1. Define K\(_sp\) for the dissolution of the following salts in water. If necessary, write a balanced chemical equation for the dissolution first.

   a. RbI, \(\text{RbI(s) } \rightleftharpoons \text{Rb}^+(aq) + \text{I}^-(aq)\), \(K_{sp} = [\text{Rb}^+] \cdot [\text{I}^-]\)
   
   b. Ca(NO\(_3\))\(_2\), \(K_{sp} = [\text{Ca}^{2+}] \cdot [\text{NO}_3^-]^2\)
   
   c. \(K_3\text{PO}_4\), \(K_{sp} = [\text{K}^+]^3 \cdot [\text{PO}_4^{3-}]\)
   
   d. SrS, \(K_{sp} = [\text{Sr}^{2+}] \cdot [\text{S}^2^-]\)
   
   e. \(\text{Fe}_2(\text{SO}_4)_3\), \(K_{sp} = [\text{Fe}^{3+}]^2 \cdot [\text{SO}_4^{2-}]^3\)
   
   f. \(\text{K}_3\text{Fe(C}_2\text{O}_4)_3\), \(K_{sp} = [\text{K}^+]^3 \cdot [\text{Fe}^{3+}] \cdot [\text{C}_2\text{O}_4^{2-}]^3\)

2. Consider each of the salts below. Express each salt's molar solubility (we'll call it x) in terms of \(K_{sp}\). It might be useful to first write a balanced equation for each salt's dissolution and complete a RICE diagram.

   a. \(\text{Cu}_3(\text{PO}_4)_2\), molar solubility = \(x = (K_{sp}/108)^{1/5}\)

   \[
   \begin{array}{|c|c|c|c|c|}
   \hline
   \text{R} & \text{Cu}_3(\text{PO}_4)_2(s) & \rightleftharpoons & 3 \text{Cu}^{2+}(aq) & + & 2 \text{PO}_4^{3-}(aq) \\
   \hline
   \text{I} & \sim & 0 & 0 \\
   \hline
   \text{C} & \sim & + 3x & +2x \\
   \hline
   \text{E} & \sim & 3x & 2x \\
   \hline
   \end{array}
   \]

   \[
   K_{sp} = [\text{Cu}^{2+}]^3 \cdot [\text{PO}_4^{3-}]^2 = (3x)^3 \cdot (2x)^2 = 108x^5
   
   K_{sp} = 108x^5
   
   x = (K_{sp}/108)^{1/5}
   
   b. MgSe, \(x = (K_{sp})^{1/2}\)
   
   c. \(\text{Li}_3\text{PO}_4\), \(x = (K_{sp}/27)^{1/4}\)
   
   d. \(\text{K}_3\text{Fe(C}_2\text{O}_4)_3\), \(x = (K_{sp}/729)^{1/7}\)

3. Estimate the actual molar solubilities of the following salts in water based on their \(K_{sp}\) values.

   a. Barium Sulfate, \(\text{BaSO}_4\), \(K_{sp} = 1.08 \times 10^{-10}\), molar solubility = \(x = 10^{-5}\) M
   
   b. Cadmium Phosphate, \(\text{Cd}_3(\text{PO}_4)_2\), \(K_{sp} = 2.53 \times 10^{-33}\), \(x = 10^{-6}\) M
   
   c. Lithium Carbonate, \(\text{Li}_2\text{CO}_3\), \(K_{sp} = 1.73 \times 10^{-3}\), \(x = 10^{-1}\) M
   
   d. Magnesium Ammonium Phosphate, \(\text{MgNH}_4\text{PO}_4\), \(K_{sp} = 2.5 \times 10^{-13}\), \(x = 10^{-4}\) M

4. Estimate the actual molar solubilities of the following salts in the following solutions based on the provided concentrations and \(K_{sp}\) values. It might be useful to first write a balanced equation for each salt's dissolution and complete a RICE diagram.

   a. Mercuric Bromide, \(\text{HgBr}_2\), \(K_{sp} = 8 \times 10^{-20}\), in \(2\) M \(\text{Hg(NO}_3)_2\), molar solubility = \(x = 10^{-10}\) M

   \[
   \begin{array}{|c|c|c|c|}
   \hline
   \text{R} & \text{HgBr}_2(s) & \rightleftharpoons & \text{Hg}^{2+}(aq) & + & \text{Br}^-(aq) \\
   \hline
   \text{I} & \sim & 2 & 0 \\
   \hline
   \text{C} & \sim & + x & +2x \\
   \hline
   \text{E} & \sim & 2 + 3x & 2x \\
   \hline
   \end{array}
   \]

   \[
   K_{sp} = [\text{Hg}^{2+}] \cdot [\text{Br}^-]^2
   
   8 \times 10^{-20} = (2 + 3x) \cdot (2x)^2
   
   For the term \((2 + 3x)\), it is safe to assume that \(2 + 3x = 2\), and the equation reduces to
   
   \[
   8 \times 10^{-20} = (2) \cdot (2x)^2
   
   x = (8 \times 10^{-20}/(2x)^2)^{1/2}
   
   b. Silver Chloride, \(\text{AgCl}\), \(K_{sp} = 1.56 \times 10^{-10}\), in \(15\) M \(\text{KCl}\), \(x = 10^{-11}\) M
   
   c. Barium Iodate, \(\text{Ba(IO}_3)_2\), \(K_{sp} = 6.5 \times 10^{-10}\), in \(2.5\) M \(\text{KIO}_3\), \(x = 10^{-10}\) M
5. Match the $K_w$ values on the left with their corresponding pH values on the right. Assume you have a sample of completely pure water.

<table>
<thead>
<tr>
<th>$K_w$ (e-14)</th>
<th>pH</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.114</td>
<td>7.08</td>
</tr>
<tr>
<td>0.293</td>
<td>7.27</td>
</tr>
<tr>
<td>0.681</td>
<td>6.14</td>
</tr>
<tr>
<td>1.008</td>
<td>6.92</td>
</tr>
<tr>
<td>1.471</td>
<td>7.47</td>
</tr>
<tr>
<td>2.916</td>
<td>6.63</td>
</tr>
<tr>
<td>5.476</td>
<td>7.00</td>
</tr>
<tr>
<td>51.3</td>
<td>6.77</td>
</tr>
</tbody>
</table>

6. Answer the following questions concerning the autoprotolysis of water;
   a. Is the autoprotolysis of water endothermic or exothermic?
     endothermic
   b. What would be a simple experiment to verify this?
     Measuring the pH of a sample of pure water at different temperatures - pH will be inversely proportional to temperature if autoprotolysis is endothermic.
   c. What would be a simple way to calculate $\Delta H_{\text{autoprotolysis}}$?
     Similar to above, measuring pH at a range of temperatures would enable us to compute $K_w$ at those temperatures and we could then use the van't Hoff equation.

7. List the 7 strong acids from memory.
   Hydrochloric (HCl), Hydrobromic (HBr), Hydroiodic (HI), Sulfuric (H$_2$SO$_4$), Nitric (HNO$_3$), Chloric (HClO$_3$) and Perchloric (HClO$_4$)

8. List the 8 strong bases from memory.
   Lithium Hydroxide (LiOH), Sodium Hydroxide (NaOH), Potassium Hydroxide (KOH), Rubidium Hydroxide (RbOH), Cesium Hydroxide (CsOH), Calcium Hydroxide [Ca(OH)$_2$], Strontium Hydroxide [Sr(OH)$_2$], Barium Hydroxide [Ba(OH)$_2$]

9. List the 14 spectator ions from memory. The answers to questions 7 and 8 are a really good starting point for this problem.
   Chloride (Cl$^-$), Bromide (Br$^-$), Iodide (I$^-$), Nitrate (NO$_3^-$), Chlorate (ClO$_3^-$), Perchlorate (ClO$_4^-$), Lithium ion (Li$^+$), Sodium ion (Na$^+$), Potassium ion (K$^+$), Rubidium ion (Rb$^+$), Cesium ion (Ce$^+$), Calcium ion (Ca$^{2+}$), Strontium ion (Sr$^{2+}$), Barium ion (Ba$^{2+}$)

10. Decide whether each of the species below is a weak acid or weak base. Note that it is possible to know this based on a chemical's name, and generally possible based on its formula.
    a. pyridinium, weak acid
    b. oxalate, weak base
    c. HIO$_3$, weak acid
    d. NH$_3$, weak base
11. Complete the following table: (Hint: \(-\log 0.4 = 0.4\), this is a good and easy reference point to remember for the log function.)

<table>
<thead>
<tr>
<th></th>
<th>([H^+]) (M)</th>
<th>pH</th>
<th>([OH^-]) (M)</th>
<th>pOH</th>
</tr>
</thead>
<tbody>
<tr>
<td>Solution A</td>
<td>0.4</td>
<td>0.4</td>
<td>2.5 \times 10^{-14}</td>
<td>13.6</td>
</tr>
<tr>
<td>Solution B</td>
<td>1</td>
<td>0</td>
<td>10^{-14}</td>
<td>14</td>
</tr>
<tr>
<td>Solution C</td>
<td>(10^{-13})</td>
<td>13</td>
<td>0.1</td>
<td>1</td>
</tr>
<tr>
<td>Solution D</td>
<td>0.01</td>
<td>2</td>
<td>10^{-12}</td>
<td>12</td>
</tr>
<tr>
<td>Solution E</td>
<td>(10^{-15})</td>
<td>15</td>
<td>10</td>
<td>-1</td>
</tr>
<tr>
<td>Solution F</td>
<td>(10^{-11})</td>
<td>11</td>
<td>0.001</td>
<td>3</td>
</tr>
<tr>
<td>Solution G</td>
<td>(10^{-5})</td>
<td>5</td>
<td>10^{-9}</td>
<td>9</td>
</tr>
<tr>
<td>Solution H</td>
<td>(2.5 \times 10^{-14})</td>
<td>13.6</td>
<td>0.4</td>
<td>0.4</td>
</tr>
<tr>
<td>Solution I</td>
<td>(10^{-7})</td>
<td>7</td>
<td>10^{-7}</td>
<td>7</td>
</tr>
<tr>
<td>Solution J</td>
<td>(10^{-9})</td>
<td>9</td>
<td>10^{-5}</td>
<td>5</td>
</tr>
</tbody>
</table>

12. What would be the pH of the following solutions?
   a. 0.01 M HClO₄, for a strong acid \([H^+] = C_a\), \(-\log[H^+] = pH = 2\)
   b. 0.05 M Ba(OH)₂, note that some strong bases yield 2 \(OH^-\), pH = 13
   c. 10 M HNO₃, \(pH = -1\)
   d. LiOH, \(pH = 15\)

13. What would be the pOH pf the following solutions?
   a. 0.1 M RbOH, for a strong base \([OH^-] = C_b\), \(-\log[OH^-] = pOH = 1\)
   b. 0.5 M Sr(OH)₂, note that some strong bases yield 2 \(OH^-\), pOH = 0
   c. 0.001 M HClO₃, pOH = 11
   d. 0.4 M HI, pOH = 13.6

14. What would be the pH of the following solutions? You may approximate if necessary; you should not need a calculator.
   a. 0.25 M HNO₂, \(K_a = 4.0 \times 10^{-4}\), for a weak acid \([H^+] = (K_a C_a)^{1/2}\), pH = 2
   b. 5.55 M NH₃, \(K_b = 1.8 \times 10^{-5}\), for a weak base \([OH^-] = (K_b C_b)^{1/2}\), pH = 12
   c. 0.0125 M ascorbic acid, \(K_a = 7.9 \times 10^{-5}\), pH = 3
   d. 0.0135 M trimethylamine, \(K_b = 7.4 \times 10^{-5}\), pH = 11
   e. 0.3 M HOCl, \(K_a = 3.5 \times 10^{-8}\), pH = 4

15. Consider each of the acids and bases below. Write the formula or name for each species' conjugate and calculate the \(K_a\) or \(K_b\) for that conjugate. Approximate if necessary.
   a. ammonium, \(K_a = 5.55 \times 10^{-10}\), ammonia, \(K_b = 1.80 \times 10^{-5}\)
   b. OCl⁻, \(K_b = 2.5 \times 10^{-7}\), HOCl, \(K_a = 4.0 \times 10^{-8}\)
   c. pyridine, \(K_b = 1.6 \times 10^{-9}\), pyridinium, \(K_a = 6.0 \times 10^{-6}\)
   d. HCN, \(K_a = 4.0 \times 10^{-10}\), CN⁻, \(K_b = 2.5 \times 10^{-5}\)