AgI 8.3 x 10^{-17} \text{ least soluble}
AgBr 5.0 x 10^{-13}
AgCl 1.8 x 10^{-10}
Ag_2CrO_4 2:1 2.6 x 10^{-12}

Calculate the molar solubility of AgCl in pure water at 25°C. \( K_{sp} = 1.8 \times 10^{-10} \)

\[
\text{AgCl}(s) \rightleftharpoons \text{Ag}^+(aq) + \text{Cl}^-(aq)
\]

\[
\begin{array}{ccc}
\text{I} & \text{C} & \text{E} \\
0 & 0 & +S \\
+S & +S & S \\
S & S & S
\end{array}
\]

\[
K_{sp} = [\text{Ag}^+][\text{Cl}^-] = (S)(S) = 1.8 \times 10^{-10}
\]

\[
S = \sqrt{1.8 \times 10^{-10}} = 1.3 \times 10^{-5} \text{ mol/L}
\]

Molar solubility = 1.3 \times 10^{-5} \text{ mol AgCl/L sol'n. (M)}

\[
[\text{Ag}^+] = [\text{Cl}^-] = 1.3 \times 10^{-5} \text{M}
\]

Calculate the molar solubility of HgI_2 in pure water at 25°C. \( K_{sp} = 3.0 \times 10^{-29} \)

\[
\text{HgI}_2(s) \rightleftharpoons \text{Hg}^{2+}(aq) + 2 \text{I}^-(aq)
\]

\[
\begin{array}{ccc}
\text{I} & \text{C} & \text{E} \\
0 & 0 & +S \\
+2S & +S & S \\
S & 2S & 2S
\end{array}
\]
\[ K_{sp} = [Hg^{2+}][I^-]^2 = (S)(2S)^2 = 3.0 \times 10^{-29} \]

\[ 4S^3 = 3.0 \times 10^{-29} \]

\[ S = \sqrt[3]{\frac{3.0 \times 10^{-29}}{4}} = 2.0 \times 10^{-10} \]

**Molar solubility** = \(2.0 \times 10^{-10}\) mol HgI\(_2\)/litre soln.

\[ [Hg^{2+}] = 2.0 \times 10^{-10} M = S \]

\[ [I^-] = 4.0 \times 10^{-10} M = 2S \]

• **Calculate the molar solubility of HgI\(_2\) in 0.010 M NaI soln.**

\[ HgI_2(s) \rightleftharpoons Hg^{2+}(aq) + 2I^-(aq) \]

\[
\begin{array}{ccc}
I^- & 0 & 0.010 \\
\hline
C & +S & +2S \\
E & S & 0.010+2S \\
\end{array}
\]

\[ K_{sp} = 3.0 \times 10^{-29} = (S)(0.010+2S)^2 \approx (S)(0.010)^2 \]

\[ 3.0 \times 10^{-29} = S(0.010)^2 \]

\[ 2S \ll 0.010 \quad 0.010+2S \approx 0.010 \]

\[ S = \frac{3.0 \times 10^{-29}}{(0.010)^2} = 3.0 \times 10^{-25} M \]

\[ \text{stupid small} \]

\[ 
\begin{align*}
\text{HCO}_3^- \quad \text{(lose H}^+) \quad \text{its conj. base?} \quad \text{CO}_3^{2-} \\
\text{H}_2\text{CO}_3 \quad \text{(gain H}^+) \quad \text{its conj. acid?} 
\end{align*} \]
Chapter 18 & 19.1–2  
**Acids, Bases, & Acid-Base Equilibria**

Many models have been proposed over hundreds of years to describe acids and bases.

Observations about acids:
- taste sour  (DO NOT taste lab chemicals!)
- tend to react with metals to produce H₂ gas
- turn blue litmus red
- tend to react with bases to produce water and a salt

Observations about bases:
- feel slippery and often taste bitter (DO NOT touch or taste lab chemicals!)
- turn red litmus blue
- tend to react with acids to produce water and a salt

Common household acids include aspirin, the juices of citrus fruits, vinegar  
Common household bases include milk of magnesia, ammonia, lye

**Brønsted-Lowry Definition/Model of Acids and Bases**  
(early 20th C., two scientists working separately who got the same results)

- **Brønsted acid**: a substance that donates H⁺ to another substance, i.e., a proton donor.  
  All Arrhenius acids are also Brønsted acids, but not vice versa.

  \[
  \text{HCl} + \text{H}_2\text{O} \rightarrow \text{Cl}^- + \text{H}_3\text{O}^+ \\
  \text{ACID} \quad \text{H} \quad \text{BASE}
  \]

- **Brønsted base**: a substance that accepts H⁺ from another substance, i.e., a proton acceptor.  
  Arrhenius bases contain the Brønsted base OH⁻.

  \[
  \text{NH}_3 + \text{H}_2\text{O} \rightarrow \text{NH}_4^+ + \text{OH}^- \\
  \text{BASE} \quad \text{H} \quad \text{ACID}
  \]

Note that water can react as an acid or a base—this behavior is referred to as amphoterism (adjective: amphoteric).

Most acid-base equilibria we discuss will be modeled using the Brønsted definition.

e.g. \( \text{CO}_2(aq) = \text{acidic} \)  
not Brønsted acid
Conjugate Acid-Base Pairs
Consider the forward and reverse reactions:

\[
\text{NH}_3(\text{aq}) + \text{H}_2\text{O}(l) \rightleftharpoons \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq}) \quad \text{(forward)}
\]

\[
\text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightleftharpoons \text{NH}_3(\text{aq}) + \text{H}_2\text{O}(l) \quad \text{(reverse)}
\]

Members of a conjugate acid-base pair differ by a proton (H\(^+\)). e.g., A conjugate acid has one more H atom and one more positive charge (or one fewer negative charge) than the base it came from.

Acid-base reactions involve a competition between the bases for the proton (H\(^+\)): where the proton ends up depends on the relative strengths of the acids and bases, i.e., on \(K\) for the equilibrium process. (Recall “strong” vs. “weak” acids and bases in Chapter 4; we will revisit this concept in much greater depth shortly.)

Acid-Base Properties of Water

\[
2\text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{OH}^-(\text{aq})
\]

\[
K_W = [\text{H}_3\text{O}^+][\text{OH}^-]
\]

Measurements of “pure” water show that

\[
[\text{H}_3\text{O}^+]_W = [\text{OH}^-]_W = 1.0 \times 10^{-7} \text{ M at 25 °C}
\]

\[
\therefore K_W = 1.0 \times 10^{-14} \text{ at 25 °C.}
\]

In any aqueous solution,

\[
K_W = [\text{H}_3\text{O}^+]_{\text{TOT}}[\text{OH}^-]_{\text{TOT}}
\]

Auto-ionization of water is already always happening in water. When an acid or base is added, another equilibrium joins in. Then there are two or more equilibria!

Since the product \([\text{H}_3\text{O}^+]_{\text{TOT}}[\text{OH}^-]_{\text{TOT}}\) is constant (= \(K_W\)), the larger \([\text{H}_3\text{O}^+]\) is, the smaller \([\text{OH}^-]\) is (and vice versa). In general:

<table>
<thead>
<tr>
<th>Condition</th>
<th>Solution Property</th>
</tr>
</thead>
<tbody>
<tr>
<td>([\text{H}_3\text{O}^+] = [\text{OH}^-])</td>
<td>NEUTRAL</td>
</tr>
<tr>
<td>([\text{H}_3\text{O}^+] &gt; [\text{OH}^-])</td>
<td>ACIDIC</td>
</tr>
<tr>
<td>([\text{H}_3\text{O}^+] &lt; [\text{OH}^-])</td>
<td>BASIC</td>
</tr>
</tbody>
</table>

**PROTIP**
You may know of a definition of “neutral” that involves the number 7. That is correct only at 25 °C. The definition above is general to all temperatures.