Common Ion Effect

Introduction

The common-ion effect is used to describe the effect on an equilibrium involving a substance that adds an ion that is a part of the equilibrium. Adding a common ion prevents the weak acid or weak base from ionizing as much as it would without the added common ion. LeChatelier’s Principle states that if an equilibrium gets out of balance it will shift to restore the balance. If a common ion is added to a weak acid or weak base equilibrium, then the equilibrium will shift towards the reactants, in this case the weak acid or base.

Common Ion Effect on Weak Acids and Bases

The common ion effect suppresses the ionization of a weak acid by adding more of an ion that is a product of this equilibrium.

EXAMPLE

The common ion effect of $\text{H}_3\text{O}^+$ on the ionization of acetic acid:
The common ion effect suppresses the ionization of a weak base by adding more of an ion that is a product of this equilibrium.

**EXAMPLE**

The common ion effect of OH\(^{-}\) on the ionization of ammonia:
When a strong base supplies the common ion $\text{OH}^-$, the equilibrium shifts to form more NH

Therefore, adding the common ion of hydroxide, shifts the reaction towards the left to decrease the stress (Le Chatelier’s Principle), forming more reactants. This decreases the reaction quotient, for the reaction is being pushed towards the left to reach equilibrium. The equilibrium constant, $K_b=1.8\times10^{-5}$, does not change. The reaction is put out of balance, or equilibrium.

$$Q_A = \frac{([\text{NH}_4^+][\text{OH}^-])}{[\text{NH}_3]}$$

At first, when more hydroxide is added, the quotient is greater than the equilibrium constant. The reaction then shifts right, causing the denominator to increase, decreasing the reaction quotient and pulling towards equilibrium and causing $Q$ to decrease towards $K$.

**Common Ion Effect on Solubility**

Adding a common ion decreases solubility, as the reaction will shift towards the left to relieve the stress of the excess product. So, adding common ion to a dissociation,
causes the equilibrium to shift left, towards the reactants, causing precipitation. (Also see: Common Ion Effect on Solubility Product Constant).

**EXAMPLE**

Consider the reaction:

\[ \text{PbCl}_2(s) \rightleftharpoons \text{Pb}^{2+}(aq) + 2\text{Cl}^-(aq) \]

What happens to the solubility of \( \text{PbCl}_2(s) \) when you add .1M NaCl?

\( K_{sp} = 1.7 \times 10^{-5} \) \( Q_{sp} = 1.8 \times 10^{-5} \)

Identify the common ion: \( \text{Cl}^- \)

Notice: \( Q_{sp} > K_{sp} \) The addition of NaCl has caused the reaction to shift out of equilibrium since there are more ions dissociated.

Typically, the solve for the molarities of the following reactions you would do the following assume the solubility of \( \text{PbCl}_2 \) to be equivalent to the concentration of \( \text{Pb}^{2+} \) produced because they are in a 1:1 ratio.

Since \( K_{sp} \) for the reaction is \( 1.7 \times 10^{-5} \), the overall reaction would be \( (s)(2s)^2 = 1.7 \times 10^{-5} \). Solving for \( s \), you would get, \( s = 1.62 \times 10^{-2} \text{ M} \). The coefficient on \( \text{Cl}^- \) is 2, so it is assumed that twice as much \( \text{Cl}^- \) is produced as \( \text{Pb}^{2+} \), hence the '2s.' You use the solubility equilibrium constant to solve for the molarities of the ions at equilibrium.

Now let’s go back to the common ion being added to the reaction at equilibrium. The molarity of \( \text{Cl}^- \) added would be .1M because \( \text{Na}^+ \) and \( \text{Cl}^- \) are in a 1:1 ratio in the ionic salt, \( \text{NaCl} \). Therefore, the OVERALL molarity of \( \text{Cl}^- \) would be \( 2s+.1 \), with \( 2s \) being the contribution of the chloride ion from the dissociation of lead chloride.

\[ Q_{sp} = [\text{Pb}^{2+}][\text{Cl}^-]^2 \]

\[ 1.8 \times 10^{-5} = (s)(2s + .1)^2 \]

\[ s = [\text{Pb}^{2+}] = 1.8 \times 10^{-3} \text{ M} , 2s = [\text{Cl}^-] = \text{approximately .1M} \]

Notice that the molarity of \( \text{Pb}^{2+} \) is lower when NaCl is added. The equilibrium constant stays the same because of the increased concentration of the chloride ion. To simplify the reaction, it can be assumed that \( [\text{Cl}^-] \) is approximately .1M since the formation of the chloride ion from the dissociation of lead chloride is so small. The reaction quotient for \( \text{PbCl}_2 \) is greater than the equilibrium constant because of the added \( \text{Cl}^- \). This therefore
shift the reaction left towards equilibrium, causing precipitation and lowering the current solubility of the reaction. Overall, the solubility of the reaction decreases with the added sodium chloride.

References


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