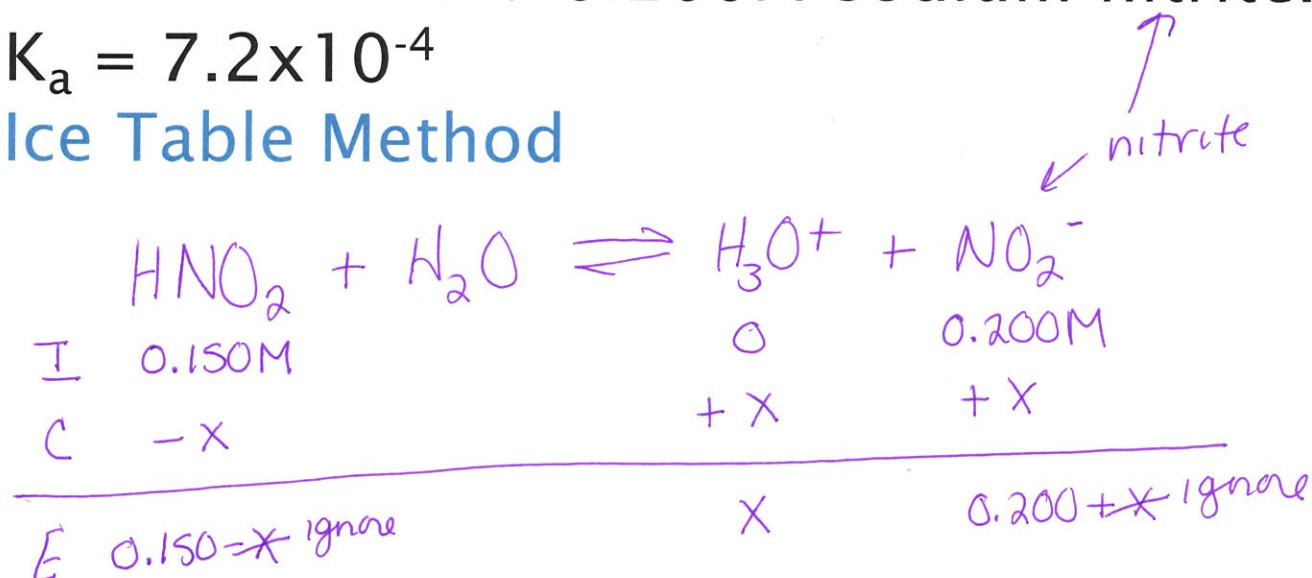


Using the Henderson-Hasselbalch Equation

1. A 1.00L buffer solution is prepared that contains 0.150M nitrous acid and 0.200M sodium nitrite. What is its pH?

$$K_a = 7.2 \times 10^{-4}$$

Ice Table Method



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{NO}_2^-]}{[\text{HNO}_2]} = \frac{[X][0.200]}{[0.150]} = 7.2 \times 10^{-4}$$

$$\frac{(0.200)(X)}{0.200} = \frac{1.08 \times 10^{-4}}{0.200}$$

$$X = 5.4 \times 10^{-4} = [\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log [5.4 \times 10^{-4}] = \boxed{3.27}$$

pH = 3.27 5

Using the Henderson-Hasselbalch Equation

1. A 1.00L buffer solution is prepared that contains 0.150M nitrous acid and 0.200M sodium nitrite. What is its pH?

$$K_a = 7.2 \times 10^{-4}$$

H-H equation method:

$$pH = pK_a + \log \frac{[A^-]}{[HA]} \quad \begin{array}{l} \swarrow \text{conj base} \\ \nwarrow \text{acid} \end{array}$$

$$A^- = \text{conj base} = \text{NaNO}_2 = 0.200\text{M}$$

$$HA = \text{acid} = \text{nitrous acid} = 0.150\text{M}$$

$$pK_a = -\log K_a = -\log(7.2 \times 10^{-4}) = 3.14267$$

$$pH = 3.14267 + \log \frac{[0.200]}{[0.150]} = \boxed{3.27}$$

$$pH = 3.27 \quad 6$$

Using the Henderson-Hasselbalch Equation

2. How many grams of sodium lactate ($\text{CH}_3\text{CH}(\text{OH})\text{COONa}$) should be added to 1.0L of a 0.150M lactic acid ($\text{CH}_3\text{CH}(\text{OH})\text{COOH}$) to form a buffer solution with pH=4.00? $K_a = 1.4 \times 10^{-4}$; molar mass of sodium lactate = 112.1g/mol

$$\text{pH} = \text{pKa} + \log \frac{[\text{A}^-]}{[\text{HA}]}$$

$$4.00 = 3.85387 + \log \frac{[\text{A}^-]}{[0.150]}$$

-3.85387 -3.85387

$$0.14613 = \log \frac{[\text{A}^-]}{[0.150]}$$

$$10^{0.14613} = \frac{[\text{A}^-]}{[0.150]}$$

$$1.400 = \frac{[\text{A}^-]}{[0.150]}$$

$$[\text{A}^-] = 0.2100 \text{ M}$$

$$\text{pKa} = -\log(K_a) = -\log(1.4 \times 10^{-4}) = 3.85387$$

$$\text{HA} = \text{lactic acid} = 0.150 \text{ M}$$

$$\text{pH} = 4.00$$

$$\text{A}^- = \text{sodium lactate [?]}$$

$$\text{pH} = 3$$

$$[\text{H}_3\text{O}^+] = 10^{-3}$$

$$0.2100 \frac{\text{mol}}{\text{L}} \times 1 \text{ L} = 0.2100 \text{ mol} \times \frac{112.1 \text{ g}}{\text{mol}}$$

$$= 23.54 \text{ g}$$

$$\hookrightarrow \boxed{24 \text{ g}}$$

Buffer calculations

1. A 1.0 L buffer solution contains 0.150 M nitrous acid and 0.200 M sodium nitrite. $K_a = 7.2 \times 10^{-4}$ (a) What is the pH of the buffer? (b) What is the pH after adding 1.00 g HBr?

$$\begin{aligned} H &= 1.00794 \text{ g/mol} \\ Br &= 79.904 \text{ g/mol} \\ MM &= 80.91194 \text{ g/mol} \end{aligned}$$

$$pH = pK_a + \log \frac{[A^-]}{[HA]} \quad pK_a = -\log(7.2 \times 10^{-4}) = 3.1427 \quad MM = 80.91194 \text{ g/mol}$$

$$(a) \quad pH = 3.1427 + \log \frac{[0.200]}{[0.150]} = \boxed{3.27}$$

$$\begin{aligned} 1.00 \text{ g HBr} \left(\frac{1 \text{ mol}}{80.91194 \text{ g}} \right) \\ = 0.01236 \text{ mol HBr} \end{aligned}$$

(b) HBr = strong acid $\rightarrow$ react with base A^-



$$NO_2^-: \text{HBr neutralize } NO_2^- \quad 0.01236 \text{ mol HBr} \left(\frac{1 \text{ mol } NO_2^-}{1 \text{ mol HBr}} \right) = 0.01236 \text{ mol } NO_2^- \text{ neutralized}$$

$$0.200 \frac{\text{mol } NO_2^-}{L} \times 1L = 0.200 \text{ mol } NO_2^- - 0.01236 \text{ mol } NO_2^- = 0.18764 \text{ mol } NO_2^- \text{ remaining}$$

$$HNO_2: \text{formed in reaction } 0.01236 \text{ mol HBr} \left(\frac{1 \text{ mol } HNO_2}{1 \text{ mol HBr}} \right)$$

$$= 0.01236 \text{ mol } HNO_2 \text{ formed}$$

$$\begin{aligned} 0.150 \frac{\text{mol } HNO_2}{L} \times 1L &= 0.150 \text{ mol } HNO_2 + 0.01236 \text{ mol } HNO_2 \text{ formed} = \frac{0.16236 \text{ mol } HNO_2}{1.0L} \\ &= 0.16236 \text{ M} \end{aligned}$$

$$\begin{aligned} pH &= 3.1427 + \log \frac{[0.18764]}{[0.16236]} \\ &= 3.1427 + 0.062846 = 3.2055 \rightarrow \boxed{3.21} \end{aligned}$$

A: (a) 3.27

(b) 3.21 12

Buffer calculations

2. A buffer is made by adding 0.600 mol CH_3COOH and 0.600 mol CH_3COONa to enough water to make 2.00L of solution. $K_a = 1.8 \times 10^{-5}$.

(a) What is the pH of the buffer? **A: 4.74**

(b) Calculate the pH after 0.040 mol HCl is added. **A: 4.69**

(c) Calculate the pH after 0.040 mol NaOH is added. **A: 4.80**

$$\text{pH} = \text{p}K_a + \log \frac{[\text{A}^-] \leftarrow \text{CH}_3\text{COO}^-}{[\text{HA}] \leftarrow \text{CH}_3\text{COOH}} \quad \text{p}K_a = -\log(1.8 \times 10^{-5}) = 4.7447$$

$$[\text{A}^-] = \frac{0.600 \text{ mol}}{2.00 \text{ L}} = 0.300 \text{ M}$$

$$[\text{HA}] = \frac{0.600 \text{ mol}}{2.00 \text{ L}} = 0.300 \text{ M}$$

$$\text{pH} = 4.7447 + \log \frac{[0.300]}{[0.300]}$$

$$\text{pH} = 4.7447 + \log(1)$$

$$\log 1 = 0$$

$$\text{pH} = 4.7447 + 0 = 4.7447 \rightarrow \boxed{4.74}$$

Buffer calculations

2. A buffer is made by adding 0.600 mol CH_3COOH and 0.600 mol CH_3COONa to enough water to make 2.00L of solution. $K_a = 1.8 \times 10^{-5}$. $= 4.7447 = \text{pKa}$

(a) What is the pH of the buffer? A: 4.74

(b) Calculate the pH after 0.040 mol HCl is added. A: 4.69



$$\text{pH} = \text{pKa} + \log \frac{[\text{A}^-]}{[\text{HA}]}$$

0.040 mol HCl:

$$\text{A}^- \text{ CH}_3\text{COO}^- \quad 0.040 \text{ mol HCl} \left(\frac{1 \text{ mol CH}_3\text{COO}^-}{1 \text{ mol HCl}} \right) = 0.040 \text{ mol CH}_3\text{COO}^-$$

$$0.600 \text{ mol initial} - 0.040 \text{ mol neutralized} = \frac{0.560 \text{ mol CH}_3\text{COO}^- \text{ remains}}{2.00 \text{ L}}$$

$$= 0.280 \text{ M CH}_3\text{COO}^-$$

$$\text{HA CH}_3\text{COOH} \quad 0.040 \text{ mol HCl} \left(\frac{1 \text{ mol CH}_3\text{COOH}}{1 \text{ mol HCl}} \right) = 0.040 \text{ mol CH}_3\text{COOH}$$

$$0.600 \text{ mol initial} + 0.040 \text{ mol formed} = \frac{0.640 \text{ mol CH}_3\text{COOH}}{2.00 \text{ L}} = 0.320 \text{ M CH}_3\text{COOH}$$

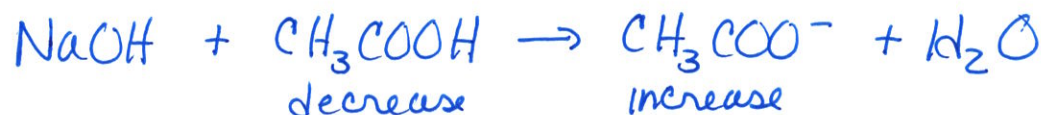
$$\text{pH} = 4.7447 + \log \frac{[0.280]}{[0.320]} = 4.7447 + (-0.05799) = 4.6867 \rightarrow \boxed{4.69}$$

Buffer calculations

2. A buffer is made by adding 0.600 mol CH_3COOH and 0.600 mol CH_3COONa to enough water to make 2.00L of solution. $K_a = 1.8 \times 10^{-5}$. $pK_a = 4.7447$

(a) What is the pH of the buffer? A: 4.74

(c) Calculate the pH after 0.040 mol NaOH is added. A: 4.80



$$\begin{aligned} [\text{HA}] \text{ CH}_3\text{COOH} & \quad 0.040 \text{ mol NaOH} \left(\frac{1 \text{ mol CH}_3\text{COOH}}{1 \text{ mol NaOH}} \right) = 0.040 \text{ mol CH}_3\text{COOH neutralized} \\ 0.600 \text{ mol} - 0.040 \text{ mol} & = \frac{0.560 \text{ mol CH}_3\text{COOH remains}}{2.00 \text{ L}} = 0.280 \text{ M CH}_3\text{COOH} \end{aligned}$$

$$\begin{aligned} [\text{A}^-] \text{ CH}_3\text{COO}^- & \quad 0.040 \text{ mol NaOH} \left(\frac{1 \text{ mol CH}_3\text{COO}^-}{1 \text{ mol NaOH}} \right) = 0.040 \text{ mol CH}_3\text{COO}^- \\ 0.600 \text{ mol} + 0.040 \text{ mol} & = \frac{0.640 \text{ mol}}{2.00 \text{ L}} = 0.320 \text{ M CH}_3\text{COO}^- \end{aligned}$$

$$\begin{aligned} \text{pH} &= 4.7447 + \log \frac{[0.320]}{[0.280]} \\ &= 4.7447 + 0.05799 = 4.803 \rightarrow 4.80 \end{aligned}$$

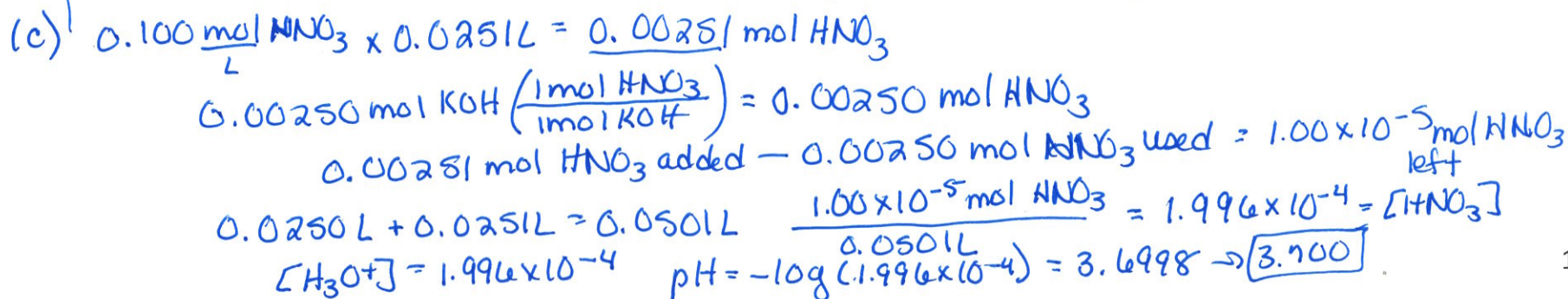
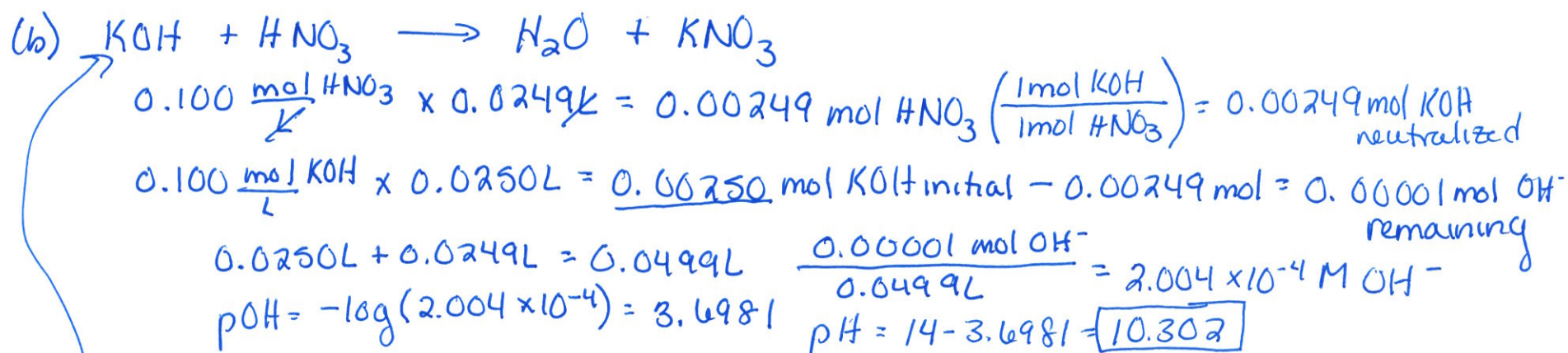
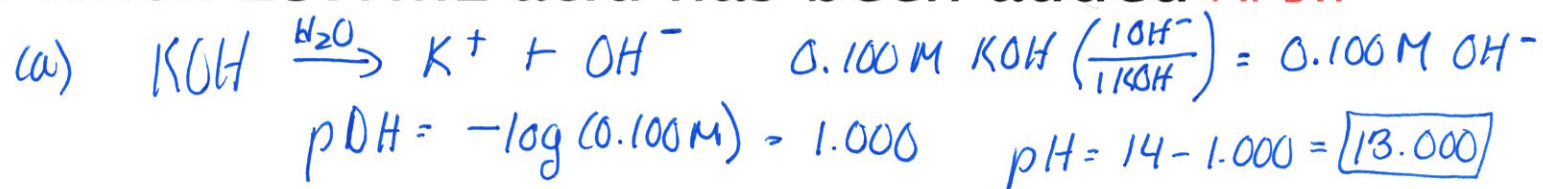
Strong Acid/Strong Base Calculations

In the titration of 25.0mL of 0.100M KOH with 0.100M HNO₃, determine the pH:

(a) At the start of the titration (no acid added) **A: 13.0**

(b) When 24.9mL acid has been added **A: 10.3**

(c) When 25.1mL acid has been added **A: 3.7**

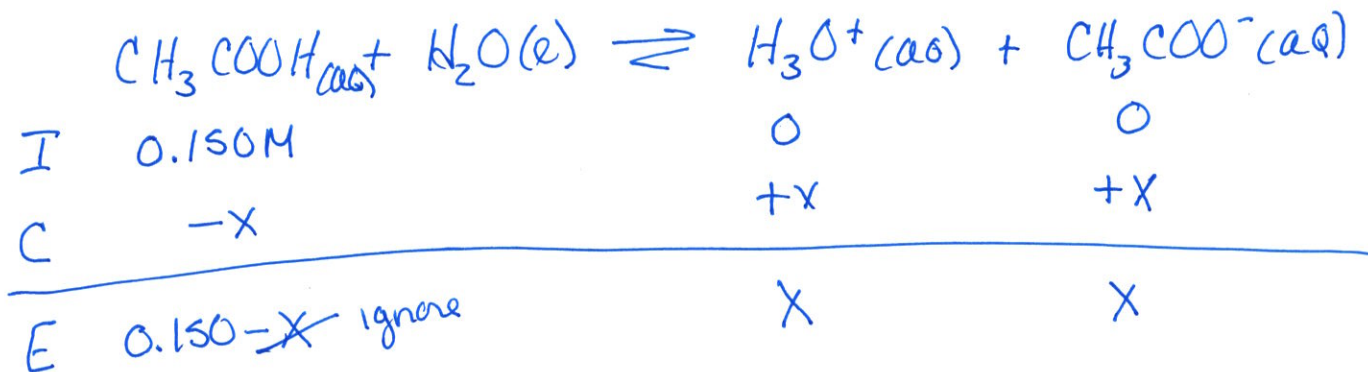


Weak Acid/Strong Base Calculations

35.0mL of 0.150M CH_3COOH ($K_a = 1.8 \times 10^{-5}$) was titrated with 0.150M NaOH. Determine the pH:

- a.) At the start of the titration A: 2.78
- b.) When 20.0mL of 0.150M NaOH has been added A: 4.87
- c.) At the equivalence point A: 8.81
- d.) When 50.0mL of 0.150M NaOH has been added A: 12.42

(a) No base added - only have acetic acid equilibrium



$$K_a = \frac{[x][x]}{0.150} = 1.8 \times 10^{-5}$$

$$x^2 = 2.7 \times 10^{-6}$$

$$x = 1.643 \times 10^{-3} = [\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log(1.643 \times 10^{-3})$$

$$= \boxed{2.784}$$

Weak Acid/Strong Base Calculations

35.0mL of 0.150M CH_3COOH ($K_a = 1.8 \times 10^{-5}$) was titrated with 0.150M NaOH. Determine the pH:

b.) When 20.0mL of 0.150M NaOH has been added **A:4.87**



$$\text{Initial moles acid: } \frac{0.150 \text{ mol}}{\text{L}} \times 0.0350 \text{ L} = 0.00525 \text{ mol CH}_3\text{COOH}$$

$$\text{Amount NaOH added: } \frac{0.150 \text{ mol}}{\text{L}} \times 0.0200 \text{ L} = 0.00300 \text{ mol NaOH}$$

$$0.00300 \text{ mol NaOH} \left(\frac{1 \text{ mol CH}_3\text{COOH}}{1 \text{ mol NaOH}} \right) = 0.00300 \text{ mol CH}_3\text{COOH neutralized}$$

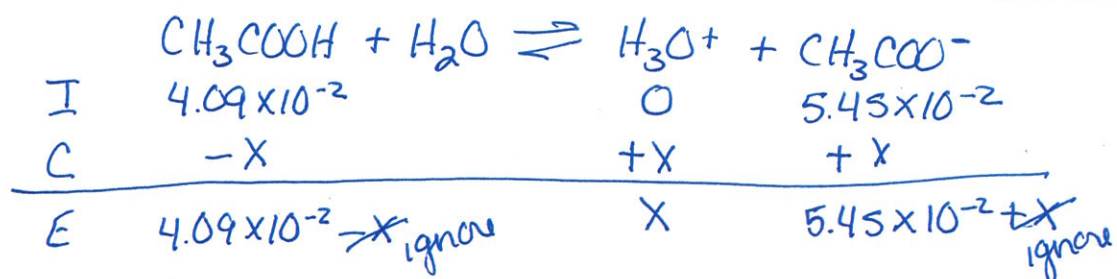
$$\text{CH}_3\text{COOH remaining: } 0.00525 \text{ mol initial} - 0.00300 \text{ mol neut} = 0.00225 \text{ mol CH}_3\text{COOH}$$

$$0.0350 \text{ L} + 0.0200 \text{ L} = 0.0550 \text{ L}$$

$$\frac{0.00225 \text{ mol}}{0.0550 \text{ L}} = 4.09 \times 10^{-2} \text{ M} = [\text{CH}_3\text{COOH}]$$

Amount CH_3COO^- :

$$0.00300 \text{ mol NaOH} \left(\frac{1 \text{ mol CH}_3\text{COO}^-}{1 \text{ mol NaOH}} \right) = \frac{0.00300 \text{ mol CH}_3\text{COO}^-}{0.0550 \text{ L}} = 5.45 \times 10^{-2} \text{ M} = [\text{CH}_3\text{COO}^-]$$



$$K_a = \frac{[X][5.45 \times 10^{-2}]}{[4.09 \times 10^{-2}]} = 1.8 \times 10^{-5}$$

$$[5.45 \times 10^{-2}][X] = 7.362 \times 10^{-7}$$

$$[X] = 1.35 \times 10^{-5} = [\text{H}_3\text{O}^+]$$

$$-\log(1.35 \times 10^{-5}) = \boxed{4.87}$$

Weak Acid/Strong Base Calculations

35.0mL of 0.150M CH_3COOH ($K_a = 1.8 \times 10^{-5}$) was titrated with 0.150M NaOH. Determine the pH:

c.) At the equivalence point **A: 8.81** $\text{CH}_3\text{COOH} + \text{NaOH} \rightarrow \text{CH}_3\text{COO}^- + \text{H}_2\text{O}$

All original acid & added strong base have been used up in neutralization

pH depends on CH_3COO^- equilibrium $\text{CH}_3\text{COO}^- + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{COOH} + \text{OH}^-$

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

Initially 0.00525 mol CH_3COOH $\left(\frac{1 \text{ mol } \text{CH}_3\text{COO}^-}{1 \text{ mol } \text{CH}_3\text{COOH}} \right) = 0.00525 \text{ mol } \text{CH}_3\text{COO}^-$

New volume: needed 0.00525 mol NaOH $\left(\frac{1 \text{ L}}{0.150 \text{ mol}} \right) = 0.035 \text{ L}$

$$0.0350 \text{ L} + 0.0350 \text{ L} = 0.070 \text{ L}$$

$$[\text{CH}_3\text{COO}^-] = \frac{0.00525 \text{ mol}}{0.070 \text{ L}} = 0.075 \text{ M}$$

$$K_b = \frac{x^2}{0.075} = 5.56 \times 10^{-10}$$

$$x^2 = 4.17 \times 10^{-11}$$

$$x = 6.458 \times 10^{-6} = [\text{OH}^-]$$

$$\text{pOH} = -\log(6.458 \times 10^{-6}) = 5.19$$

$$\text{pH} = 14 - 5.19 = \boxed{8.81}$$

	$\text{CH}_3\text{COO}^- + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{COOH} + \text{OH}^-$		
I	0.075	0	0
C	-x	+x	+x
E	0.075 - x ignore	x	x

Weak Acid/Strong Base Calculations

35.0mL of 0.150M CH_3COOH ($K_a = 1.8 \times 10^{-5}$) was titrated with 0.150M NaOH. Determine the pH:

d.) When 50.0mL of 0.150M NaOH has been added **A: 12.42**

pH depends on excess base $\text{CH}_3\text{COOH} + \text{NaOH} \rightarrow \text{H}_2\text{O} + \text{Na}^+ + \text{CH}_3\text{COO}^-$

Base added: $\frac{0.150 \text{ mol}}{L} \times 0.0500 L = 0.00750 \text{ mol NaOH}$

neutralized 0.00525 mol acid $\left(\frac{1 \text{ mol NaOH}}{1 \text{ mol acid}} \right) = 0.00525 \text{ mol NaOH used}$

0.00750 mol added - 0.00525 moles used = 0.00225 moles NaOH remaining

New volume: $0.0350 L + 0.0500 L = 0.085 L$

$[\text{NaOH}] = \frac{0.00225 \text{ mol}}{0.085 L} = 0.02647 M$ $\text{NaOH} \rightarrow \text{Na}^+ + \text{OH}^-$

$0.02647 M \text{ NaOH} = 0.02647 M \text{ OH}^-$

$\text{pOH} = -\log(0.02647) = 1.577$

$\text{pH} = 14 - 1.577 = \boxed{12.42}$