The following equilibrium equation for a typical weak base, B, is used to derive the generalized expression for K_b, the base dissociation constant.

$$B(aq) + H_2O(l) \rightleftharpoons BH^+(aq) + OH^-(aq)$$

$$K_b = \frac{[BH^+][OH^-]}{[B]}$$

The hydrolysis reaction between water and the cation, BH^+, produced by the dissociation of the weak base, B, is represented by the general equilibrium equation that follows.

$$BH^+(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + B(aq)$$

In the forward reaction, the cation BH^+ donates a proton to the water molecule to form the hydronium ion and the weak base, B. The extent of H_3O^+ ion formation and the position of the equilibrium depend on the relative strength of the cation, BH^+. The lower the K_b value of B, the stronger the donation of protons that BH^+ will have compared with H_3O^+, and the greater the production of H_3O^+ ions will be. Therefore, the weaker the base, the stronger its conjugate acid will be.

Ammonium chloride, NH_4Cl, dissociates in water to produce NH_4^+ ions, Cl^- ions, and an acidic solution. Chloride ions are the conjugate base of a strong acid, HCl, so they show no noticeable tendency to hydrolyze in aqueous solution. Ammonium ions, however, are the conjugate acid of a weak base, NH_3. Ammonium ions donate protons to water molecules. Equilibrium is established with an increased $[H_3O^+]$, so the pH is *lower* than 7.

Neutralization Curve for 100 mL of 0.100 M CH_3COOH Titrated with 0.100 M NaOH

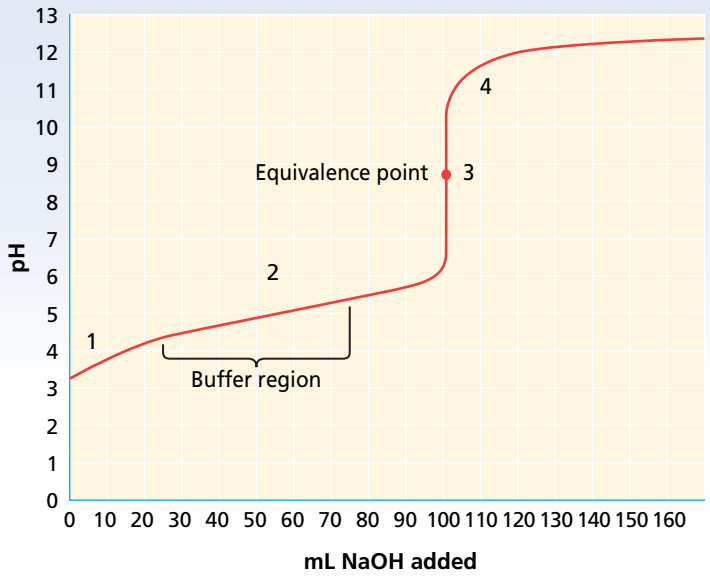

FIGURE 11 At point *1* on the titration curve, only acetic acid is present. The pH depends on the weak acid alone. At *2* there is a mixture of CH_3COOH and CH_3COO^-. Adding NaOH changes the pH slowly. At point *3* all acid has been converted to CH_3COO^-. This hydrolyzes to produce a slightly basic solution. At *4* the pH is due to the excess OH^- that has been added.