CHE 131

Lecture 31 – Reactions in Aqueous Solutions

Chapter 5: pp. 165-185.

Table 5.1: Solubility Rules for Ionic Compounds

<table>
<thead>
<tr>
<th>Usually Soluble</th>
<th>Usually Insoluble</th>
</tr>
</thead>
<tbody>
<tr>
<td>Group IA, ammonium, NO₃⁻</td>
<td>All phosphates are insoluble except those of NH₄⁺ and Group 1A elements (alkali metals).</td>
</tr>
<tr>
<td>NH₄⁺, KO₂, RO₂⁻, Ca⁺², Sr⁺², Ba⁺²</td>
<td>Carbonates, CO₃⁻</td>
</tr>
<tr>
<td>Br⁻, I⁻, I₂⁻</td>
<td>All carbonates are insoluble except those of NH₄⁺ and Group 1A elements (alkali metals).</td>
</tr>
<tr>
<td>Chlorides, bromides, iodides, C₇⁻, Br⁻, I⁻</td>
<td>Hydroxides, OH⁻</td>
</tr>
<tr>
<td>All chlorides and bromides, and iodides are soluble; except AgCl, Hg₂Cl₂, PbCl₂, AgBr, Hg₂Br₂, PbBr₂, AgI, Hg₂I₂, PbI₂</td>
<td>All hydroxides are insoluble except those of NH₄⁺ and Group 1A (alkali metal cations), Sr(OH)₂, Ba(OH)₂, and Ca(OH)₂ are slightly soluble.</td>
</tr>
<tr>
<td>Sulfates, SO₄²⁻</td>
<td>Oxalates, C₂O₄²⁻</td>
</tr>
<tr>
<td>Most sulfates are soluble; exceptions include CaSO₄, MgSO₄, BaSO₄, and PbSO₄</td>
<td>All oxalates are insoluble except those of NH₄⁺ and Group 1A (alkali metal cations).</td>
</tr>
<tr>
<td>Chromates, CrO₄²⁻</td>
<td>Sulfides, S⁻</td>
</tr>
<tr>
<td>All chromates are soluble</td>
<td>All sulfides are insoluble except those of NH₄⁺ Group 1A (alkali metal cations), and Group 2A (Mg, Ca, and Ba) are sparingly soluble.</td>
</tr>
<tr>
<td>Pseudochromates, Cr₂O₇⁻</td>
<td></td>
</tr>
<tr>
<td>All pseudochromates are soluble.</td>
<td></td>
</tr>
<tr>
<td>Acetates, CH₃COO⁻</td>
<td></td>
</tr>
<tr>
<td>All acetates are soluble.</td>
<td></td>
</tr>
</tbody>
</table>

Table 5.1a, p.168

Table 5.1b, p.168
1. All nitrates are soluble.
2. All compounds of Group IA metals and the ammonium ion, NH$_4^+$, are soluble.
3. All chlorides are soluble except: AgCl, Hg$_2$Cl$_2$ and PbCl$_2$.
4. All sulfates are soluble except: PbSO$_4$, BaSO$_4$, and SrSO$_4$.
Solubility Rules

5. All hydroxides and sulfides are insoluble except those of the Group IA metals and the ammonium ion.
6. All carbonates and phosphates are insoluble except those of the Group IA metals and the ammonium ion.

Precipitation Reactions

The process of separating a substance from a solution as a solid.
The formation of a solid from solution.
The opposite of dissolution.

Precipitation of Barium Sulfate

$$\text{BaCl}_2(\text{aq}) + \text{Na}_2\text{SO}_4(\text{aq}) \rightarrow 2\text{NaCl}(\text{aq}) + \text{BaSO}_4(\text{s})$$

Precipitate

Sodium and chloride ions appear on both sides of the equation. They are spectator ions, which assure charge neutrality but do not take part directly in a chemical reaction.

Ionic Equations

Total Ionic Equation:
$$\text{Ag}^+ + \text{NO}_3^- + \text{Na}^+ + \text{Cl}^- \rightarrow \text{AgCl} + \text{Na}^+ + \text{NO}_3^-$$

Net Ionic Equation:
$$\text{Ag}^+ + \text{Cl}^- \rightarrow \text{AgCl}$$
Precipitation of Silver Chloride

\[ \text{AgNO}_3 + \text{NaCl} \rightarrow \text{AgCl} + \text{NaNO}_3 \]

Precipitate

Ionic Equations

Total Ionic Equation:

\[ \text{Ba}^{2+} + 2\text{Cl}^{-} + 2\text{Na}^{+} + \text{SO}_4^{2-} \rightarrow 2\text{Na}^{+} + \text{Cl}^{-} + \text{BaSO}_4(s) \]

Net Ionic Equation:

\[ \text{Ba}^{2+} + \text{SO}_4^{2-} \rightarrow \text{BaSO}_4 \]

Neutralization Reactions

- acid
  - substance that donates H\(^+\) ions to solution
  - sour-tasting substances
  - substances whose aqueous solutions are capable of turning blue litmus indicators red
  - dissolves certain metals to form salts
  - react with bases or alkalis to form salts

Neutralization Reactions

- base
  - substance that donates a OH\(^-\) ion to solution
  - hydroxides and oxides of metals
  - bitter tasting, slippery solutions
  - turn litmus blue
  - react with acids to form salts
Neutralization Reactions

- substances produced by the reaction of an acid with a base
- characterized by ionic bonds, relatively high melting points, electrical conductivity when melted or when in solution, and a crystalline structure when in the solid state

Common Acids and Bases

<table>
<thead>
<tr>
<th>Strong acids</th>
<th>Strong bases</th>
</tr>
</thead>
<tbody>
<tr>
<td>HCl</td>
<td>LiOH</td>
</tr>
<tr>
<td>HNO3</td>
<td>NaOH</td>
</tr>
<tr>
<td>H2SO4</td>
<td>Ca(OH)2</td>
</tr>
<tr>
<td>H2CO3</td>
<td>HCl</td>
</tr>
<tr>
<td>HClO3</td>
<td>HNO3</td>
</tr>
<tr>
<td>H2SO4</td>
<td>LiOH</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Weak acids</th>
<th>Weak bases</th>
</tr>
</thead>
<tbody>
<tr>
<td>H2PO4</td>
<td>NH4Cl</td>
</tr>
<tr>
<td>CH3COOH</td>
<td>HCO2Na</td>
</tr>
<tr>
<td>H2CO3</td>
<td>NH4Cl</td>
</tr>
<tr>
<td>HCN</td>
<td>HCO2Na</td>
</tr>
<tr>
<td>H2SO3</td>
<td>NH4Cl</td>
</tr>
</tbody>
</table>

*Monoatomic acids are not weak acids.
*Monoatomic bases are not weak bases.
Neutralization Reactions

acid + base → “salt” + water

HCl + NaOH → NaCl + H₂O

H₂SO₄ + 2KOH → K₂SO₄ + 2H₂O

Strong vs. Weak Acids and Bases

• strong – completely ionized
• weak – partially ionized

Ionic Equations

HCl + NaOH → NaCl + H₂O

Total Ionic Equation:
H⁺ + Cl⁻ + Na⁺ + OH⁻ → Na⁺ + Cl⁻ + H₂O

Net Ionic Equation:
H⁺ + OH⁻ → H₂O

H₂SO₄ + 2KOH → K₂SO₄ + 2H₂O

Total Ionic Equation:
2H⁺ + SO₄²⁻ + 2Na⁺ + 2OH⁻ → 2Na⁺ + 2Cl⁻ + 2H₂O

Net Ionic Equation:
2H⁺ + 2OH⁻ → 2H₂O
Gas-Forming Exchange Reaction

Reaction of calcium carbonate with an acid.

Antacid reacting with HCl

Reaction of Metal Carbonates with Acids

**CaCO₃(s) + 2HCl(aq) → CaCl₂(aq) + H₂CO₃(aq)**

**H₂CO₃(aq) → H₂O + CO₂(g)**

Total Ionic Equation:

\[ \text{CaCO}_3(s) + 2\text{H}^+ + 2\text{Cl}^- \rightarrow \text{Ca}^{2+} + 2\text{Cl}^- + \text{H}_2\text{O} + \text{CO}_2(g) \]

Net Ionic Equation:

\[ \text{CaCO}_3(s) + 2\text{H}^+ \rightarrow \text{Ca}^{2+} + \text{H}_2\text{O} + \text{CO}_2(g) \]

**Reaction of Metal Carbonates with Acids**

Alka-Seltzer

\[ \text{NaHCO}_3(aq) + \text{HCl(aq)} \rightarrow \text{NaCl(aq)} + \text{H}_2\text{O} + \text{CO}_2(g) \]

Net Ionic Equation:

\[ \text{HCO}_3^- + 2\text{H}^+ \rightarrow \text{H}_2\text{O} + \text{CO}_2(g) \]

Tums

\[ \text{CaCO}_3(s) + 2\text{HCl(aq)} \rightarrow \text{CaCl}_2(aq) + \text{H}_2\text{O} + \text{CO}_2(g) \]

Net Ionic Equation:

\[ \text{CO}_3^{2-} + 2\text{H}^+ \rightarrow \text{H}_2\text{O} + \text{CO}_2(g) \]
**Reaction of Metal Sulfites with Acids**

\[
\text{CaSO}_3(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{CaCl}_2(\text{aq}) + \text{H}_2\text{SO}_3(\text{aq})
\]
\[
\text{H}_2\text{SO}_3(\text{aq}) \rightarrow \text{H}_2\text{O} + \text{SO}_2(\text{g})
\]

Total Ionic Equation:

\[
\text{CaSO}_3(\text{s}) + 2\text{H}^+ + 2\text{Cl}^- \rightarrow \text{Ca}^{2+} + 2\text{Cl}^- + \text{H}_2\text{O} + \text{SO}_2(\text{g})
\]

Net Ionic Equation:

\[
\text{CaSO}_3(\text{s}) + 2\text{H}^+ \rightarrow \text{Ca}^{2+} + \text{H}_2\text{O} + \text{SO}_2(\text{g})
\]

**Reaction of Metal Sulfides with Acids**

\[
\text{Na}_2\text{S}(\text{aq}) + 2\text{HCl}(\text{aq}) \rightarrow 2\text{NaCl}(\text{aq}) + \text{H}_2\text{S}(\text{g})
\]

Total Ionic Equation:

\[
2\text{Na}^+ + \text{S}^{2-} + 2\text{H}^+ + 2\text{Cl}^- \rightarrow 2\text{Na}^+ + 2\text{Cl}^- + \text{H}_2\text{S}(\text{g})
\]

Net Ionic Equation:

\[
\text{S}^{2-} + 2\text{H}^+ \rightarrow \text{H}_2\text{S}(\text{g})
\]

**Oxidation-Reduction Reactions**

Oxidation – loss of electrons
Reduction – gain of electrons
Redox reaction
oxidizing agent – substance that causes oxidation
reducing agent – substance that causes reduction
SnO\textsubscript{2} loses oxygen and is reduced.

\[
\text{SnO}_2(s) + 2 \text{C}(s) \rightarrow \text{Sn}(s) + 2 \text{CO}(g)
\]

Mg combines with oxygen and is oxidized.

\[
2 \text{Mg}(s) + \text{O}_2(g) \rightarrow 2 \text{MgO}(s)
\]

Each Ag\textsuperscript{+} accepts an electron and is reduced to Ag.

\[
2 \text{Ag}^+(aq) + \text{Cu}(s) \rightarrow 2 \text{Ag}(s) + \text{Cu}^{2+}(aq)
\]

Oxidation of Copper metal by silver ion

\[
2\text{Ag}^+(aq) + \text{Cu}(s) \rightarrow 2 \text{Ag}(s) + \text{Cu}^{2+}(aq)
\]