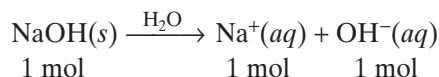## Calculating [H₃O⁺] and [OH⁻]

Recall that strong acids and bases are considered completely ionized or dissociated in weak aqueous solutions. A review of strong acids and bases is given in **Table 2.** Because NaOH is a strong base, 1 mol of it will yield 1 mol of OH⁻ in an aqueous solution.

$$\text{NaOH}(s) \xrightarrow{H_2O} \text{Na}^+(aq) + \text{OH}^-(aq)$$
$$1\text{ mol} \qquad 1\text{ mol} \qquad 1\text{ mol}$$

Therefore, a 1.0×10^{-2} M NaOH solution has an [OH⁻] of 1.0×10^{-2} M, as shown by the following.

$$\frac{1.0 \times 10^{-2} \text{ mol NaOH}}{1 \text{ L solution}} \times \frac{1 \text{ mol OH}^-}{1 \text{ mol NaOH}} = \frac{1.0 \times 10^{-2} \text{ mol OH}^-}{1 \text{ L solution}}$$
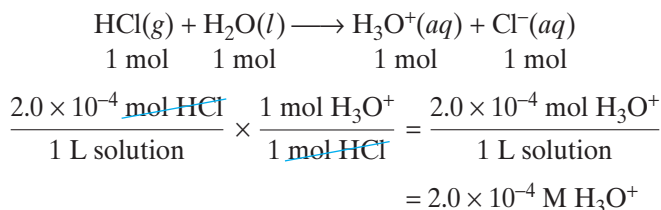$$= 1.0 \times 10^{-2} \text{ M OH}^-$$

Notice that the [OH⁻] is greater than 1.0×10^{-7} M. This solution is basic.

Because the K_w of an aqueous solution is a relatively constant 1.0×10^{-14} at ordinary room temperatures, the concentration of either ion can be determined if the concentration of the other ion is known. The [H₃O⁺] of this solution is calculated as follows.

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$$

$$[\text{H}_3\text{O}^+] = \frac{1.0 \times 10^{-14}}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-2}} = 1.0 \times 10^{-12} \text{ M}$$

The [OH⁻], 1.0×10^{-2} M, is greater than the [H₃O⁺], 1.0×10^{-12} M, as is true for all basic solutions.

Now consider a 2.0×10^{-4} M HCl solution. Because HCl is a strong acid, the [H₃O⁺] is 2.0×10^{-4} M, as shown by the following.

$$\text{HCl}(g) + \text{H}_2\text{O}(l) \longrightarrow \text{H}_3\text{O}^+(aq) + \text{Cl}^-(aq)$$
$$1\text{ mol} \qquad 1\text{ mol} \qquad 1\text{ mol} \qquad 1\text{ mol}$$

$$\frac{2.0 \times 10^{-4} \text{ mol HCl}}{1 \text{ L solution}} \times \frac{1 \text{ mol H}_3\text{O}^+}{1 \text{ mol HCl}} = \frac{2.0 \times 10^{-4} \text{ mol H}_3\text{O}^+}{1 \text{ L solution}}$$
$$= 2.0 \times 10^{-4} \text{ M H}_3\text{O}^+$$

Notice that the [H₃O⁺] is greater than 1.0×10^{-7} M. This solution is acidic. The [OH⁻] of this solution is calculated as follows.

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$$

$$[\text{OH}^-] = \frac{1.0 \times 10^{-14}}{[\text{H}_3\text{O}^+]} = \frac{1.0 \times 10^{-14}}{2.0 \times 10^{-4}} = 5.0 \times 10^{-10} \text{ M}$$

The [H₃O⁺] is greater than the [OH⁻] for all acidic solutions.

You may have realized that in order for K_w to remain constant, an increase in either the [H₃O⁺] or the [OH⁻] in an aqueous solution causes a decrease in the concentration of the other ion. Sample Problem A also shows calculation of the [H₃O⁺] and [OH⁻] of an acidic solution.

TABLE 2 Common Strong Acids and Bases	
Strong Acids	**Strong Bases**
HCl	LiOH
HBr	NaOH
HI	KOH
$HClO_4$	RbOH
$HClO_3$	CsOH
HNO_3	$Ca(OH)_2$
H_2SO_4	$Sr(OH)_2$
	$Ba(OH)_2$