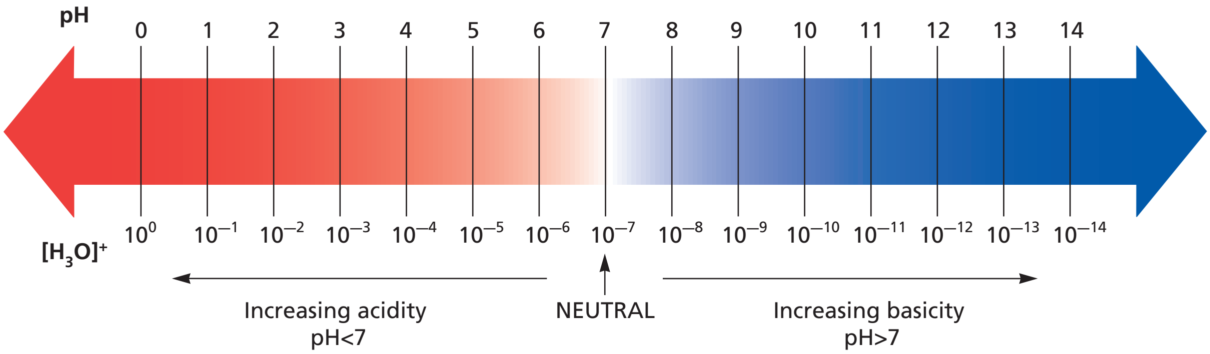

FIGURE 3 As the concentration of hydronium ions increases, the solution becomes more acidic and the pH decreases. As the concentration of hydronium ions decreases, the solution becomes more basic and the pH increases.

The pH Scale

Expressing acidity or basicity in terms of the concentration of H_3O^+ or OH^- can be cumbersome because the values tend to be very small. A more convenient quantity, called pH, also indicates the hydronium ion concentration of a solution. The letters *pH* stand for the French words *pouvoir hydrogène,* meaning "hydrogen power." *The* **pH** *of a solution is defined as the negative of the common logarithm of the hydronium ion concentration,* $[H_3O^+]$. The pH is expressed by the following equation.

$$\text{pH} = -\log [H_3O^+]$$

The common logarithm of a number is the power to which 10 must be raised to equal the number. A neutral solution at 25°C has a $[H_3O^+]$ of 1×10^{-7} M. The logarithm of 1×10^{-7} is −7.0. The pH is determined as follows.

$$\text{pH} = -\log [H_3O^+] = -\log (1 \times 10^{-7}) = -(-7.0) = 7.0$$

The relationship between pH and $[H_3O^+]$ is shown on the scale in **Figure 3.**

Likewise, *the* **pOH** *of a solution is defined as the negative of the common logarithm of the hydroxide ion concentration,* $[OH^-]$.

$$\text{pOH} = -\log[OH^-]$$

A neutral solution at 25°C has a $[OH^-]$ of 1×10^{-7} M. Therefore, the pOH is 7.0.

Remember that the values of $[H_3O^+]$ and $[OH^-]$ are related by K_w. The negative logarithm of K_w at 25°C, 1×10^{-14}, is 14.0. You may have noticed that the sum of the pH and the pOH of a neutral solution at 25°C is also equal to 14.0. The following relationship is true at 25°C.

$$\text{pH} + \text{pOH} = 14.0$$

At 25°C the range of pH values of aqueous solutions generally falls between 0 and 14, as shown in **Table 3.**