

The Buffering Shortcut: Henderson-Hasselbach Equation

Although RICE table calculations can always be used to determine $[H_3O^+]$ and thus pH, there is a more efficient way. This process works because buffers are always weak acids or bases and their conjugate, so the "x is negligible" approximation will always be accurate: we can consider the equilibrium concentrations of the weak acid, HA, and its conjugate base, A^- to be essentially equal to their initial concentrations.

I'm not lazy



Just buffering

We can derive an equation that relates the pH of a buffer solution to the initial concentration of the buffer components by rearranging the acid ionization constant expression.

not on F.C.

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

$$[H_3O^+] = K_a \frac{[HA]}{[A^-]}$$

or

$$[H_3O^+] = K_a \frac{[Acid]}{[Base]}$$

This equation can also be written in the following format:

on F.C.

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

or

$$pH = pK_a + \log\left(\frac{[Base]}{[Acid]}\right)$$

K_a – is the acid dissociation constant

[Acid] – is the concentration of the substance behaving as the **weak acid** in the buffered solution

- Either the weak acid or the salt of the conjugate base

[Base] – is the concentration of the substance behaving as the **weak base** in the buffered solution

- Either the weak base or the salt of the conjugate acid

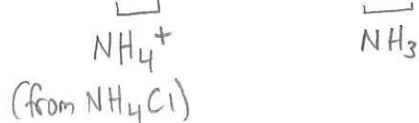
How To Solve Buffer Problems

- If the buffer is a weak acid, HA, and its conjugate salt, A^- , then HA is the acid and A^- is the base.
- If the buffer is a weak base, B, and its conjugate salt, BH^+ , then B is the base and BH^+ is the acid.
 - Even if the buffer uses a weak base, you still use K_a : you must calculate it first using the K_b provided.

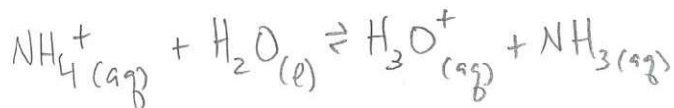
Let's Practice!

1. A buffered solution contains 0.250 M NH_3 ($K_b = 1.80 \times 10^{-5}$), and 0.400 M NH_4Cl .

a. What species is the acid and which is the base?



b. Write an equilibrium reaction to describe this buffer.



c. Calculate the pH.

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{NH}_3]}{[\text{NH}_4^+]} = \frac{K_w}{K_b}$$

$$\frac{1 \times 10^{-14}}{1.80 \times 10^{-5}} = \frac{[\text{H}_3\text{O}^+](0.250)}{0.400}$$

$$[\text{H}_3\text{O}^+] = \frac{(1 \times 10^{-14})(0.400)}{(1.80 \times 10^{-5})(0.250)} = 8.89 \times 10^{-10}$$

$$\Rightarrow \text{pH} = -\log(8.89 \times 10^{-10}) = \boxed{9.051}$$

OR

$$\begin{aligned} \text{pH} &= \text{p}K_a + \log \frac{[\text{A}^-]}{[\text{HA}]} \\ &= -\log\left(\frac{1 \times 10^{-14}}{1.80 \times 10^{-5}}\right) + \log\left(\frac{0.250}{0.400}\right) \\ &= 9.255 - 0.204 = \boxed{9.051} \end{aligned}$$

HCOO^- CB

2. A student makes a 0.500 L of a buffer solution of pH 3.89 by combining formic acid, HCOOH , and sodium formate, NaHCOO . The K_a of formic acid is 1.8×10^{-4} .

a. Write an equilibrium reaction to describe this buffer.



b. What ratio of formic acid to sodium formate would be required to make this buffer?

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

$$3.89 = -\log(1.8 \times 10^{-4}) + \log \frac{[\text{HCOO}^-]}{[\text{HCOOH}]}$$

$$\Rightarrow \log \frac{[\text{HCOO}^-]}{[\text{HCOOH}]} = 3.89 + 3.74 = 0.15 \Rightarrow \frac{[\text{HCOO}^-]}{[\text{HCOOH}]} = 10^{0.15} = 1.4$$

$$\Rightarrow \boxed{1 \text{ HCOOH} : 1.4 \text{ NaHCOO}}$$