



### **Extra Practice: Acid and Base pH Calculations 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<sup>-</sup> and H<sub>3</sub>O<sup>+</sup> 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 <sub>2</sub> O(l) →	Cl <sup>-</sup> (aq) +	H <sub>3</sub> O <sup>+</sup> (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 H<sub>3</sub>O<sup>+</sup> ions, OH<sup>-</sup> ions, weak acids and weak bases. **So here we have [H<sub>3</sub>O<sup>+</sup>] = 0.1 M.**

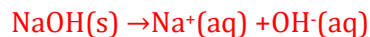
#### **Step 3: What is the pH of the solution?**

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

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 <sup>+</sup> (aq) +	OH <sup>-</sup> (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  $[\text{OH}^{\text{-}}] = 0.1 \text{ M}$ .**

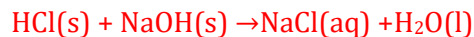
### Step 3: What is the pH of the solution?

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

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  $\text{OH}^-$  and  $\text{Na}^+$  ions and all the HCl disassociates into  $\text{Cl}^-$  and  $\text{H}_3\text{O}^+$  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 <sub>2</sub> 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  $\text{H}_3\text{O}^+$  ions,  $\text{OH}^-$  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  $[\text{H}_3\text{O}^+] = 1 \cdot 10^{-7} \text{ M}$ .

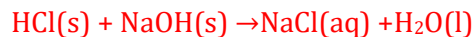
#### Step 3: What is the pH of the solution?

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

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  $\text{OH}^-$  and  $\text{Na}^+$  ions and all the HCl disassociates into  $\text{Cl}^-$  and  $\text{H}_3\text{O}^+$  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 <sub>2</sub> 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 <sub>2</sub> O(l) →	Cl <sup>-</sup> (aq) +	H <sub>3</sub> O <sup>+</sup> (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  $\text{H}_3\text{O}^+$  ions,  $\text{OH}^-$  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.

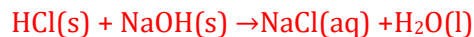
Step 3: What is the pH of the solution?

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

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  $\text{OH}^-$  and  $\text{Na}^+$  ions and all the HCl disassociates into  $\text{Cl}^-$  and  $\text{H}_3\text{O}^+$  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 <sub>2</sub> 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 <sup>+</sup> (aq) +	OH <sup>-</sup> (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  $\text{H}_3\text{O}^+$  ions,  $\text{OH}^-$  ions, weak acids and weak bases. So here we have  $[\text{OH}^-] = n/V = 0.01 \text{ moles} / 0.3 \text{ L} = 0.033 \text{ M}$ .

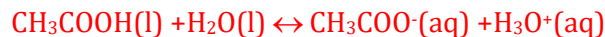
Note we had to convert back to concentrations from moles. The volume of the solution is 0.3 L because 200 mL was added to 100 mL.

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 CH<sub>3</sub>COOH



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 CH<sub>3</sub>COOH. This way, the number of moles and molarity is the same value (they just have different units).

R	CH <sub>3</sub> COOH(l)+	H <sub>2</sub> O(l) ↔	CH <sub>3</sub> COO <sup>-</sup> (aq) +	H <sub>3</sub> O <sup>+</sup> (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 CH<sub>3</sub>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.

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

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.



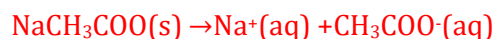
**7. 100 mL 0.1 M CH<sub>3</sub>COOH + 100 mL 0.1 M NaOH**



This reaction goes to completion because NaOH is a strong base and dissociates fully. The OH<sup>-</sup> pulls all the H<sup>+</sup> off of the CH<sub>3</sub>COOH 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 NaCH<sub>3</sub>COO is soluble in water.

After this reaction we are left with CH<sub>3</sub>COO<sup>-</sup> ions in solution. CH<sub>3</sub>COO<sup>-</sup> is a weak base and thus will reach an equilibrium state associated with K<sub>b</sub> 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 <sub>3</sub> COOH(l)+	NaOH(s) ↔	NaCH <sub>3</sub> COO(aq) +	H <sub>2</sub> O(aq)
I	0.01	0.01	0	-
C	-0.01	-0.01	+0.01	-
E	0	0	0.01	-

Then,

R	NaCH <sub>3</sub> COO(s) →	Na <sup>+</sup> (aq) +	CH <sub>3</sub> COO <sup>-</sup> (aq)
I	0.01	0	0
C	-0.01	+0.01	+0.01
E	0	0.01	0.01

Finally,

R	CH <sub>3</sub> COO <sup>-</sup> (aq) +	H <sub>2</sub> O(l) ↔	CH <sub>3</sub> COOH(aq)+	OH <sup>-</sup> (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  $\text{CH}_3\text{COO}^-$  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  $\text{CH}_3\text{COO}^-$  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 $\text{CH}_3\text{COOH}$ + 200 mL 0.1 M $\text{NaOH}$



This reaction goes to completion because  $\text{NaOH}$  is a strong base and dissociates fully. The  $\text{OH}^-$  pulls all the  $\text{H}^+$  off of the  $\text{CH}_3\text{COOH}$  that it can get.

Here the limiting reactant. It is  $\text{CH}_3\text{COOH}$ . So there is left over  $\text{NaOH}$ .

We then have two things going on

- $\text{NaCH}_3\text{COO}$  is dissociating and forming  $\text{CH}_3\text{COO}^-$  which is a weak base. The  $\text{CH}_3\text{COO}^-$  reaches equilibrium. However, the amount of  $\text{OH}^-$  produced is negligible compared to the amount of  $\text{OH}^-$  produced by the left over  $\text{NaOH}$ . So we ignore the effect of this on the pH of the solution.
  - $\text{NaCH}_3\text{COO}(\text{s}) \rightarrow \text{Na}^+(\text{aq}) + \text{CH}_3\text{COO}^-(\text{aq})$
  - $\text{CH}_3\text{COO}^-(\text{aq}) + \text{H}_2\text{O}(\text{l}) \leftrightarrow \text{CH}_3\text{COOH}(\text{aq}) + \text{OH}^-(\text{aq})$
- The leftover  $\text{NaOH}$  dissociated 100% forming producing  $\text{OH}^-$  ions and making the solution strongly basic.
  - $\text{NaOH}(\text{s}) \rightarrow \text{Na}^+(\text{aq}) + \text{OH}^-(\text{aq})$

Calculate the pH

Step 1: What is left in solution?

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

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





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

R	$\text{CH}_3\text{COOH(l)} +$	$\text{NaOH(s)} \leftrightarrow$	$\text{NaCH}_3\text{COO(aq)} +$	$\text{H}_2\text{O(aq)}$
I	0.01	0.02	0	-
C	-0.01	-0.01	+0.01	-
E	0	0.01	0.01	-

Then,

R	$\text{NaOH(s)} \rightarrow$	$\text{Na}^{\text{(aq)}} +$	$\text{OH}^{\text{(aq)}}$
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  $\text{H}_3\text{O}^+$  ions,  $\text{OH}^-$  ions, weak acids and weak bases. Again, here we can ignore the presence of the weak base in solution because the amounts of  $\text{OH}^-$  it produces is negligible compared to those produced by the 0.01 mole of NaOH. 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.

**9. 200 mL 0.1 M CH<sub>3</sub>COOH + 100 mL 0.1 M NaOH**



This reaction goes to completion because NaOH is a strong base and dissociates fully. The OH<sup>-</sup> pulls all the H<sup>+</sup> off of the CH<sub>3</sub>COOH that it can get. NaCH<sub>3</sub>COO is formed and dissociates completely forming CH<sub>3</sub>COO<sup>-</sup> which is a weak base. The limiting reactant is NaOH. So there is left over CH<sub>3</sub>COOH which is a weak acid.

Both the CH<sub>3</sub>COO<sup>-</sup> and CH<sub>3</sub>COOH are weak acids and bases, so they establish an equilibrium corresponding to their K<sub>b</sub> or K<sub>a</sub> 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 <sub>3</sub> COOH(l)+	NaOH(s) ↔	NaCH <sub>3</sub> COO(aq) +	H <sub>2</sub> O(aq)
I	0.02	0.01	0	-
C	-0.01	-0.01	+0.01	-
E	0.01	0	0.01	-

Then,

R	NaCH <sub>3</sub> COO(s) →	Na <sup>+</sup> (aq) +	CH <sub>3</sub> COO <sup>-</sup> (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 <sub>3</sub> COO <sup>-</sup> (aq) +	H <sub>2</sub> O(l) ↔	CH <sub>3</sub> COOH(aq)+	OH <sup>-</sup> (aq)
I	0.01	-	0	0
C	-x	-	+x	+x
E	0.01-x(~0.01)	-	x	x

R	CH <sub>3</sub> COOH(l)+	H <sub>2</sub> O(l) ↔	CH <sub>3</sub> COO <sup>-</sup> (aq) +	H <sub>3</sub> O <sup>+</sup> (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  $\text{CH}_3\text{COOH}$  and  $\text{CH}_3\text{COO}^-$  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  $\text{CH}_3\text{COOH}$  and  $\text{CH}_3\text{COO}^-$  and are solving for only the concentration of either  $\text{H}_3\text{O}^+$  or  $\text{OH}^-$  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.

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

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

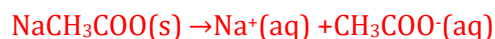
So here we have  $[\text{H}_3\text{O}^+] = 1.8 \times 10^{-5} \text{ M}$

Step 3: What is the pH of the solution?

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

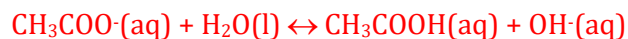
The pH of the solution is 4.74.

### 10.0.1 M $\text{NaCH}_3\text{COO}$



This reaction goes to completion because  $\text{NaCH}_3\text{COO}$  is soluble salt in water.

After this reaction we are left with  $\text{CH}_3\text{COO}^-$  ions in solution.  $\text{CH}_3\text{COO}^-$  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  $\text{NaCH}_3\text{COO}$ . This way, the number of moles and molarity is the same value (they just have different units).

R	$\text{NaCH}_3\text{COO}(s) \rightarrow$	$\text{Na}^+(aq) +$	$\text{CH}_3\text{COO}^-(aq)$
I	0.1	0	0
C	-0.1	+0.1	+0.1



E      0                                      0.1                                      0.1

Then,

R      CH<sub>3</sub>COO<sup>-</sup>(aq) +                      H<sub>2</sub>O(l) ↔                      CH<sub>3</sub>COOH(aq)+                      OH<sup>-</sup>(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  $K_b(\text{CH}_3\text{COO}^-)$ . 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  $\text{CH}_3\text{COO}^-$  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  $\text{CH}_3\text{COO}^-$  is  $[\text{CH}_3\text{COO}^-] = 0.1\text{M}$ .

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

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

Step 3: What is the pH of the solution?

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

The pH of the solution is 8.87.

**11.0.1 M NH<sub>4</sub>Cl**



This reaction goes to completion because  $\text{NH}_4\text{Cl}$  is soluble salt in water.

After this reaction we are left with  $\text{NH}_4^+$  ions in solution.  $\text{NH}_4^+$  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.



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  $\text{NH}_4\text{Cl}$ . This way, the number of moles and molarity is the same value (they just have different units).

R	$\text{NH}_4\text{Cl}(\text{s}) \rightarrow$	$\text{NH}_4^+(\text{aq}) +$	$\text{Cl}^-(\text{aq})$
I	0.1	0	0
C	-0.1	+0.1	+0.1
E	0	0.1	0.1

Then,

R	$\text{NH}_4^+(\text{aq}) +$	$\text{H}_2\text{O}(\text{l}) \leftrightarrow$	$\text{NH}_3(\text{aq}) +$	$\text{H}_3\text{O}^+(\text{aq})$
I	0.1	-	0	0
C	-x	-	+x	+x
E	$0.1-x(\sim 0.1)$	-	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  $K_a(\text{NH}_4^+)$ . The equilibrium concentration of  $\text{NH}_4^+$  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.

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

So here we have  $[\text{H}_3\text{O}^+] = 7.58 * 10^{-6} \text{ M}$

Step 3: What is the pH of the solution?

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

The pH of the solution is 5.12.