Charges are essential for the stability and function of many proteins. To understand the effect of the environment, particularly its pH on stability and function, a few concepts are needed.

### 19.1 Acids and Bases

An *acid* is a proton donor, a *base* a proton acceptor; they always occur in conjugated pairs [1]:

\[
\text{acid} \leftrightarrow \text{base} + \text{H}^+.
\]  

(19.1)

Examples are:

- \(\text{NH}_4^+ \leftrightarrow \text{NH}_3 + \text{H}^+\)
- \(\text{H}_2\text{SO}_4 \leftrightarrow \text{HSO}_4^- + \text{H}^+\)
- \(\text{H}_3\text{O}^+ \leftrightarrow \text{H}_2\text{O} + \text{H}^+\)
- \(\text{H}_2\text{O} \leftrightarrow \text{OH}^- + \text{H}^+\)

Some compounds can act as either base or acid. Physicists often believe that the \(\text{H}^+\) appears as a proton. In reality, however, protons have the tendency to sneak into the electron shell of neighboring molecules. In an aqueous solution, protons appear in the form of hydronium ions, \(\text{H}_3\text{O}^+\). The dissociation of an acid, Eq. (19.1), thus occurs as shown in Fig. 19.1.

The acid transfers its proton directly onto a neighboring water molecule. The correct way to describe the acid-base equilibrium hence is

\[
\text{AH} + \text{H}_2\text{O} \leftrightarrow \text{A}^- + \text{H}_3\text{O}^+.
\]  

(19.2)

For “simplicity” (and to confuse physicists), the reaction is usually written in the form of Eq. (19.1).
19.2 The pH Scale

In water, Eq. (19.2) becomes

\[ \text{H}_2\text{O} + \text{H}_2\text{O} \leftrightarrow \text{HO}^- + \text{H}_3\text{O}^+ \]  \hspace{1cm} (19.3)

or

\[ \text{H}_2\text{O} \leftrightarrow \text{OH}^- + \text{H}^+. \]

The equilibrium coefficient of this reaction is given by

\[ K = [\text{H}^+][\text{OH}^-]/[\text{H}_2\text{O}]. \]  \hspace{1cm} (19.4)

Usually, the concentration \([\text{H}_2\text{O}]\) of water is very large, about 55 M, and the concentrations of the \(\text{H}^+\) and \(\text{OH}^-\) ions very small, \([\text{H}_2\text{O}]\) is essentially constant during the reaction (19.4). A new coefficient \(K_{eq}\) is consequently introduced through the definition

\[ [\text{H}^+][\text{OH}^-] = K_{eq}. \]  \hspace{1cm} (19.5)

The coefficient \(K_{eq}\) is called the ion product of water. Its value at 25°C is \(1.0 \times 10^{-14}\) M². In an acid solution, the concentration \([\text{H}^+]\) is relatively high; in a basic solution, \([\text{OH}^-]\) is relatively high.

Equation (19.5) is the basis for the pH scale. Since \(K_{eq}\) can change over many orders of magnitude, it is more convenient to use logarithms than powers. The term pH is defined as [2]

\[ \text{pH} = -\log[\text{H}^+]. \]

In pure water, Eq. (19.4) tells us that \([\text{H}^-] = [\text{OH}^+] = 10^{-7}\) M and hence pH (pure water) = 7. The pH of some solutions is given in Table 19.1.