The hydronium ion concentration and the pH of the solution remain practically unchanged.

Suppose a small amount of a base is added to the original solution. The OH^- ions of the base react with and remove hydronium ions to form nonionized water molecules. Acetic acid molecules then ionize and mostly replace the hydronium ions neutralized by the added OH^- ions.

$$CH_3COOH(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + CH_3COO^-(aq)$$

The pH of the solution again remains practically unchanged.

A solution of a weak base containing a salt of the base behaves in a similar manner. The hydroxide ion concentration and the pH of the solution remain essentially constant with small additions of acids or bases. Suppose a base is added to an aqueous solution of ammonia that also contains ammonium chloride. Ammonium ions donate a proton to the added hydroxide ions to form nonionized water molecules.

$$NH_4^+(aq) + OH^-(aq) \longrightarrow NH_3(aq) + H_2O(l)$$

If a small amount of an acid is added to the solution instead, hydroxide ions from the solution accept protons from the added hydronium ions to form nonionized water molecules. Ammonia molecules in the solution then ionize and mostly replace the hydroxide ions neutralized by added H_3O^+.

$$NH_3(aq) + H_2O(l) \longrightarrow NH_4^+(aq) + OH^-(aq)$$

Buffer action has many important applications in chemistry and physiology. Human blood is naturally buffered to maintain a pH of between 7.3 and 7.5. This is essential because large changes in pH would lead to serious disturbances of normal body functions. **Figure 9** shows an example of one of the many medicines buffered to prevent large and potentially damaging changes in pH.

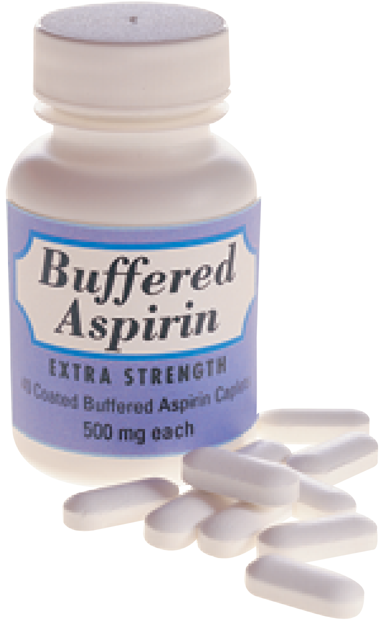

FIGURE 9 Many consumer products are buffered to protect the body from potentially harmful pH changes.

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Topic: Buffers

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Ionization Constant of Water

Recall from Chapter 15 that the self-ionization of water is an equilibrium reaction.

$$H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$$

Equilibrium is established with a very low concentration of H_3O^+ and OH^- ions. The expression for the equilibrium constant, $K_w = [H_3O^+][OH^-]$, is derived from the balanced chemical equation. The numerical value of K_w, obtained experimentally, is 1.0×10^{-14} at 25°C.