Learning Objectives

• To understand why the solubility of many compounds depends on pH.

The solubility of many compounds depends strongly on the pH of the solution. For example, the anion in many sparingly soluble salts is the conjugate base of a weak acid that may become protonated in solution. In addition, the solubility of simple binary compounds such as oxides and sulfides, both strong bases, is often dependent on pH. In this section, we discuss the relationship between the solubility of these classes of compounds and pH.

The Effect of Acid–Base Equilibria the Solubility of Salts

We begin our discussion by examining the effect of pH on the solubility of a representative salt, \(\ce{M^{+}A^{-}}\), where \(\ce{A^{-}}\) is the conjugate base of the weak acid \(\ce{HA}\). When the salt dissolves in water, the following reaction occurs:

\[
\ce{MA (s) \rightleftharpoons M^{+} (aq) + A^{-} (aq)} \quad \text{Equation 17.13a}
\]

with

\[
K_{sp} = [\ce{M^{+}}][\ce{A^{-}}] \quad \text{Equation 17.13b}
\]

The anion can also react with water in a hydrolysis reaction:

\[
\ce{A^{-} (aq) + H2O (l) \rightleftharpoons OH^{-} (aq) + HA (aq)} \quad \text{Equation 17.14}
\]

Because of the reaction described in Equation \(\text{ref{17.14}}\), the predicted solubility of a sparingly soluble salt that has a basic anion such as \(\ce{S^{2-}}, \ce{PO_{4}^{3-}}, \text{or} \ce{CO_{3}^{2-}}\) is increased. If instead a strong acid is added to the solution, the added \(\ce{H^{+}}\) will react essentially completely with \(\ce{A^{-}}\) to form \(\ce{HA}\). This reaction decreases \([\ce{A^{-}}]\), which decreases the magnitude of the ion product

\[
Q = [\ce{M^{+}}][\ce{A^{-}}]
\]

According to Le Chatelier's principle, more MA will dissolve until \(Q = K_{sp}\)). Hence an acidic pH dramatically increases the solubility of virtually all sparingly soluble salts whose anion is the conjugate base of a weak acid. In contrast, pH has little to no effect on the solubility of salts whose anion is the conjugate base of a stronger weak acid or a strong acid, respectively (e.g., chlorides, bromides, iodides, and sulfates). For example, the hydroxide salt Mg(OH)\(_2\) is relatively insoluble in water:

\[
\ce{Mg(OH)_{2(s)} \rightleftharpoons Mg^{2+} (aq) + 2OH^{-} (aq)} \quad \text{Equation 17.15a}
\]

with

\[
K_{sp} = 5.61 \times 10^{−12} \quad \text{Equation 17.15b}
\]

When acid is added to a saturated solution that contains excess solid Mg(OH)\(_2\), the following reaction occurs, removing...
OH\textsuperscript{−} from solution:

\[\text{H}^+ (\text{aq}) + \text{OH}^− (\text{aq}) \rightarrow \text{H}_2\text{O} (\text{l}) \\text{\ref{17.16}}\]

The overall equation for the reaction of Mg(OH)\textsubscript{2} with acid is thus

\[\text{Mg(OH)}_{2(s)} + 2\text{H}^+ (\text{aq}) \rightleftharpoons \text{Mg}^{2+} (\text{aq}) + 2\text{H}_2\text{O} (\text{l}) \\text{\ref{17.17}}\]

As more acid is added to a suspension of Mg(OH)\textsubscript{2}, the equilibrium shown in Equation \(\text{\ref{17.17}}\) is driven to the right, so more Mg(OH)\textsubscript{2} dissolves.

Such pH-dependent solubility is not restricted to salts that contain anions derived from water. For example, CaF\textsubscript{2} is a sparingly soluble salt:

\[\text{CaF}_{2(s)} \rightleftharpoons \text{Ca}^{2+} (\text{aq}) + 2\text{F}^− (\text{aq}) \\text{\ref{17.18a}}\]

with

\[\text{K}_{\text{sp}} = 3.45 \times 10^{−11} \\text{\ref{17.18b}}\]

When strong acid is added to a saturated solution of CaF\textsubscript{2}, the following reaction occurs:

\[\text{H}^+ (\text{aq}) + \text{F}^− (\text{aq}) \rightleftharpoons \text{HF} (\text{aq}) \\text{\ref{17.19}}\]

Because the forward reaction decreases the fluoride ion concentration, more CaF\textsubscript{2} dissolves to relieve the stress on the system. The net reaction of CaF\textsubscript{2} with strong acid is thus

\[\text{CaF}_{2(s)} + 2\text{H}^+ (\text{aq}) \rightarrow \text{Ca}^{2+} (\text{aq}) + 2\text{HF} (\text{aq}) \\text{\ref{17.20}}\]

Example \(\text{\PageIndex{1}}\) shows how to calculate the solubility effect of adding a strong acid to a solution of a sparingly soluble salt.

Sparingly soluble salts derived from weak acids tend to be more soluble in an acidic solution.

Example \(\text{\PageIndex{1}}\)

Lead oxalate (PbC\textsubscript{2}O\textsubscript{4}), lead iodide (PbI\textsubscript{2}), and lead sulfate (PbSO\textsubscript{4}) are all rather insoluble, with \(K_{\text{sp}}\) values of \(4.8 \times 10^{−10}\), \(9.8 \times 10^{−9}\), and \(2.53 \times 10^{−8}\), respectively. What effect does adding a strong acid, such as perchloric acid, have on their relative solubilities?

**Given:** \(K_{\text{sp}}\) values for three compounds

**Asked for:** relative solubilities in acid solution

**Strategy:**
Write the balanced chemical equation for the dissolution of each salt. Because the strongest conjugate base will be most affected by the addition of strong acid, determine the relative solubilities from the relative basicity of the anions.

**Solution**

The solubility equilibria for the three salts are as follows:

\[
PbC_2O_4(s) \rightleftharpoons Pb^{2+}(aq) + C_2O^{2−}(aq) \\
PbI_2(s) \rightleftharpoons Pb^{2+}(aq) + 2I^−(aq) \\
PbSO_4(s) \rightleftharpoons Pb^{2+}(aq) + SO^{2−}(aq)
\]

The addition of a strong acid will have the greatest effect on the solubility of a salt that contains the conjugate base of a weak acid as the anion. Because HI is a strong acid, we predict that adding a strong acid to a saturated solution of PbI\(_2\) will not greatly affect its solubility; the acid will simply dissociate to form H\(^+\)(aq) and the corresponding anion. In contrast, oxalate is the fully deprotonated form of oxalic acid (HO\(_2\)CCO\(_2\)H), which is a weak diprotic acid (pK\(_{a1}\) = 1.23 and pK\(_{a2}\) = 4.19). Consequently, the oxalate ion has a significant affinity for one proton and a lower affinity for a second proton. Adding a strong acid to a saturated solution of lead oxalate will result in the following reactions:

\[
C_2O^{2−}(aq) + H^+(aq) \rightarrow HO_2CCO^−(aq) \\
HO_2CCO^−(aq) + H^+(aq) \rightarrow HO_2CCO_2H(aq)
\]

These reactions will decrease [C\(_2\)O\(^{2−}\)], causing more lead oxalate to dissolve to relieve the stress on the system. The pK\(_a\) of HSO\(^{-}\) (1.99) is similar in magnitude to the pK\(_{a1}\) of oxalic acid, so adding a strong acid to a saturated solution of PbSO\(_4\) will result in the following reaction:

\[
SO^{2−}(aq) + H^+(aq) \rightleftharpoons HSO^−(aq)
\]

Because HSO\(^{-}\) has a pKa of 1.99, this reaction will lie largely to the left as written. Consequently, we predict that the effect of added strong acid on the solubility of PbSO\(_4\) will be significantly less than for PbC\(_2\)O\(_4\).

**Exercise**

Which of the following insoluble salts—AgCl, Ag\(_2\)CO\(_3\), Ag\(_3\)PO\(_4\), and/or AgBr—will be substantially more soluble in 1.0 M HNO\(_3\) than in pure water?

**Answer**

Ag\(_2\)CO\(_3\) and Ag\(_3\)PO\(_4\)
Caves and their associated pinnacles and spires of stone provide one of the most impressive examples of pH-dependent solubility equilibria (part (a) in Figure \(\PageIndex{1}\)). Perhaps the most familiar caves are formed from limestone, such as Carlsbad Caverns in New Mexico, Mammoth Cave in Kentucky, and Luray Caverns in Virginia. The primary reactions that are responsible for the formation of limestone caves are as follows:

\[
\ce{CO2(aq) + H2O (l) & H^+(aq) + HCO^-3(aq)} \label{17.21}
\]

\[
\ce{HCO^-3(aq) & H^+(aq) + CO^{2-}_3(aq)} \label{17.22}
\]

\[
\ce{Ca^{2+}(aq) + CO^{2-}_3(aq) & CaCO3(s)} \label{17.23}
\]

Limestone deposits that form caves consist primarily of \(\ce{CaCO_3}\) from the remains of living creatures such as clams and corals, which used it for making structures such as shells. When a saturated solution of \(\ce{CaCO_3}\) in CO\(_2\)-rich water rises toward Earth’s surface or is otherwise heated, CO\(_2\) gas is released as the water warms. \(\ce{CaCO_3}\) then precipitates from the solution according to the following equation (part (b) in Figure \(\PageIndex{1}\)):

\[
\ce{Ca^{2+}(aq) + 2HCO^-_3(aq) & rightleftharpoons CaCO_3(s) + CO_{2(g)} + H_2O (l)} \label{17.24}
\]

The forward direction is the same reaction that produces the solid called scale in teapots, coffee makers, water heaters, boilers, and other places where hard water is repeatedly heated.
Figure \(\PageIndex{1}\): The Chemistry of Cave Formation. (a) This cave in Campanet, Mallorca, Spain, and its associated formations are examples of pH-dependent solubility equilibriums. (b) A cave forms when groundwater containing atmospheric CO\(_2\), forming an acidic solution, dissolves limestone (CaCO\(_3\)) in a process that may take tens of thousands of years. As groundwater seeps into a cave, water evaporates from the solution of CaCO\(_3\) in CO\(_2\)-rich water, producing a supersaturated solution and a shift in equilibrium that causes precipitation of the CaCO\(_3\). The deposited limestone eventually forms stalactites and stalagmites.

When groundwater-containing atmospheric CO\(_2\) (Equations \(\ref{17.21}\) and \(\ref{17.22}\)) finds its way into microscopic cracks in the limestone deposits, CaCO\(_3\) dissolves in the acidic solution in the reverse direction of Equation \(\ref{17.24}\). The cracks gradually enlarge from 10–50 \(\mu\)m to 5–10 mm, a process that can take as long as 10,000 yr. Eventually, after about another 10,000 yr, a cave forms. Groundwater from the surface seeps into the cave and clings to the ceiling, where the water evaporates and causes the equilibrium in Equation \(\ref{17.24}\) to shift to the right. A circular layer of solid CaCO\(_3\) is deposited, which eventually produces a long, hollow spire of limestone called a stalactite that grows down from the ceiling. Below, where the droplets land when they fall from the ceiling, a similar process causes another spire, called a stalagmite, to grow up. The same processes that carve out hollows below ground are also at work above ground, in some cases producing fantastically convoluted landscapes like that of Yunnan Province in China (Figure \(\PageIndex{2}\)).
Acidic, Basic, and Amphoteric Oxides and Hydroxides

One of the earliest classifications of substances was based on their solubility in acidic versus basic solution, which led to the classification of oxides and hydroxides as being either basic or acidic. **Basic oxides** and hydroxides either react with water to produce a basic solution or dissolve readily in aqueous acid. **Acidic oxides** or hydroxides either react with water to produce an acidic solution or are soluble in aqueous base. There is a clear correlation between the acidic or the basic character of an oxide and the position of the element combined with oxygen in the periodic table. Oxides of metallic elements are generally basic oxides, and oxides of nonmetallic elements are acidic oxides. Compare, for example, the reactions of a typical metal oxide, cesium oxide, and a typical nonmetal oxide, sulfur trioxide, with water:

\[
\text{Cs}_2\text{O (s) + H}_2\text{O (l) } \rightarrow 2\text{Cs}^+ (aq) + 2\text{OH}^- (aq) \quad \text{(17.25)}
\]

\[
\text{SO}_3\text{(g) + H}_2\text{O (l) } \rightarrow \text{H}_2\text{SO}_4\text{(aq)} \quad \text{(17.26)}
\]

Cesium oxide reacts with water to produce a basic solution of cesium hydroxide, whereas sulfur trioxide reacts with
Metal oxides generally react with water to produce basic solutions, whereas nonmetal oxides produce acidic solutions.

The difference in reactivity is due to the difference in bonding in the two kinds of oxides. Because of the low electronegativity of the metals at the far left in the periodic table, their oxides are best viewed as containing discrete $\text{M}^{n+}$ cations and $\text{O}^{2−}$ anions. At the other end of the spectrum are nonmetal oxides; due to their higher electronegativities, nonmetals form oxides with covalent bonds to oxygen. Because of the high electronegativity of oxygen, however, the covalent bond between oxygen and the other atom, $\text{E}$, is usually polarized: $\text{E}^{δ+}\text{O}^{δ−}$. The atom $\text{E}$ in these oxides acts as a Lewis acid that reacts with the oxygen atom of water to produce an oxoacid. Oxides of metals in high oxidation states also tend to be acidic oxides for the same reason: they contain covalent bonds to oxygen. An example of an acidic metal oxide is $\text{MoO}_3$, which is insoluble in both water and acid but dissolves in strong base to give solutions of the molybdate ion ($\text{MoO}_4^{2−}$):

$$[\text{MoO}_3(s) + 2\text{OH}^− (aq) \rightarrow \text{MoO}_4^{2−}(aq) + \text{H}_2\text{O} (l)]$$

As shown in Figure (PageIndex{3}), there is a gradual transition from basic metal oxides to acidic nonmetal oxides as we go from the lower left to the upper right in the periodic table, with a broad diagonal band of oxides of intermediate character separating the two extremes. Many of the oxides of the elements in this diagonal region of the periodic table are soluble in both acidic and basic solutions; consequently, they are called amphoteric oxides (from the Greek amphi, meaning “both,” as in amphiprotic). Amphoteric oxides either dissolve in acid to produce water or dissolve in base to produce a soluble complex ion. (Radioactive elements are not classified.)
base to produce a soluble complex. As shown in Video \(\PageIndex{1}\), for example, mixing the amphoteric oxide Cr(OH)\(_3\) (also written as Cr\(_2\)O\(_3\)•3H\(_2\)O) with water gives a muddy, purple-brown suspension. Adding acid causes the Cr(OH)\(_3\) to dissolve to give a bright violet solution of Cr\(^{3+}\) (aq), which contains the [Cr(H\(_2\)O)\(_6\)]\(^{3+}\) ion, whereas adding strong base gives a green solution of the [Cr(OH)\(_4\)]\(^{-}\) ion. The chemical equations for the reactions are as follows:

\[
\text{Cr(OH)}_3(s) + 3\text{H}^+(aq) \rightarrow \text{Cr}^{3+}(aq) + 3\text{H}_2\text{O}(l) \label{17.28}\]

\[
\text{Cr(OH)}_3(s) + \text{OH}^-(aq) \rightarrow [\text{Cr(OH)}_4]^-\text{(aq)} \label{17.29}\]

**Video \(\PageIndex{1}\):** Chromium(III) Hydroxide [Cr(OH)\(_3\) or Cr\(_2\)O\(_3\)•3H\(_2\)O] is an Example of an Amphoteric Oxide. All three beakers originally contained a suspension of brownish purple Cr(OH)\(_3\)(s) (center). When concentrated acid (6 M H\(_2\)SO\(_4\)) was added to the beaker on the left, Cr(OH)\(_3\) dissolved to produce violet [Cr(H\(_2\)O)\(_6\)]\(^{3+}\) ions and water. The addition of concentrated base (6 M NaOH) to the beaker on the right caused Cr(OH)\(_3\) to dissolve, producing green [Cr(OH)\(_4\)]\(^{-}\) ions. For a more complete description, see [https://www.youtube.com/watch?v=IQNcLH6OZK0](https://www.youtube.com/watch?v=IQNcLH6OZK0)

Example \(\PageIndex{2}\)

Aluminum hydroxide, written as either Al(OH)\(_3\) or Al\(_2\)O\(_3\)•3H\(_2\)O, is amphoteric. Write chemical equations to describe the dissolution of aluminum hydroxide in (a) acid and (b) base.

**Given:** amphoteric compound
Asked for: dissolution reactions in acid and base

Strategy:

Using Equations \(\ref{17.28}\) and \(\ref{17.29}\) as a guide, write the dissolution reactions in acid and base solutions.

Solution

a. An acid donates protons to hydroxide to give water and the hydrated metal ion, so aluminum hydroxide, which contains three OH\(^{-}\) ions per Al, needs three H\(^{+}\) ions:

\[
[\text{Al(OH)}_{3(s)} + 3\text{H}^{+} (\text{aq}) \rightarrow \text{Al}^{3+} (\text{aq}) + 3\text{H}_2\text{O} (\text{l})]
\]

In aqueous solution, Al\(^{3+}\) forms the complex ion \([\text{Al(H}_2\text{O})_6]^{3+}\).

b. In basic solution, OH\(^{-}\) is added to the compound to produce a soluble and stable poly(hydroxo) complex:

\[
[\text{Al(OH)}_{3(s)} + \text{OH}^{-} (\text{aq}) \rightarrow [\text{Al(OH)}_4]^{-} (\text{aq})]
\]

Exercise \(\PageIndex{2}\)

Copper(II) hydroxide, written as either Cu(OH)\(_2\) or CuO•H\(_2\)O, is amphoteric. Write chemical equations that describe the dissolution of cupric hydroxide both in an acid and in a base.

Answer

\[
[\text{Cu(OH)}_{2(s)} + 2\text{H}^{+} (\text{aq}) \rightarrow \text{Cu}^{2+} (\text{aq}) + 2\text{H}_2\text{O} (\text{l})]
\]

\[
[\text{Cu(OH)}_{2(s)} + 2\text{OH}^{-} (\text{aq}) \rightarrow [\text{Cu(OH)}_4]^{2-} (\text{aq})]
\]

Selective Precipitation Using pH

Many dissolved metal ions can be separated by the selective precipitation of the cations from solution under specific conditions. In this technique, pH is often used to control the concentration of the anion in solution, which controls which cations precipitate.

The concentration of anions in solution can often be controlled by adjusting the pH, thereby allowing the selective precipitation of cations.

Suppose, for example, we have a solution that contains 1.0 mM Zn\(^{2+}\) and 1.0 mM Cd\(^{2+}\) and want to separate the two metals by selective precipitation as the insoluble sulfide salts, ZnS and CdS. The relevant solubility equilibria can be written as follows:

\[
[Z\text{nS (s)} \rightleftharpoons \text{Zn}^{2+} (\text{aq}) + \text{S}^{2-} (\text{aq})]
\]
with
\[ K_{sp} = 1.6 \times 10^{-24} \quad \text{label(17.30b)} \]
and
\[ CdS (s) \rightleftharpoons Cd^{2+} (aq) + S^{2-} (aq) \quad \text{label(17.31a)} \]
with
\[ K_{sp} = 8.0 \times 10^{-27} \quad \text{label(17.31b)} \]
Because the S^{2-} ion is quite basic and reacts extensively with water to give HS^{-} and OH^{-}, the solubility equilibria are more accurately written as \( MS (s) \rightleftharpoons M^{2+} (aq) + HS^{-} (aq) + OH^{-}\) rather than \( MS (s) \rightleftharpoons M^{2+} (aq) + S^{2-} (aq) \). Here we use the simpler form involving S^{2-}, which is justified because we take the reaction of S^{2-} with water into account later in the solution, arriving at the same answer using either equilibrium equation.

The sulfide concentrations needed to cause \( ZnS \) and \( CdS \) to precipitate are as follows:
\[ K_{sp} = [Zn^{2+}][S^{2-}] \quad \text{label(17.32a)} \]
\[ 1.6 \times 10^{-24} = (0.0010 \; M)[S^{2-}] \quad \text{label(17.32b)} \]
\[ 1.6 \times 10^{-21} \; M = [S^{2-}] \quad \text{label(17.32c)} \]
and
\[ K_{sp} = [Cd^{2+}][S^{2-}] \quad \text{label(17.33a)} \]
\[ 8.0 \times 10^{-27} = (0.0010 \; M)[S^{2-}] \quad \text{label(17.33b)} \]
\[ 8.0 \times 10^{-24} \; M = [S^{2-}] \quad \text{label(17.33c)} \]
Thus sulfide concentrations between \( 1.6 \times 10^{-21} \) M and \( 8.0 \times 10^{-24} \) M will precipitate CdS from solution but not ZnS. How do we obtain such low concentrations of sulfide? A saturated aqueous solution of H_{2}S contains \( 0.10 \; M \) H_{2}S at 20°C. The pK_{a1} for H_{2}S is 6.97, and pK_{a2} corresponding to the formation of [S^{2-}] is 12.90. The equations for these reactions are as follows:
\[ H_{2}S (aq) \rightleftharpoons H^{+} (aq) + HS^{-} (aq) \quad \text{label(17.34a)} \]
with
\[ pK_{a1} = 6.97; \; \text{and hence} \; K_{a1} = 1.1 \times 10^{-7} \quad \text{label(17.34b)} \]
\[ HS^{-} (aq) \rightleftharpoons H^{+} (aq) + S^{2-} (aq) \quad \text{label(17.34c)} \]
\( pK_{a2} = 12.90 \) \( \text{(and hence)} \) \( K_{a2} = 1.3 \times 10^{-13} \) \( \label{17.34d} \)

We can show that the concentration of \( S^{2-} \) is \( 1.3 \times 10^{-13} \) by comparing \( K_{a1} \) and \( K_{a2} \) and recognizing that the contribution to \( [H^+] \) from the dissociation of \( HS^- \) is negligible compared with \( [H^+] \) from the dissociation of \( H_2S \). Thus substituting 0.10 M in the equation for \( K_{a1} \) for the concentration of \( H_2S \), which is essentially constant regardless of the pH, gives the following:

\[
[K_{\text{a1}}]=1.1 \times 10^{-7} \dfrac{[\text{H}^+][\text{HS}^-]}{[\text{H}_2\text{S}]}=\dfrac{x^2}{0.10 \text{ M}}
\]

\[
x=1.1 \times 10^{-4} \text{ M}=[\text{H}^+]=[\text{HS}^-] \label{17.35}
\]

Substituting this value for \( [H^+] \) and \( [HS^-] \) into the equation for \( K_{a2} \),

\[
[K_{\text{a2}}]=1.3 \times 10^{-13} \dfrac{[\text{H}^+][\text{S}^{2-}]}{[\text{HS}^-]}=\dfrac{(1.1 \times 10^{-4} \text{ M})x}{1.1 \times 10^{-4} \text{ M}}=x=[\text{S}^{2-}]
\]

Although \( [S^{2-}] \) in an \( H_2S \) solution is very low \( (1.3 \times 10^{-13} \text{ M}) \), bubbling \( H_2S \) through the solution until it is saturated would precipitate both metal ions because the concentration of \( S^{2-} \) would then be much greater than \( 1.6 \times 10^{-21} \text{ M} \). Thus we must adjust \( [S^{2-}] \) to stay within the desired range. The most direct way to do this is to adjust \( [H^+] \) by adding acid to the \( H_2S \) solution (recall \textit{Le Chatelier's principle}), thereby driving the equilibrium in Equation \( \ref{17.34d} \) to the left. The overall equation for the dissociation of \( H_2S \) is as follows:

\[
[H_2S \text{ (aq)} \rightleftharpoons 2H^+ \text{ (aq)} + S^{2-} \text{ (aq)}] \label{17.36}
\]

Now we can use the equilibrium constant \( K \) for the overall reaction, which is the product of \( K_{a1} \) and \( K_{a2} \), and the concentration of \( H_2S \) in a saturated solution to calculate the \( H^+ \) concentration needed to produce \( [S^{2-}] \) of \( 1.6 \times 10^{-21} \text{ M} \):

\[
[K]=K_{\text{a1}}K_{\text{a2}}=(1.1 \times 10^{-7})(1.3 \times 10^{-13})=1.4 \times 10^{-20} \dfrac{[\text{H}^+]^2[\text{S}^{2-}]}{[\text{H}_2\text{S}]} \label{17.37}
\]

\[
\begin{align}
[\text{H}^+]^2&=\dfrac{K[\text{H}_2\text{S}]}{[\text{S}^{2-}]}=\dfrac{(1.4 \times 10^{-20})(0.10 \text{ M})}{1.6 \times 10^{-21}}\text{ M} \\
[\text{H}^+]&=0.94
\end{align} \label{17.38}
\]

Thus adding a strong acid such as HCl to make the solution 0.94 M in \( H^+ \) will prevent the more soluble ZnS from precipitating while ensuring that the less soluble CdS will precipitate when the solution is saturated with \( H_2S \).

Example \( \PageIndex{3} \)

A solution contains 0.010 M \( Ca^{2+} \) and 0.010 M \( La^{3+} \). What concentration of HCl is needed to precipitate \( La_2(C_2O_4)_3\cdot9H_2O \) but not \( Ca(C_2O_4)\cdot5H_2O \) if the concentration of oxalic acid is 1.0 M? \( K_{sp} \) values are \( 2.32 \times 10^{-9} \) for \( Ca(C_2O_4) \) and \( 2.5 \times 10^{-27} \) for \( La_2(C_2O_4)_3 \); \( pK_{a1} = 1.25 \) and \( pK_{a2} = 3.81 \) for oxalic acid.

\textbf{Given:} concentrations of cations, \( K_{sp} \) values, and concentration and \( pK_a \) values for oxalic acid
**Asked for:** concentration of HCl needed for selective precipitation of La$_2$(C$_2$O$_4$)$_3$

**Strategy:**

A. Write each solubility product expression and calculate the oxalate concentration needed for precipitation to occur. Determine the concentration range needed for selective precipitation of La$_2$(C$_2$O$_4$)$_3$•9H$_2$O.

B. Add the equations for the first and second dissociations of oxalic acid to get an overall equation for the dissociation of oxalic acid to oxalate. Substitute the [ox$^{2-}$] needed to precipitate La$_2$(C$_2$O$_4$)$_3$•9H$_2$O into the overall equation for the dissociation of oxalic acid to calculate the required [H$^+$].

**Solution**

A Because the salts have different stoichiometries, we cannot directly compare the magnitudes of the solubility products. Instead, we must use the equilibrium constant expression for each solubility product to calculate the concentration of oxalate needed for precipitation to occur. Using ox$^{2-}$ for oxalate, we write the solubility product expression for calcium oxalate as follows:

$$K_{sp} = [Ca^{2+}][ox^{2-}] = (0.010)[ox^{2-}] = 2.32 \times 10^{-9}$$

$$[ox^{2-}] = 2.32 \times 10^{-7} \text{ M}$$

The expression for lanthanum oxalate is as follows:

$$K_{sp} = [La^{3+}]^2[ox^{2-}]^3 = (0.010)^2[ox^{2-}]^3 = 2.5 \times 10^{-27}$$

$$[ox^{2-}] = 2.9 \times 10^{-8} \text{ M}$$

Thus lanthanum oxalate is less soluble and will selectively precipitate when the oxalate concentration is between $2.9 \times 10^{-8} \text{ M}$ and $2.32 \times 10^{-7} \text{ M}$.

B To prevent Ca$^{2+}$ from precipitating as calcium oxalate, we must add enough H$^+$ to give a maximum oxalate concentration of $2.32 \times 10^{-7}$ M. We can calculate the required [H$^+$] by using the overall equation for the dissociation of oxalic acid to oxalate:

$$HO_2CCO_2H (aq) \rightleftharpoons 2H^+ (aq) + C_2O^{2-}_{4(aq)}$$

$$K = K_{a1}K_{a2} = (10^{-1.25})(10^{-3.81}) = 10^{-5.06} = 8.7 \times 10^{-6}$$

Substituting the desired oxalate concentration into the equilibrium constant expression,

$$\begin{align}8.7 \times 10^{-6} = \dfrac{[\text{H}^+]^2[\text{ox}^{2-}]}{[\text{HO}_2\text{CCO}_2\text{H}]} &= \dfrac{[\text{H}^+]^2(2.32 \times 10^{-7})}{1.0} \\ [\text{H}^+] &= \text{6.1 M} \end{align}$$

Thus adding enough HCl to give [H$^+$] = 6.1 M will cause only La$_2$(C$_2$O$_4$)$_3$•9H$_2$O to precipitate from the solution.

Exercise \(\PageIndex{3}\)
A solution contains 0.015 M Fe$^{2+}$ and 0.015 M Pb$^{2+}$. What concentration of acid is needed to ensure that Pb$^{2+}$ precipitates as PbS in a saturated solution of H$_2$S, but Fe$^{2+}$ does not precipitate as FeS? $K_{sp}$ values are 6.3 $\times$ 10$^{-18}$ for FeS and 8.0 $\times$ 10$^{-28}$ for PbS.

Answer

0.018 M H$^+$

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**Summary**

The anion in sparingly soluble salts is often the conjugate base of a weak acid that may become protonated in solution, so the solubility of simple oxides and sulfides, both strong bases, often depends on pH. The anion in many sparingly soluble salts is the conjugate base of a weak acid. At low pH, protonation of the anion can dramatically increase the solubility of the salt. Oxides can be classified as acidic oxides or basic oxides. Acidic oxides either react with water to give an acidic solution or dissolve in strong base; most acidic oxides are nonmetal oxides or oxides of metals in high oxidation states. Basic oxides either react with water to give a basic solution or dissolve in strong acid; most basic oxides are oxides of metallic elements. Oxides or hydroxides that are soluble in both acidic and basic solutions are called amphoteric oxides. Most elements whose oxides exhibit amphoteric behavior are located along the diagonal line separating metals and nonmetals in the periodic table. In solutions that contain mixtures of dissolved metal ions, the pH can be used to control the anion concentration needed to selectively precipitate the desired cation.

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**Contributors and Attributions**

- Anonymous