Exercise 3: A weak acid

Given the reaction of an acid HB ⇌ H⁺ + B⁻, you want to determine the equilibrium constant (actually, the dissociation constant of an acid, K_a). You will do this by observing the concentration of various species in a test tube.

1. Write the equilibrium constant expression for HB. Use K_a as the equilibrium constant. Use all or some of the following: [H⁺]_equilibrium, [H⁺]_initial, [B⁻]_equilibrium, [B⁻]_initial, [HB]_equilibrium, [HB]_initial.

2. The HB flask has a concentration of 2.5 × 10⁻⁴ M written on the side. You pipet 10 mL into a test tube, then add 40 mL of 0.010 M sodium chloride. What is the concentration of the diluted HB? Hint: Is the concentration of the sodium chloride relevant?

3. Call this [HB]_initial. How come the original 2.5 × 10⁻⁴ M isn’t called the “initial concentration”?

4. Suppose you measure the pH of the diluted solution (using a pH meter) as 7.66. What is the concentration of the hydrogen ion in the diluted solution? Is this the initial or the equilibrium concentration?

5. What must be the concentration of the B⁻ ion? Is this the initial or the equilibrium concentration? Why must it be this concentration?
6. What is the equilibrium concentration of HB? How do you calculate it from the answer to question 3?

7. Calculate $K_a$.

8. Suppose, instead of 0.010 M NaCl, to the original 10 mL of HB, 40 mL of 0.010 M HA (where HA is some generic weak acid) is added. Will the measured pH reflect only the protons dissociated from HB? Why or why not?

9. Given the scenario above, a pH of 6.67 is measured. If you did not know $K_a$ from question 7, could $K_a$ be calculated from the information given? If so, calculate it. If not, what other piece of information is needed?

10. Given the two scenarios (questions 7 and 9), should the HB $K_a$’s be different, since one scenario has HA and the other one doesn’t? Explain your answer.