Which of these statements is **false**, given that pK_b of aniline ($C_6H_5NH_2$) is 9.40?

a. The reaction $C_6H_5NH_2 + H_2O \rightleftharpoons C_6H_5NH_3^+ + OH^-$ has an equilibrium constant of 4.0×10^{-10}.

b. Aniline is a weak base.

c. The K_a of aniline is 2.5×10^{-5}.

d. The reaction $C_6H_5NH_3^+ + H_2O \rightleftharpoons C_6H_5NH_2 + H_3O^+$ has an equilibrium constant of 2.5×10^{-5}.

e. Aniline will not react OH^- to an appreciable degree.

f. The reaction $C_6H_5NH_2 + H_3O^+ \rightleftharpoons C_6H_5NH_3^+ + H_2O$ has a very large equilibrium constant.

Common-Ion Effect: Solutions of >1 compound

17.1

Solutions need not be only formed from one compound at a time.

The Common-Ion Effect describes the effect on an equilibrium by a second substance that furnishes ions that can participate in that equilibrium.

The added ions are said to be *common* to the equilibrium.

e.g. a solution of acetic acid and sodium acetate

Calcium carbonate and carbon dioxide (carbonic acid) in ocean water

Scenario 1: Weak Acid + Strong Base (or vice versa)

If a weak acid (or base) is combined with a strong base (or acid), we consider the reaction in two steps:

1. React the strong acid/base with the weak component in solution, if possible. Use as much of the strong component as possible (remember, equilibrium lies *away* from the strong acid/base in a reaction)

2. Allow the mixture to reach equilibrium (ICE table).

Example 2: What is the pH of a solution made by combining 1.00 L each of 0.200 M sodium fluoride ($K_{a,\ HF}$ = 6.6x10^{-4}) and 0.0100 M hydrochloric acid?

Scenario 2: Weak + Weak

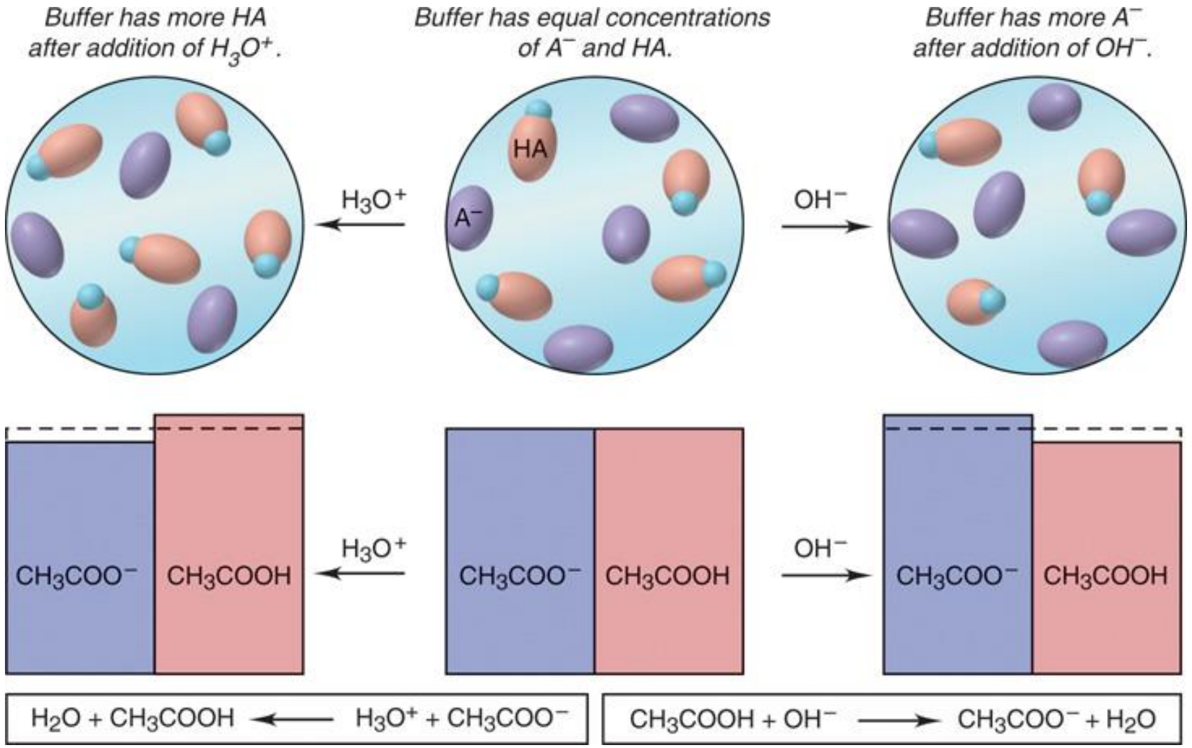
A solution that contains both a weak acid (or base) and the salt of its' conjugate is dealt with by an equilibrium calculation:

e.g.: What is the pH of a solution made by combining 100 mL of a 0.1 M sodium fluoride solution with 100 mL of a 0.01 M hydrofluoric acid solution?

Buffers

A solution that contains both components of a conjugate acid-base pair is called a **buffer**.

Because these solutions contain both a weak acid and a weak base, the pH will not change significantly on addition of a small amount of strong acid or base.



In order for a system to act as a buffer, the equilibrium concentration of the weak acid/base components must be approximately the same as the initial concentrations of these components. (i.e. K_a is small compared to the concentrations)

Therefore, if we know that a solution is behaving as a buffer, we can say (assuming a HA/A$^-$ generic system):

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

Taking the –log of both sides:

This is the **Hendersen-Hasselbach (HH) Equation**.

Revisit the last calculation: What is the pH of a solution made by combining 100 mL of a 0.1 M sodium fluoride solution with 100 mL of a 0.01 M hydrofluoric acid solution?

Preparing Buffer Solutions

Buffers can be prepared in several ways:

- Calculating the [acid] and [base] required for the desired pH and adding HA and A^- salt (or the base equivalents).
  - This method works but is tedious in practice.

- Starting with a solution of the weak acid and titrating in strong base until the desired pH is reached
- Starting with a solution of the weak base and titrating in strong acid until the desired pH is reached
  - These methods are easier (and you can calculate the amount of strong acid/base required to speed the process)

Which of the following, when mixed, could produce a “good” buffer solution (with approximately equal amounts of conjugate acid and base)?

1. $HCl/NaCl$
2. $HC_2H_3O_2/NH_3$
3. NaH_2PO_4/Na_2HPO_4
4. $HNO_3/Ca(OH)_2$
5. $KNO_3/$ $NaOH$

Fig. 16.9

Buffer Capacity

The ability of a buffer to resist pH change is the *buffer capacity*. Typically, we define it as **the amount (in mmol) of strong acid or base required to change the pH by 1 unit**.