• Ethylenediamine tetraacetate (EDTA\(^{4–}\)) is a ligand that forms complexes with many metal ions and consequently may be used to treat heavy metal toxicity in the body. The reaction with lead ions is represented by the following equilibrium:

\[
Pb^{2+} + EDTA^{4–} \rightleftharpoons [PbEDTA]^{2–}
\]

If a solution had an initial concentration of \(1 \times 10^{-4}\) M \(Pb^{2+}\) and 0.05 M EDTA, what will be the concentration of uncomplexed lead ions once equilibrium is established? \(K_{\text{stab}}\) for \([PbEDTA]^{2–}\) is \(1 \times 10^{18}\) M\(^{-1}\).

\(K_{\text{stab}}\) is very large so the amount of uncomplexed \(Pb^{2+}\) will be tiny.

If \([Pb^{2+}] = x\) M,

\[
[[PbEDTA]^{2–}] = (1 \times 10^4 – x) \text{ M} \approx 1 \times 10^4 \text{ M as } x \text{ is so small.}
\]

\[
[EDTA^{4–}] = 0.05 – (1 \times 10^4 – x) \text{ M} \approx 0.05 – (1 \times 10^4) \text{ M} = 0.0499 \text{ M}
\]

Hence,

\[
K_{\text{stab}} = \frac{[[PbEDTA]^{2–}]}{[Pb^{2+}][EDTA^{4–}]} = \frac{(1 \times 10^{-4})}{x(0.0499)} = 1 \times 10^{18}
\]

\[
x = [Pb^{2+}] \text{ M} = 2 \times 10^{-21} \text{ M}
\]

Answer: \([Pb^{2+}] = 2 \times 10^{-21} \text{ M}\)
• Calculate the molar solubility of Fe(OH)$_3$ in a pH = 5.0 buffer solution. The solubility product constant of Fe(OH)$_3$ is $4 \times 10^{-38}$ M$^4$.

Using $pOH = -\log_{10}([OH^-])$ and pH + pOH = 14.0:

\[ pOH = (14.0 - 5.0) = 9.0 \text{ and } [OH^-] = 1 \times 10^{-9} \text{ M} \]

The solubility equilibrium and product are:

\[ \text{Fe(OH)}_3(s) \rightleftharpoons \text{Fe}^{3+}(aq) + 3\text{OH}^-\text{(aq)} \quad \text{ } K_{sp} = [\text{Fe}^{3+}(aq)][\text{OH}^-\text{(aq)}]^3 \]

Hence,

\[ [\text{Fe}^{3+}(aq)] = \frac{K_{sp}}{[\text{OH}^-\text{(aq)}]^3} = \frac{(4 \times 10^{-38})}{(1.0 \times 10^{-9})^3} = 4 \times 10^{-11} \text{ M} \]

As Fe(OH)$_3$(s) dissolves to give 1 Fe$^{3+}$(aq), this is also the molar solubility.

Answer: $4 \times 10^{-11}$ M
The presence of iron in inorganic qualitative analysis is detected by the precipitation of the hydroxide using a buffer of pH 8. The solubility product constant of Fe(OH)$_3$ is $4 \times 10^{-38}$ M$^4$ and that of Fe(OH)$_2$ is $4 \times 10^{-15}$ M$^3$. Is it more sensible to try and detect the presence of Fe$^{2+}$ ions or Fe$^{3+}$ ions? Show all working and then give a reason for your answer.

Using pOH = $-\log_{10}([\text{OH}^-(aq)])$ and pH + pOH = 14.0:

\[ \text{pOH} = (14.0 - 8.0) = 6.0 \text{ and} \]
\[ [\text{OH}^-(aq)] = 1 \times 10^{-6} \text{ M} \]

The solubility equilibria and products are:

$\text{Fe(OH)}_2(s) \rightleftharpoons \text{Fe}^{2+}(aq) + 2\text{OH}^-(aq) \quad K_{sp}(\text{Fe(OH)}_2) = [\text{Fe}^{2+}(aq)][\text{OH}^-(aq)]^2$

$\text{Fe(OH)}_3(s) \rightleftharpoons \text{Fe}^{3+}(aq) + 3\text{OH}^-(aq) \quad K_{sp}(\text{Fe(OH)}_3) = [\text{Fe}^{3+}(aq)][\text{OH}^-(aq)]^3$

Hence,

\[ [\text{Fe}^{2+}(aq)] = \frac{K_{sp}(\text{Fe(OH)}_2)}{[\text{OH}^-(aq)]^2} = \frac{(4 \times 10^{-15})}{(1.0 \times 10^{-6})^2} = 4 \times 10^{-3} \text{ M} \]

\[ [\text{Fe}^{3+}(aq)] = \frac{K_{sp}(\text{Fe(OH)}_3)}{[\text{OH}^-(aq)]^3} = \frac{(4 \times 10^{-38})}{(1.0 \times 10^{-6})^3} = 4 \times 10^{-20} \text{ M} \]

As Fe(OH)$_2(s)$ Fe(OH)$_3(s)$ dissolves to give 1 Fe$^{2+}(aq)$ and 1 Fe$^{3+}(aq)$, these are also the molar solubilities.

The solubility of Fe(OH)$_3$ is much lower so it will precipitate at much lower iron concentrations. It is therefore easier to detect Fe$^{3+}$.

- Name the following complexes.

<table>
<thead>
<tr>
<th><a href="NO$_3$">Cr(OH$_2$)$_6$</a>$_3$</th>
<th>hexaaquachromium(III) nitrate</th>
</tr>
</thead>
<tbody>
<tr>
<td>[CoBr$_2$(en)$_2$]Cl</td>
<td>dibromidobis(ethylenediamine)cobalt(III) chloride</td>
</tr>
</tbody>
</table>

en = ethylenediamine = NH$_2$CH$_2$CH$_2$NH$_2$