NH_4^+ ions in solution. NH_4^+ is a weak acid and thus will reach an equilibrium state associated with K_a of ammonium and that has the reaction equation stated below.
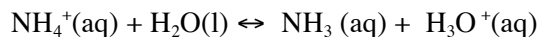
$NH_4^+(aq) + H_2O(l) \leftrightarrow \ NH_3 (aq) + \ H_3O^+(aq)$

Calculate the pH

Step 1: What is left in solution?

In RICE tables, we need to convert all concentrations into moles. To make the calculations of concentrations easier later in the problem, we assume a volume of 1 L of 0.1 M NH_4Cl. This way, the number of moles and molarity is the same value (they just have different units).

R	$NH_4Cl(s) \rightarrow$	$NH_4^+(aq) +$	$Cl^-(aq)$
I	0.1	0	0
C	-0.1	+0.1	+0.1
E	0	0.1	0.1

Then,

R	NH_4^+(aq) +	H_2O(l) ⟷	NH_3 (aq) +	H_3O^+(aq)
I	0.1	-	0	0
C	-x	-	+x	+x
E	0.1-x(~0.1)	-	x	x

Step 2: What are the equilibrium concentrations of the species in solution?

We only need to concern ourselves with the last RICE table since that is what is left in solution. We calculate the equilibrium concentrations using K_a(NH_4^+). The equilibrium concentration of NH_4^+ is about equal to 0.1. We can ignore the –x because it is so small which we know due to the very small K_a value.

$$K_a(NH_4^+) = 5.75 * 10^{-10} = \frac{x^2}{0.1} \Rightarrow x = \sqrt{(5.75 * 10^{-10}) * (0.1)} = 7.58 * 10^{-6}$$

So here we have [H_3O^+] = 7.58*10⁻⁶ M

Step 3: What is the pH of the solution?

pH = -log[H3O⁺] = -log[7.58*10⁻⁶ M] = 5.12

The pH of the solution is 5.12.