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17

ACIDS AND BASES

In previous chapters we have referred from time to time to water solutions of acids and bases. Such solutions are among the most useful of laboratory reagents. They are also common in the home. Vinegar, orange juice, and battery fluid are, to varying degrees, acidic. Oven and drain cleaners, whitewash (like Tom Sawyer used), and antacid tablets are basic.

This chapter examines the properties of acidic and basic water solutions. We start with working definitions of the terms acid and base. To begin, we take an *acid* to be a substance which, when added to water, produces *hydrogen ions* (**H^+**). *Bases* produce *hydroxide ions* (**OH^-**) in water solution.

All acidic water solutions have certain properties in common. They evolve hydrogen gas, H_2, with zinc or magnesium. They react with compounds containing the CO_3^{2-} ion to form CO_2 (Fig. 17.1). They affect the color of certain organic dyes, called indicators. For example, litmus turns red in acidic solution. These properties of acidic solutions are due to the presence of H^+ ions.

Water solutions of bases also have identifying properties. They feel slippery and

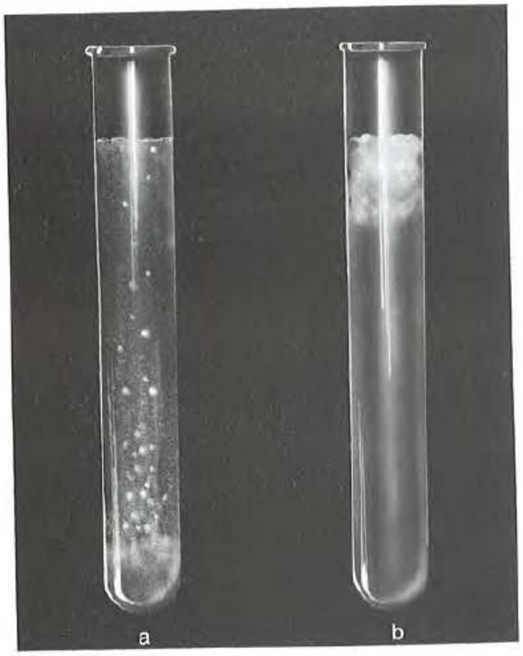

Figure 17.1 The H^+ ions present in a water solution of an acid react with calcium carbonate to form carbon dioxide: $CaCO_3(s) + 2\,H^+(aq) \rightarrow Ca^{2+}(aq) + CO_2(g) + H_2O$ (tube a). The OH^- ions present in a solution of a base precipitate $Mg(OH)_2$ from a solution containing Mg^{2+} ions: $Mg^{2+}(aq) + 2\,OH^-(aq) \rightarrow Mg(OH)_2(s)$ (tube b).

turn litmus blue. In the presence of Mg^{2+} ions, they precipitate magnesium hydroxide, $Mg(OH)_2$. The characteristic properties of basic solutions are caused by OH^- ions.

17.1 WATER DISSOCIATION; ACIDIC, NEUTRAL, AND BASIC SOLUTIONS

The acidic and basic properties of aqueous solutions are dependent upon an equilibrium that involves the solvent, water. Water, when pure or as a solvent, tends to dissociate to some extent into hydrogen ions and hydroxide ions:

$$H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq) \tag{17.1}$$

The forward reaction proceeds only slightly before equilibrium is reached. Only a small fraction of the total number of water molecules is dissociated.

Applying the general rules from Chapter 13 for equilibrium systems, we can write the equilibrium constant expression for Reaction 17.1 as

$$K_c = \frac{[H^+] \times [OH^-]}{[H_2O]}$$

In aqueous solutions, the concentration of H_2O is very high, typically about 55 M, and is essentially constant. Hence the term $[H_2O]$ can be combined with K_c to give a new constant, K_w, called the dissociation constant of water:

$$K_c \times [H_2O] = K_w = [H^+] \times [OH^-] = 1.0 \times 10^{-14} \tag{17.2}$$

In the equation for K_w, $[H_2O]$ is always omitted

At 25°C, K_w is 1.0×10^{-14}. This small value reflects the slight dissociation of water into H^+ and OH^- ions. In any aqueous solution or in pure water, the product of $[H^+]$ times $[OH^-]$ at 25°C is always 1.0×10^{-14}.

We can readily calculate $[H^+]$ and $[OH^-]$ in pure water. From Equation 17.1, we see that equal amounts of these two ions form when water dissociates. Applying Equation 17.2 to pure water:

$$[H^+] = [OH^-]; [H^+] \times [OH^-] = [H^+]^2 = 1.0 \times 10^{-14}$$

$$[H^+] = 1.0 \times 10^{-7}\ M = [OH^-]$$

Any aqueous solution in which $[H^+]$ *equals* $[OH^-]$ is called a *neutral* solution. It has a $[H^+]$ equal to 1.0×10^{-7} M at 25°C.

Ordinarily, the concentrations of H^+ and OH^- in a solution are not equal. Note from Equation 17.2 that as the concentration of one ion goes up, that of the other must go down so that the product $[H^+] \times [OH^-]$ stays constant. If the concentration of one ion is known, then that of the other can be calculated using Equation 17.2 (Example 17.1).

Example 17.1 An aqueous solution has a $[H^+]$ of 2.0×10^{-4} M. What is its $[OH^-]$?

Solution From Equation 17.2:

$$[OH^-] = \frac{K_w}{[H^+]} = \frac{1.0 \times 10^{-14}}{2.0 \times 10^{-4}} = 5.0 \times 10^{-11}\ M$$

Exercise What is $[H^+]$ in a solution in which $[OH^-]$ is 2.5×10^{-7} M? Answer: 4.0×10^{-8} M.

An aqueous solution, like that described in Example 17.1 where $[H^+]$ *is greater than* $[OH^-]$, is said to be *acidic*. An aqueous solution in which $[OH^-]$ *is greater than* $[H^+]$ is *basic*. Therefore,

if $[H^+] > 1.0 \times 10^{-7}$ M, $[OH^-] < 1.0 \times 10^{-7}$ M, **solution is acidic**

if $[OH^-] > 1.0 \times 10^{-7}$ M, $[H^+] < 1.0 \times 10^{-7}$ M, **solution is basic**

Table 17.1 indicates some possible combinations of concentrations of these ions. Solutions 1 through 4 are decreasingly acidic, 5 is neutral and 6 through 9 are progressively more basic.

17.2 pH

As we have seen, the acidity or basicity of a solution can be described by giving its H^+ ion concentration. Sorensen, in 1909, proposed an alternative method of accomplishing this purpose. He suggested using a term called pH, defined as

$$pH = -\log_{10}[H^+] = \log_{10} 1/[H^+] \qquad (17.3)$$

Thus we have

$$[H^+] = 10^{-4}\text{ M};\ pH = -\log_{10}(10^{-4}) = -(-4) = 4$$
$$[H^+] = 10^{-7}\text{ M};\ pH = -\log_{10}(10^{-7}) = -(-7) = 7$$
$$[H^+] = 10^{-10}\text{ M};\ pH = -\log_{10}(10^{-10}) = -(-10) = 10$$

A scientist will usually refer to a solution in terms of its pH rather than its $[H^+]$

Most aqueous solutions have hydrogen ion concentrations between 1 M and 10^{-14} M. By Equation 17.3, such solutions have pH's lying between 0 and 14. In this case, it is perhaps more convenient to express acidity in terms of pH rather than $[H^+]$. This avoids using small fractions or negative exponents.

Looking at the pH values in Table 17.1, we see that $[H^+]$ and pH are inversely

TABLE 17.1 RELATIONS BETWEEN $[H^+]$, $[OH^-]$, and pH IN AQUEOUS SOLUTIONS

	SOLUTION								
	No. 1	No. 2	No. 3	No. 4	No. 5	No. 6	No. 7	No. 8	No. 9
$[H^+]$	10^0	10^{-2}	10^{-4}	10^{-6}	10^{-7}	10^{-8}	10^{-10}	10^{-12}	10^{-14}
$[OH^-]$	10^{-14}	10^{-12}	10^{-10}	10^{-8}	10^{-7}	10^{-6}	10^{-4}	10^{-2}	10^0
pH	0	2	4	6	7	8	10	12	14
	Acidic				Neutral	Basic			

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