Skills to Develop

- Calculate ion concentrations to maintain a heterogeneous equilibrium.
- Calculate pH required to precipitate a metal hydroxide.
- Design experiments to separate metal ions in a solution of mixtures of metals.

Chemical Separation by Precipitation

A mixture of metal ions in a solution can be separated by precipitation with anions such as $\text{Cl}^-$, $\text{Br}^-$, $\text{SO}_4^{2-}$, $\text{CO}_3^{2-}$, $\text{S}^2-$, $\text{Cr}_2\text{O}_4^{2-}$, $\text{PO}_4^{2-}$, $\text{OH}^-$ etc. When a metal ion or a group of metal ions form insoluble salts with a particular anion, they can be separated from others by precipitation. We can also separate the anions by precipitating them with appropriate metal ions.

There are no definite dividing lines between **insoluble salts**, **sparingly soluble**, and **soluble salts**, but concentrations of their saturated solutions are small, medium, and large. Solubility products are usually listed for insoluble and sparingly soluble salts, but they are not given for soluble salts. Solubility products for soluble salts are very large.

What type of salts are usually soluble, sparingly soluble and insoluble? The following are some general guidelines, but these are not precise laws.

- **All nitrates are soluble.** The singly charged large $\text{NO}_3^-$ ions form salts with high solubilities. So do $\text{ClO}_4^-$, $\text{ClO}_3^-$, $\text{NO}_2^-$, $\text{HCOO}^-$, and $\text{CH}_3\text{COO}^-$. $\text{CaF}_2$, $\text{BaF}_2$, and $\text{PbF}_2$ are also insoluble.

- **All chlorides, bromides, and iodides are soluble except those of** $\text{Ag}^+$, $\text{Hg}_2^{2+}$, and $\text{Pb}^{2+}$. $\text{CaF}_2$, $\text{BaF}_2$, and $\text{PbF}_2$ are also insoluble.

- **All sulfates are soluble, except those of** $\text{Ba}^{2+}$, $\text{Sr}^{2+}$, and $\text{Pb}^{2+}$. The doubly charged sulfates are usually less soluble than halides and nitrates.

- **Most singly charge cations** $\text{K}^+$, $\text{Na}^+$, $\text{NH}_4^+$ form soluble salts. However, $\text{K}_3\text{Co(NO}_2)_6$ and $\text{NH}_4\text{H}_2\text{Co(NO}_2)_6$ are insoluble.

These are handy rules for us to have if we deal with salts often. On the other hand, solubility is an important physical property of a substance, and these properties are listed in handbooks.

Chemical Separation of Metal Ions

Formation of crystals from a saturated solution is a **heterogeneous equilibrium** phenomenon, and it can be applied to separate various chemicals or ions in a solution. When solubilities of two metal salts are very different, they can be separated by precipitation. The $K_{sp}$ values for various salts are valuable information, and some data are given in the Handbook of this website. On the Handbook Menu, clicking the item Salts $K_{sp}$ will give the $K_{sp}$’s of some salts. In the first two examples, we show how barium and strontium can be separated as chromate.

Example 1

The $K_{sp}$ for strontium chromate is $3.6\text{E-5}$ and the $K_{sp}$ for barium chromate is $1.2\text{E-10}$. What concentration of potassium
chromate will precipitate the maximum amount of either the barium or the strontium chromate from an equimolar 0.30 M solution of barium and strontium ions without precipitating the other?

**SOLUTION**

Since the $K_{sp}$ for barium chromate is smaller, we know that $\ce{BaCrO4}$ will form a precipitate first as $\ce{[CrO4^2-]}$ increases so that $Q_{sp}$ for $\ce{BaCrO4}$ also increases from zero to $K_{sp}$ of $\ce{BaCrO4}$, at which point, $\ce{BaCrO4}$ precipitates. As $\ce{[CrO4^2-]}$ increases, $\ce{[Ba^2+]}'$s decreases. Further increase of $\ce{[CrO4^2-]}$ till $Q_{sp}$ for $\ce{SrCrO4}$ increases to $K_{sp}$ of $\ce{SrCrO4}$; it then precipitates.

Let us write the equilibrium equations and data down to help us think. Further, let $x$ be the concentration of chromate to precipitate $\ce{Sr^2+}$, and $y$ be that to precipitate $\ce{Ba^2+}$). According to the definition of $K_{sp}$ we have,

\[
\begin{array}{cccccl}
\ce{SrCrO4 &\rightarrow &Sr^2+ &+ &CrO4^2-}, &K_{\ce{sp}} = 3.6\times 10^{-5}\\
&&0.30 &&x & \\
\end{array}
\]

\[
[x = \dfrac{\text{3.6e-5}}{0.30} = 1.2 \times 10^{-4} \; M]
\]

\[
\begin{array}{cccccl}
\ce{BaCrO4 &\rightarrow &Ba^2+ &+ &CrO4^2-}, &K_{\ce{sp}} = 1.2 \times 10^{-10}\\
&&0.30 &&y & \\
\end{array}
\]

\[
[y = \dfrac{\text{1.2e-10}}{0.30} = 4.0 \times 10^{-10} \; M]
\]

The $K_{sp}$'s for the two salts indicate $\ce{BaCrO4}$ to be much less soluble, and it will precipitate before any $\ce{SrCrO4}$ precipitates. If chromate concentration is maintained a little less than 1.2e-4 M, $\ce{Sr^2+}$ ions will remain in the solution.

**DISCUSSION**

In reality, to control the increase of $\ce{[CrO4^2-]}$ is very difficult.

Example 2

The $K_{sp}$ for strontium chromate is $\langle 3.6\times 10^{-4} \rangle$ and the $K_{sp}$ for barium chromate is $\langle 1.2\times 10^{-10} \rangle$. Potassium chromate is added a small amount at a time to first precipitate $\ce{BaCrO4})$. Calculate $\ce{[CrO4^2-]}$ when the first trace of $\ce{SrCrO4}$ precipitate starts to form in a solution that contains 0.30 M each of $\ce{Ba^2+}$ and $\ce{Sr^2+}$ ions.

**SOLUTION**

From the solution given in **Example 1**, $\langle 3.6\times 10^{-4} \rangle$ when $\ce{SrCrO4}$ starts to form. At this concentration, the $\ce{[CrO4^2-]}$ is estimated as follows.

\[
\langle \ce{Ba^2+} \rangle \times 3.6\times 10^{-4} = 1.2\times 10^{-10}
\]
The $K_{sp}$ of $\text{BaCrO}_4$.

Thus,

$$\text{Ba}^{2+} = 3.33 \times 10^{-7} \text{ M}$$

Very small indeed, compared to 0.30. In the fresh precipitate of $\text{SrCrO}_4$, the mole ratio of $\text{SrCrO}_4$ to $\text{BaCrO}_4$ is $0.30 / 3.33 \times 10^{-7} = 9.0 \times 10^5$. In other words, the amount of $\text{Ba}^{2+}$ ion in the solid is only $1 \times 10^{-6}$ (1 ppm) of all metal ions, providing that all the solid was removed when $\text{CrO}_4^{2-} = 3.6 \times 10^{-4} \text{ M}$.

**DISCUSSION**

The calculation shown here indicates that the separation of $\text{Sr}$ and $\text{Ba}$ is pretty good. In practice, an impurity level of 1 ppm is a very small value.

Example 3

What reagent should you use to separate silver and lead ions that are present in a solution? What data or information will be required for this task?

**SOLUTION**

The $K_{sp}$s for salts of silver and lead are required. We list the $K_{sp}$s for chlorides and sulfates in a table here. These values are found in the Handbook Menu of our website as Salts $K_{sp}$.

<table>
<thead>
<tr>
<th>Salt</th>
<th>$K_{sp}$</th>
<th>Salt</th>
<th>$K_{sp}$</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\text{AgCl}$</td>
<td>$1.8 \times 10^{-10}$</td>
<td>$\text{Ag}_2\text{SO}_4$</td>
<td>$1.4 \times 10^{-5}$</td>
</tr>
<tr>
<td>$\text{Hg}_2\text{Cl}_2$</td>
<td>$1.3 \times 10^{-18}$</td>
<td>$\text{BaSO}_4$</td>
<td>$1.1 \times 10^{-10}$</td>
</tr>
<tr>
<td>$\text{PbCl}_2$</td>
<td>$1.7 \times 10^{-5}$</td>
<td>$\text{CaSO}_4$</td>
<td>$2.4 \times 10^{-5}$</td>
</tr>
<tr>
<td></td>
<td></td>
<td>$\text{PbSO}_4$</td>
<td>$6.3 \times 10^{-7}$</td>
</tr>
<tr>
<td></td>
<td></td>
<td>$\text{SrSO}_4$</td>
<td>$3.2 \times 10^{-7}$</td>
</tr>
</tbody>
</table>

Because the $K_{sp}$s of $\text{AgCl}$ and $\text{PbCl}_2$ are very different, chloride, $\text{Cl}^-$, appears a good choice of negative ions for their separation.

The literature also indicates that $\text{PbCl}_2$ is rather soluble in warm water, and by heating the solution to 350 K (80°C), you can keep $\text{Pb}^{2+}$ ions in solution and precipitate $\text{AgCl}$ as a solid. The solubility of $\text{AgCl}$ is very small even at high temperatures.

**DISCUSSION**

Find more detailed information about the solubility of lead chloride as a function of temperature.
Can sulfate be used to separate silver and lead ions? Which one will form a precipitate first as the sulfate ion concentration increases? What is the \(\text{[Pb}^2+]\) when \(\text{[Ag}_2\text{SO}_4]\) begins to precipitate in a solution that contains 0.10 M \(\text{[Ag}^+]\)?

Questions

1. Iron(II) hydroxide is only sparingly soluble in water at 25°C; its \(K_{sp}\) is \((7.9 \times 10^{-16})\). Calculate the solubility of iron(II) hydroxide in a buffer solution with pH = 7.00.

2. A solution contains 0.60 M \(\text{[Ba}^2+]\) and 0.30 M \(\text{[Ca}^2+]\); \(K_{sp}\) values for \(\text{[BaCrO}_4\text{]}\) and \(\text{[CaCrO}_4\text{]}\) are \((1.2 \times 10^{-10})\) and \((7.1 \times 10^{-4})\) respectively. What value of \(\text{[CrO}_4^{2-}]\) will result in a maximum separation of these two ions?

3. A solution contains 0.60 M \(\text{[Ba}^2+]\) and 0.30 M \(\text{[Ca}^2+]\); \(K_{sp}\) values for \(\text{[BaCrO}_4\text{]}\) and \(\text{[CaCrO}_4\text{]}\) are \((1.2 \times 10^{-10})\) and \((7.1 \times 10^{-4})\) respectively. What is \(\text{[Ca}^2+] / \text{[Ba}^2+]\) when \(\text{[CrO}_4^{2-}]\) is maintained at \((1.2 \times 10^{-3})\)?

Solutions

1. **Answer** \(\text{[Fe}^2+] = 0.079\ M\)

   **Consider...**

   \(\text{[OH}^-] = 10^{(-14+7)} = \text{1.00e-7 (buffer)})\).

   \(\text{[Fe}^2+] \times (\text{1.00e-7})^2 = K_{sp}: \text{[Fe}^2+] = ?\)

   This \(\text{[Fe}^2+]\) concentration is low; it is not very soluble in a neutral solution (pH = 7).

   What is \(\text{[Fe}^2+]\) in a solution whose pH = 6.00?

2. **Answer** \(\text{[CrO}_4^{2-}] = 2.37e-3\ M\)

   **Consider...**

   Solid \(\text{[BaCrO}_4\text{]}\) will form first as \(\text{[CrO}_4^{2-}]\) increases. The maximum \(\text{[CrO}_4^{2-}]\) to precipitate \(\text{[CaCrO}_4\text{]}\) is estimated as follows.

   \(\text{[CrO}_4^{2-}] = \frac{\text{7.1e-4}}{0.30} = \text{2.37e-3 M}\)

   Estimate \(\text{[Ba}^2+]\) when \(\text{[CrO}_4^{2-}] = \text{2.3e-3 M}\), slightly below the maximum concentration.

3. **Answer** \(\text{[Ca}^2+] / \text{[Ba}^2+] = 3e6\)

   **Consider...**

   \(\text{[Ba}^2+] = \text{[Ca}^2+] \times (\text{1.2e-10}) / (\text{1.2e-3}) = \text{1e-7}\);

   \(\text{[Ca}^2+] = \text{[Ba}^2+] / (\text{0.3\times1e-7}) = ?\) The ratio of three million is large!

Contributors

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