In the section on amphiprotic species, we saw that water can act as a very weak acid and a very weak base, donating protons to itself to a limited extent:

\[
2\text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{OH}^-
\]

Applying the equilibrium law to this reaction, we obtain

\[
K_c = \frac{[\text{H}_3\text{O}^+][\text{OH}^-]}{[\text{H}_2\text{O}]}^{2}
\]

However, as can be seen in the section on the law of chemical equilibrium, the concentration of water has a constant value of 55.5 mol/L, and so its square can be multiplied by \(K_c\) to give a new constant \(K_w\), called the ion-product constant of water:

\[
K_w = K_c \times (55.5 \text{ mol L}^{-1})^2
\]

Measurements of the electrical conductivity of carefully purified water indicate that at 25°C \([\text{H}_3\text{O}^+] = [\text{OH}^-] = 1.00 \times 10^{-7}\) mol/L, so that

\[
K_w = 1.00 \times 10^{-14} \text{ mol}^2\text{L}^{-2}
\]

(Since the equilibrium law is not obeyed exactly, even in dilute solutions, results of most equilibrium calculations are rounded to three significant figures. Hence the value of \(K_w = 1.00 \times 10^{-14}\) mol²/L² is sufficiently accurate for all such calculations.)

The equilibrium constant \(K_w\) applies not only to pure water but to any aqueous solution at 25°C. Thus, for example, if we add 1.00 mol of the strong acid HNO₃ to H₂O to make a total volume of 1 L, essentially all the HNO₃ molecules donate their protons to H₂O:

\[
\text{HNO}_3 + \text{H}_2\text{O} \rightarrow \text{NO}_3^- + \text{H}_3\text{O}^+
\]

and a solution in which \([\text{H}_3\text{O}^+] = 1.00 \text{ mol/L}\) is obtained. Although this solution is very acidic, there are still hydroxide ions present. We can calculate their concentration by rearranging Eq. (3):

\[
[\text{OH}^-] = \frac{K_w}{[\text{H}_3\text{O}^+]} = \frac{1.00 \times 10^{-14} \text{ mol}^2\text{L}^{-2}}{1.00 \text{ mol L}^{-1}} = 1.00 \times 10^{-14} \text{ mol}
\]
The addition of the HNO$_3$ to H$_2$O not only increases the hydronium-ion concentration but also reduces the hydroxide-ion concentration from an initially minute $10^{-7}$ mol/L to an even more minute $10^{-14}$ mol/L.

Example (PageIndex{2}): Ion Concentration

Calculate the hydronium-ion concentration in a solution of 0.306 $M$ Ba(OH)$_2$.

**Solution**

Since 1 mol Ba(OH)$_2$ produces 2 mol OH$^-$ in solution, we have

$$[\text{OH}^-] = 2 \times 0.306 \text{ mol/L} = 0.612 \text{ mol/L}$$

Then

$$[\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.00 \times 10^{-14} \text{ mol}^2 \text{ L}^{-2}}{0.612 \text{ mol L}^{-1}} = 1.63 \times 10^{-14} \text{ mol L}^{-1}$$

Note

Note that since strong acids like HNO$_3$ are completely converted to H$_3$O$^+$ in aqueous solution, it is a simple matter to determine [H$_3$O$^+$], and from it, [OH$^-$]. Similarly, when a strong base dissolves in H$_2$O it is entirely converted to OH$^-$, so that [OH$^-$], and from it [H$_3$O$^+$] are easily obtained.

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