

Chemistry 146 Lecture Problems

Generic Buffer

Let's take a look at how this buffer system is useful.

Given a weak acid system where:



And the equilibrium starting conditions for 1 liter of solution:

$$C_{\text{HA}} := 1 \cdot \frac{\text{mole}}{\text{liter}} \quad C_{\text{A}^-} := 1 \cdot \frac{\text{mole}}{\text{liter}} \quad C_{\text{H}_3\text{O}^+} := 1 \cdot \frac{\text{mole}}{\text{liter}} \quad M := \frac{\text{mole}}{\text{liter}}$$

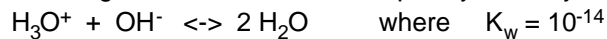
This gives an equilibrium constant (K_a):

$$K_a := \frac{C_{\text{A}^-} \cdot C_{\text{H}_3\text{O}^+}}{C_{\text{HA}}}$$

$$K_a = 1 \cdot M$$

Now, let's disrupt the equilibrium by adding 0.5 moles NaOH

OH^- is a strong base so it will react completely with any best proton donor.



Almost all of the OH^- is converted to H_2O .

As a result of this reaction the new, non-equilibrium concentrations are approximately:

$$C_{\text{H}_3\text{O}^+} := C_{\text{H}_3\text{O}^+} - 0.5 \cdot M \quad C_{\text{HA}} = 1 \cdot M \quad C_{\text{A}^-} = 1 \cdot M$$

$$C_{\text{H}_3\text{O}^+} = 0.5 \cdot M \quad C_{\text{H}_3\text{O}^+} = 0.5 \cdot M$$

Assuming X moles shift, the new equilibrium may be described as:

$$C_{\text{H}_3\text{O}^+} = C_{\text{H}_3\text{O}^+} + X \quad C_{\text{A}^-} = C_{\text{A}^-} + X \quad C_{\text{HA}} = C_{\text{HA}} - X$$

Then K_a is:

$$K_a = \frac{(C_{\text{H}_3\text{O}^+} + X) \cdot (C_{\text{A}^-} + X)}{(C_{\text{HA}} - X)}$$

Substituting in for the known values results in the expression:

$$1 = \frac{(0.5 + X) \cdot (1.0 + X)}{(1.0 - X)}$$

expands to

$$1 = \frac{(.5 + 1.5 \cdot X + X^2)}{(1.0 - X)}$$

rearranges to

$$-.5 + 2.5 \cdot X + X^2 = 0$$

has solution(s)

$$X = \begin{bmatrix} -2.686140661634507165 \\ .186140661634507165 \end{bmatrix}$$

Picking the physically meaningful root gives:

$$X := 0.18614 \cdot M$$

Which gives new equilibrium concentrations as:

$$C_{H_3O^+} := C_{H_3O^+} + X \quad C_A := C_A + X \quad C_{HA} := C_{HA} - X$$

$$C_{H_3O^+} = 0.68614 \cdot M \quad C_A = 1.18614 \cdot M \quad C_{HA} = 0.81386 \cdot M$$

Even though 0.5 mole of NaOH was added, the [H₃O⁺], changed by a much smaller amount. If you assume the volume of the solution was constant at 1 L:

$$pH_{initial} := -\log(1)$$

$$pH_{initial} = 0$$

$$pH_{final} := -\log(0.686)$$

$$pH_{final} = 0.16368$$

Compare this to the pH change if you add 0.5 mole NaOH to 1 L of water

$$pH_{water} := -\log(10^{-7})$$

$$pH_{water} = 7$$

$$pOH_{base} := -\log(0.5)$$

$$pOH_{base} = 0.30103$$

$$pH_{base} := 14 - pOH_{base}$$

$$pH_{base} = 13.69897$$