

Acid and Base pH Calculations – Supplemental Worksheet KEY

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

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) + | H ₂ O(l) → | Cl ⁻ (aq) + | H ₃ O ⁺ (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?

$$\text{pH} = -\log[\text{H}_3\text{O}^+] = -\log[0.1\text{M}] = 1$$

The pH of the solution is 1.



2. 0.1 M NaOH



This reaction goes to completion because NaOH is a strong base. So, all the NaOH disassociates into $\text{OH}^{\text{-}}$ and $\text{Na}^{\text{+}}$ 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) → | 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 $\text{H}_3\text{O}^{\text{+}}$ ions, $\text{OH}^{\text{-}}$ ions, weak acids and weak bases. **So here we have $[\text{OH}^{\text{-}}] = 0.1 \text{ M}$.**

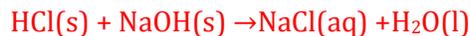
Step 3: What is the pH of the solution?

$$\text{pH} = 14 - \text{pOH} = 14 - (-\log[\text{OH}^{\text{-}}]) = 14 - (-\log[0.1]) = 14 - 1 = 13$$

The pH of the solution is 13.



3. 100 mL 0.1 M HCl + 100 mL 0.1 M NaOH



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 H_3O^+ 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.

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

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

| | | | | |
|---|----------|-----------|------------|---------------------|
| R | HCl(s) + | NaOH(s) → | NaCl(aq) + | H ₂ O(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 H_3O^+ 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, $[\text{OH}^-] = 1 \cdot 10^{-7} \text{ M}$ and $[\text{H}_3\text{O}^+] = 1 \cdot 10^{-7} \text{ M}$.

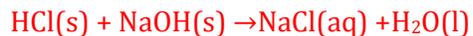
Step 3: What is the pH of the solution?

$$\text{pH} = -\log[\text{H}_3\text{O}^+] = -\log[1 \cdot 10^{-7} \text{ M}] = 7$$

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



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 H_3O^+ 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.

$$n(\text{HCl}) = c \cdot V = 0.1 \text{ mol/L} \cdot 0.2 \text{ L} = 0.02 \text{ moles}$$

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

| | | | | |
|---|----------|-----------|------------|---------------------|
| R | HCl(s) + | NaOH(s) → | NaCl(aq) + | H ₂ O(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) + | H ₂ O(l) → | Cl ⁻ (aq) + | H ₃ O ⁺ (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 H_3O^+ ions, OH^- ions, weak acids and weak bases. So here we have $[\text{H}_3\text{O}^+] = n/V = 0.01 \text{ moles} / 0.3 \text{ L} = 0.033 \text{ M}$.

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.

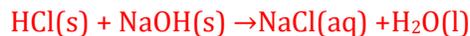
Step 3: What is the pH of the solution?

$$\text{pH} = -\log[\text{H}_3\text{O}^+] = -\log[0.033 \text{ M}] = 1.477$$

The pH of the solution is 1.477.



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



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 H_3O^+ 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.

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

$$n(\text{NaOH}) = c \cdot V = 0.1 \text{ mol/L} \cdot 0.2 \text{ L} = 0.02 \text{ moles}$$

| | | | | |
|---|----------|-----------|------------|---------------------|
| R | HCl(s) + | NaOH(s) → | NaCl(aq) + | H ₂ O(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 H_3O^+ ions, OH^- ions, weak acids and weak bases. So here we have $[\text{OH}^-] = n/V = 0.01 \text{ moles} / 0.3 \text{ L} = 0.033 \text{ M}$.

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?

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

The pH of the solution is 12.523.

6. 0.1 M CH₃COOH



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. $K_a(\text{CH}_3\text{COOH}) = 1.8 * 10^{-5}$.

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 CH₃COOH. 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) ↔ | 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 $K_a(\text{CH}_3\text{COOH})$. The equilibrium concentration of CH₃COOH 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.

$$K_a(\text{CH}_3\text{COOH}) = 1.8 * 10^{-5} = \frac{x^2}{0.1} \Rightarrow x = \sqrt{(1.8 * 10^{-5}) * (0.1)} = 1.34 * 10^{-3}$$

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

Step 3: What is the pH of the solution?

$$\text{pH} = -\log[\text{H}_3\text{O}^+] = -\log[1.34 * 10^{-3} \text{ M}] = 2.87$$

The pH of the solution is 2.87.



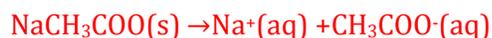
7. 100 mL 0.1 M CH₃COOH + 100 mL 0.1 M NaOH



This reaction goes to completion because NaOH is a strong base and dissociates fully. The OH⁻ pulls all the H⁺ off of the CH₃COOH that it can get.

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

Then we have,



This reaction goes to completion because NaCH₃COO is soluble in water.

After this reaction we are left with CH₃COO⁻ ions in solution. CH₃COO⁻ 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.



Calculate the pH

Step 1: What is left in solution?

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

$$n(\text{CH}_3\text{COOH}) = n(\text{NaOH}) = c \cdot V = 0.1 \text{ mol/L} \cdot 0.1 \text{ L} = 0.01 \text{ moles}$$

| R | CH ₃ COOH(l)+ | NaOH(s) ↔ | NaCH ₃ COO(aq) + | H ₂ O(aq) |
|---|--------------------------|-----------|-----------------------------|----------------------|
| I | 0.01 | 0.01 | 0 | - |
| C | -0.01 | -0.01 | +0.01 | - |
| E | 0 | 0 | 0.01 | - |

Then,

| R | NaCH ₃ COO(s) → | Na ⁺ (aq) + | CH ₃ COO ⁻ (aq) |
|---|----------------------------|------------------------|---------------------------------------|
| I | 0.01 | 0 | 0 |
| C | -0.01 | +0.01 | +0.01 |
| E | 0 | 0.01 | 0.01 |

Finally,

| R | CH ₃ COO ⁻ (aq) + | H ₂ O(l) ↔ | CH ₃ COOH(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 $K_b(\text{CH}_3\text{COO}^-)$. Because acetate is the conjugate base of acetic acid, $K_b(\text{CH}_3\text{COO}^-) = K_w/K_a(\text{CH}_3\text{COOH}) = 1 \times 10^{-14}/1.8 \times 10^{-5} = 5.56 \times 10^{-10}$. 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 $[\text{CH}_3\text{COO}^-] = n/V = 0.01 \text{ moles}/0.2 \text{ L} = 0.05 \text{ M}$.

$$K_b(\text{CH}_3\text{COO}^-) = 5.56 \times 10^{-10} = \frac{x^2}{0.05} \Rightarrow x = \sqrt{(5.56 \times 10^{-10}) * (0.05)} = 5.27 \times 10^{-6}$$

So here we have $[\text{OH}^-] = 5.27 \times 10^{-6} \text{ M}$

Step 3: What is the pH of the solution?

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

The pH of the solution is 8.72.



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



This reaction goes to completion because NaOH is a strong base and dissociates fully. The OH⁻ pulls all the H⁺ off of the CH₃COOH that it can get.

Here the limiting reactant. It is CH₃COOH. So there is left over NaOH.

We then have two things going on

- NaCH₃COO is dissociating and forming CH₃COO⁻ which is a weak base. The CH₃COO⁻ 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.
 - NaCH₃COO(s) → Na⁺(aq) + CH₃COO⁻(aq)
 - CH₃COO⁻(aq) + H₂O(l) ↔ CH₃COOH(aq) + OH⁻(aq)
- The leftover NaOH dissociated 100% forming producing OH⁻ ions and making the solution strongly basic.
 - NaOH(s) → 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.

$$n(\text{CH}_3\text{COOH}) = c \cdot V = 0.1 \text{ mol/L} \cdot 0.1 \text{ L} = 0.01 \text{ moles}$$

$$n(\text{NaOH}) = c \cdot V = 0.1 \text{ mol/L} \cdot 0.2 \text{ L} = 0.02 \text{ moles}$$

| R | CH ₃ COOH(l) | + NaOH(s) ↔ | NaCH ₃ COO(aq) + | H ₂ O(aq) |
|---|-------------------------|-------------|-----------------------------|----------------------|
| I | 0.01 | 0.02 | 0 | - |
| C | -0.01 | -0.01 | +0.01 | - |
| E | 0 | 0.01 | 0.01 | - |

Then,

| 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. 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?

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

The pH of the solution is 12.523.

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



This reaction goes to completion because NaOH is a strong base and dissociates fully. The OH⁻ pulls all the H⁺ off of the CH₃COOH that it can get. NaCH₃COO is formed and dissociates completely forming CH₃COO⁻ which is a weak base. The limiting reactant is NaOH. So there is left over CH₃COOH which is a weak acid.

Both the CH₃COO⁻ and CH₃COOH 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.

$$n(\text{CH}_3\text{COOH}) = c \cdot V = 0.1 \text{ mol/L} \cdot 0.2 \text{ L} = 0.02 \text{ moles}$$

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

| R | CH ₃ COOH(l)+ | NaOH(s) ↔ | NaCH ₃ COO(aq) + | H ₂ O(aq) |
|---|--------------------------|-----------|-----------------------------|----------------------|
| I | 0.02 | 0.01 | 0 | - |
| C | -0.01 | -0.01 | +0.01 | - |
| E | 0.01 | 0 | 0.01 | - |

Then,

| R | NaCH ₃ COO(s) → | Na ⁺ (aq) + | CH ₃ COO ⁻ (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 ₃ COO ⁻ (aq) + | H ₂ O(l) ↔ | CH ₃ COOH(aq)+ | OH ⁻ (aq) |
|---|---|-----------------------|---------------------------|----------------------|
| I | 0.01 | - | 0 | 0 |
| C | -x | - | +x | +x |
| E | 0.01-x(~0.01) | - | x | x |

| R | CH ₃ COOH(l)+ | H ₂ O(l) ↔ | CH ₃ COO ⁻ (aq) + | H ₃ O ⁺ (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 $K_a(\text{CH}_3\text{COOH})$ or $K_b(\text{CH}_3\text{COO}^-)$. 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,

$$[\text{CH}_3\text{COO}^-] = n/V = 0.01 \text{ moles}/0.3 \text{ L} = 0.033 \text{ M}$$

$$[\text{CH}_3\text{COOH}] = n/V = 0.01 \text{ moles}/0.3 \text{ L} = 0.033 \text{ M}.$$

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 $K_a(\text{CH}_3\text{COOH})$ because then we can solve for $[\text{H}_3\text{O}^+]$ and then be able to calculate pH more directly.

$$K_a(\text{CH}_3\text{COOH}) = 1.8 * 10^{-5} = \frac{[\text{H}_3\text{O}^+] * [\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]} = \frac{[\text{H}_3\text{O}^+] * (0.033)}{(0.033)}$$

$$\Rightarrow [\text{H}_3\text{O}^+] = \frac{(1.8 * 10^{-5}) * (0.033)}{(0.033)} = 1.8 * 10^{-5}$$

So here we have $[\text{H}_3\text{O}^+] = 1.8 * 10^{-5} \text{ M}$

Step 3: What is the pH of the solution?

$$\text{pH} = -\log[\text{H}_3\text{O}^+] = -\log[1.8 * 10^{-5} \text{ M}] = 4.74$$

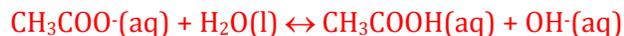
The pH of the solution is 4.74.

10.0.1 M NaCH₃COO



This reaction goes to completion because NaCH₃COO is soluble salt in water.

After this reaction we are left with CH₃COO⁻ ions in solution. CH₃COO⁻ 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.



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 NaCH₃COO. This way, the number of moles and molarity is the same value (they just have different units).

| | | | |
|---|----------------------------|------------------------|---------------------------------------|
| R | NaCH ₃ COO(s) → | Na ⁺ (aq) + | CH ₃ COO ⁻ (aq) |
| I | 0.1 | 0 | 0 |
| C | -0.1 | +0.1 | +0.1 |
| E | 0 | 0.1 | 0.1 |

Then,

| | | | | |
|---|---|-----------------------|---------------------------|----------------------|
| R | CH ₃ COO ⁻ (aq) + | H ₂ O(l) ↔ | CH ₃ COOH(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(CH₃COO⁻). Because acetate is the conjugate base of acetic acid, K_b(CH₃COO⁻)=K_w/K_a(CH₃COOH)=1*10⁻¹⁴/1.8*10⁻⁵= 5.56*10⁻¹⁰. The equilibrium number of moles of CH₃COO⁻ 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 CH₃COO⁻ is [CH₃COO⁻] = 0.1M.

$$K_b(\text{CH}_3\text{COO}^-) = 5.56 * 10^{-10} = \frac{x^2}{0.1} \Rightarrow x = \sqrt{(5.56 * 10^{-10}) * (0.1)} = 7.46 * 10^{-6}$$

So here we have [OH⁻] = 7.46*10⁻⁶ M

Step 3: What is the pH of the solution?



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

The pH of the solution is 8.87.

11.0.1 M NH₄Cl



This reaction goes to completion because NH₄Cl is soluble salt in water.

After this reaction we are left with NH₄⁺ ions in solution. NH₄⁺ 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.



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 NH₄Cl. This way, the number of moles and molarity is the same value (they just have different units).

| | | | |
|---|-------------------------|-------------------------------------|----------------------|
| R | NH ₄ Cl(s) → | 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 ₂ O(l) ↔ | NH ₃ (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 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(NH₄⁺). The equilibrium concentration of NH₄⁺ 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.

$$K_a(\text{NH}_4^+) = 5.75 \times 10^{-10} = \frac{x^2}{0.1} \Rightarrow x = \sqrt{(5.75 \times 10^{-10}) * (0.1)} = 7.58 \times 10^{-6}$$

So here we have [H₃O⁺] = 7.58*10⁻⁶ M

Step 3: What is the pH of the solution?

$$\text{pH} = -\log[\text{H}_3\text{O}^+] = -\log[7.58 \times 10^{-6} \text{ M}] = 5.12$$

The pH of the solution is 5.12.