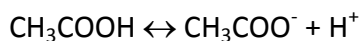


### Answer on Question #83024, Chemistry/ General Chemistry

Determine the pH change when 0.091 mol HCl is added to 1.00 L of a buffer solution that is 0.488 M in CH<sub>3</sub>COOH and 0.310 M in CH<sub>3</sub>COO<sup>-</sup>.

pH after addition – pH before addition = pH change =

#### Solution



To answer this question we should use Henderson-Hasselbalch equation:

$$\text{pH} = \text{p}K_a + \log\left(\frac{[\text{A}^-]}{[\text{HA}]}\right)$$

Where pH is the concentration of [H<sup>+</sup>]

pK<sub>a</sub> is the acid dissociation constant

$\frac{[\text{A}^-]}{[\text{HA}]}$  is the ratio of the concentrations of the conjugate base and starting acid.

Before addition:

$$\text{pH}(\text{before addition}) = \text{p}K_a + \log\left(\frac{[\text{A}^-]}{[\text{HA}]}\right) = \text{p}K_a + \log\left(\frac{0.310}{0.488}\right) = \text{p}K_a - 0.197$$

After addition:

Concentration of HCl added is 0.091 mol/1 L = 0.091 M

As we add a strong acid we make an assumption that all moles of conjugate base react with this acid to give a new concentration of [CH<sub>3</sub>COO<sup>-</sup>] = 0.310 - 0.091 = 0.219 M and additional concentration of CH<sub>3</sub>COOH is formed [CH<sub>3</sub>COOH] = 0.488 + 0.091 = 0.579 M

$$\text{pH}(\text{after addition}) = \text{p}K_a + \log\left(\frac{[\text{A}^-]}{[\text{HA}]}\right) = \text{p}K_a + \log\left(\frac{0.310 - 0.091}{0.488 + 0.091}\right) = \text{p}K_a - 0.422$$

pH change = pH( after addition) – pH (before addition) = pK<sub>a</sub> – 0.422 – (pK<sub>a</sub> – 0.197) = pK<sub>a</sub> – 0.422 – pK<sub>a</sub> + 0.197 = –0.225 pH units

**Answer: pH change = -0.225 PH units**

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