Chemistry Lecture #94: Common Ion Effect

Hydrofluoric acid is a weak acid that ionizes in water to form hydrogen ion and fluoride.

$$HF \rightleftharpoons H^+ + F^-$$

If HF exists in equilibrium with H^+ and F^-, would happen if KF were added to the solution? Adding KF is the equivalent of only adding F^-. This is because KF is soluble in water and forms K^+ and F^-. K^+ doesn't react with anything. It is a spectator ion that can be ignored.

According to Le Chatelier's principle, if product is added to the system, the equilibrium would shift toward the reactants (to the left). Thus, adding F^- would shift the equilibrium to the left and more HF would be made.

$$\begin{array}{ccccc} & & & & F^- \\ & & & & + \\ HF & \rightleftharpoons & H^+ & + & F^- \end{array}$$

Both HF and KF share the common ion F^-. The addition of this common ion that causes the equilibrium to shift toward the opposite side of the equation is the common ion effect.

We can calculate the concentration of product after a common ion has been added. For these problems, remember that the ions from group IA (Li^+, Na^+, K^+) are soluble in water and can be ignored as spectator ions. We also assume that all of the common ion that has been added is dissolved in the solution, and that the acid dissociates after the common ion has been added.

What is the H^+ concentration in a solution that is 0.200 M HF and 0.100M KF? Ka = 6.61×10^{-4}

Answer

The K in KF can be ignored. We assume 0.100 M KF means 0.100 M F^- has been added.

We solve the problem like a regular Ka problem, but make sure to include 0.100 M F^- as part of the concentration of F^-. Let x = the amount of HF that ionizes.

$$\underset{HF}{0.200 - x} \rightleftharpoons \underset{H^+}{x} + \underset{F^-}{x + 0.100}$$

$$Ka = \frac{[H^+][F^-]}{[HF]}$$

$$6.61 \times 10^{-4} = \frac{[x][x+0.100]}{0.200 - x}$$

Since 6.62×10^{-4} is such a small number, it means x is a negligible number.
$x + 0.100 \approx 0.100$, & $0.200 - x \approx 0.200$

$$6.61 \times 10^{-4} = \frac{[x][0.100]}{0.200}$$

$x = 1.32 \times 10^{-3}$ M H^+

What is the H^+ concentration in a solution that is 0.100 M HN_3 and 0.0500 M NaN_3? $Ka = 1.90 \times 10^{-5}$

Answer

Ignore the Na in NaN_3 and assume 0.05 M of N_3^- has been added.

$$\underset{HN_3}{0.100 - x} \rightleftharpoons \underset{H^+}{x} + \underset{N_3^-}{x + 0.0500}$$

$$Ka = \frac{[H^+][N_3^-]}{[HN_3]}$$

$$1.90 \times 10^{-5} = \frac{[x][x+0.0500]}{[0.100 - x]}$$ x is a negligible value

$$1.90 \times 10^{-5} = \frac{[x][0.0500]}{0.100}$$

$x = 3.80 \times 10^{-6}$ M H^+