CHEM1102 Worksheet 9: Weak Acids and Titrations

Model 1: Addition of a Strong Acid to a Weak Base

In Worksheet 8, you used an "ICE" approach to work out the pH of a solution containing a weak acid. For example, you worked out that a 0.500 M solution of $CH_3COOH(aq)$ has a pH of 2.531.

If a strong base, such as NaOH, is added to this solution, it will react with the weak acid.

$$CH_3COOH(aq) + OH^-(aq) \rightarrow CH_3COO^-(aq) + H_2O(aq)$$

As long as the amount of $OH^-(aq)$ added is *less* than the amount of $CH_3COOH(aq)$ present, the solution will contain both $CH_3COO^-(aq)$ and left over $CH_3COOH(aq)$. A solution like this containing both a weak acid and its conjugate base is a buffer and its pH is given by the *Henderson-Hasselbalch* equation:

$$\text{pH} = \text{p}K_a + \log\frac{[\text{base}]}{[\text{acid}]} = \text{p}K_a + \log\frac{[CH_3COO^-(aq)]}{[CH_3COOH(aq)]}$$

Critical thinking questions

1. If 0.100 mol of NaOH(s) is added to a 1.00 L solution of 0.500 M CH_3COOH, it will react to form a solution which is 0.100 M in $CH_3COO^-(aq)$ and 0.400 M $CH_3COOH(aq)$. What is the pH of this solution? (pK_a (CH_3COOH) = 4.76).

2. Complete the table below showing the concentrations of $CH_3COOH(aq)$ and $CH_3COO^-(aq)$ and the pH of the solution as more NaOH(s) is added to this solution.

Amount of NaOH(s) added (mol)	0.000	0.100	0.200	0.300	0.400	0.500
$[CH_3COOH(aq)]$ (M)	0.500	0.400				
$[CH_3COO^-(aq)]$ (M)	0.000	0.100				
pH	2.531					

3. To react completely with the original CH_3COOH, 0.500 mol of NaOH must be added. What is the pH of the solution when exactly *half* this amount is added?

4. How can you obtain the value for pK_a for an acid?

Model 2: Neutralizing a Weak Acid

Model 1 describes the pH changes as a strong base is added to a solution containing a weak acid. The strong base reacts with the weak acid leading to a solution containing the conjugate base of the weak acid and any left over acid. The *equivalence* point occurs when enough base has been added so that there is no acid left.

At this point, the solution contains the conjugate base and essentially none of the original acid.

The conjugate base will then set up its own equilibrium:

$$CH_3COO^-(aq) + H_2O(l) \rightleftharpoons CH_3COOH(aq) + OH^-(aq) \qquad K_b = \frac{[CH_3COOH(aq)][OH^-(aq)]}{[CH_3COO^-(aq)]}$$

From the chemical equation:

$$[CH_3COOH(aq)]_{equilirium} = [OH^-(aq)]_{equilirium}$$

As hardly any base reacts, $[CH_3COO^-(aq)]_{equilibrium} \approx [CH_3COO^-(aq)]_{initial}$ and so:

$$K_b = \frac{[OH^-(aq)]^2}{[CH_3COO^-(aq)]_{initial}} \qquad \text{and} \qquad [OH^-(aq)] = \sqrt{K_b \times [CH_3COO^-(aq)]_{initial}}$$

After working out $[OH^-(aq)]$, the pOH can be calculated:

$$pOH = -\log[OH^-(aq)]$$

Finally, the pH can then be calculated using pH = 14.00 – pOH.

Critical thinking questions

1. To react completely with the original CH_3COOH in Q5, 0.500 mol of NaOH must be added. What will be $[CH_3COO^-(aq)]$ when this occurs?

2. Calculate the pH of the solution the reaction produced in Q1.
(*Hint*: remember that $pK_a + pK_b = 14.00$ or $K_a \times K_b = 10^{-14.00}$)

3. Correct your entry in the final column of the table in Model 1 if required!

4. What is the pH at the equivalence point of the titration of a *strong* acid with a strong base?

5. Is the pH at the equivalence point of the titration of a *weak* acid with a strong base less than or higher than 7?