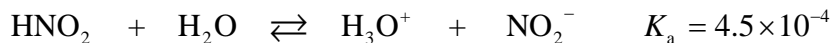


## Equilibrium V: Acid-Base Buffers and Titrimetry

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1. Calculate the pH of the solution which is prepared to be 0.10 M in nitrous acid and 0.15 M sodium nitrite.



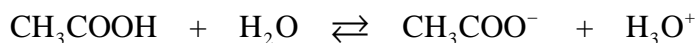
Recall that for a typical acid-base buffer of relatively high concentration, the equilibrium concentrations of the acid and conjugate base differ very little from their analytical concentrations

$$K_a = 4.5 \times 10^{-4} = \frac{[\text{H}_3\text{O}^+](0.15 \text{ M})}{(0.10 \text{ M})}$$

$$[\text{H}_3\text{O}^+] = 3.0 \times 10^{-4} \text{ M} \quad \text{pH} = 3.5$$

2. What is the calculated pH of the buffer solution formed by dissolving 4.2 g of acetic acid and 9.0 g of sodium acetate in water and diluting to 500.0 mL?

$$C_{\text{HOAc}} = \frac{4.2 \text{ g CH}_3\text{COOH} / \frac{60.05 \text{ g}}{\text{mol}}}{0.5000 \text{ L}} = 0.140 \text{ M} \quad C_{\text{OAc}^-} = \frac{9.0 \text{ g NaCH}_3\text{COO} / \frac{82.04 \text{ g}}{\text{mol}}}{0.5000 \text{ L}} = 0.22 \text{ M}$$



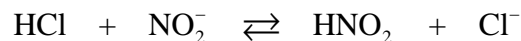
$$K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{COOH}]} \cong \frac{C_{\text{CH}_3\text{COO}^-}}{C_{\text{CH}_3\text{COOH}}} [\text{H}_3\text{O}^+]$$

$$[\text{H}_3\text{O}^+] = \frac{C_{\text{CH}_3\text{COOH}}}{C_{\text{CH}_3\text{COO}^-}} (1.75 \times 10^{-5}) = \frac{0.14}{0.22} (1.75 \times 10^{-5}) = 1.11 \times 10^{-5} \text{ M}$$

$$\text{pH} = 4.95$$

3. A 10.0 mL quantity of 0.1 M HCl is mixed with 95 mL of 0.15 M sodium nitrite. What is the pH of the new solution?

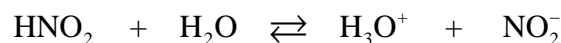
$$n_{\text{NO}_2^-} = 95 \text{ mL} \times 0.15 \text{ M} = 14.25 \text{ mmol NO}_2^- \quad n_{\text{HCl}} = 10.0 \text{ mL} \times 0.1 \text{ M} = 1.0 \text{ mmol HCl}$$



$$n_{\text{HNO}_2, \text{formed}} = 1.0 \text{ mmol HCl} \times \frac{1 \text{ mmol HNO}_2}{1 \text{ mmol HCl}} = 1.0 \text{ mmol HNO}_2 \quad C_{\text{HNO}_2} = \frac{1.0 \text{ mmol HNO}_2}{105 \text{ mL}} = 9.52 \times 10^{-3} \text{ M}$$

$$n_{\text{NO}_2^-, \text{remaining}} = 14.25 \text{ mmol NO}_2^- - 1.0 \text{ mmol HCl} \times \frac{1 \text{ mmol NO}_2^-}{1 \text{ mmol HCl}} = 13.25 \text{ mmol NO}_2^-$$

$$C_{\text{NO}_2^-} = \frac{13.25 \text{ mmol NO}_2^-}{105 \text{ mL}} = 0.126 \text{ M}$$

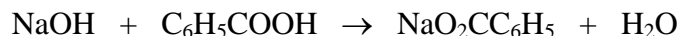


$$4.5 \times 10^{-4} = \frac{[\text{H}_3\text{O}^+][\text{NO}_2^-]}{[\text{HNO}_2]}$$

$$[\text{H}_3\text{O}^+] = \frac{9.52 \times 10^{-3} \text{ M}}{0.126 \text{ M}} (4.5 \times 10^{-4}) = 3.39 \times 10^{-5} \text{ M} \quad \text{pH} = 4.47$$

4. What is the calculated pH of the solution formed by mixing 10.0 mL of 0.15 M NaOH with 85 mL of 0.10 M benzoic acid?

This problem starts as simple stoichiometry problem and ends as a buffer problem. Start with the acid-base neutralization:



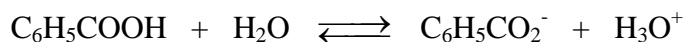
$$n_{\text{C}_6\text{H}_5\text{COOH}}^{\text{initial}} = 0.085 \text{ L} \times 0.10 \text{ M C}_6\text{H}_5\text{COOH} = 0.0085 \text{ mol C}_6\text{H}_5\text{COOH}$$

$$n_{\text{NaOH}}^{\text{added}} = 0.0100 \text{ L} \times 0.15 \text{ M NaOH} = 0.0015 \text{ mol NaOH}$$

$$n_{\text{C}_6\text{H}_5\text{COOH}}^{\text{remaining}} = 0.0085 \text{ mol C}_6\text{H}_5\text{COOH} - 0.0015 \text{ mol NaOH} \times \left( \frac{1 \text{ mol C}_6\text{H}_5\text{COOH}}{1 \text{ mol NaOH}} \right) = 0.0070 \text{ mol C}_6\text{H}_5\text{COOH}$$

$$n_{\text{C}_6\text{H}_5\text{CO}_2^-}^{\text{made}} = 0.0015 \text{ mol NaOH} \times \left( \frac{1 \text{ mol C}_6\text{H}_5\text{CO}_2^-}{1 \text{ mol NaOH}} \right) = 0.0015 \text{ mol C}_6\text{H}_5\text{CO}_2^-$$

Now, work out the equilibrium. This is a buffer solution so dividing each of the quantities by volume is convenient but not necessary.



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{C}_6\text{H}_5\text{CO}_2^-]}{[\text{C}_6\text{H}_5\text{COOH}]} = \frac{[\text{H}_3\text{O}^+](0.0015 \text{ mol})}{0.0070 \text{ mol}} = 6.3 \times 10^{-5}$$

$$[\text{H}_3\text{O}^+] = 2.94 \times 10^{-4} \text{ M} \quad \text{pH} = 3.53$$