Solutions Worksheet - The Solution Process

1. Define the following:
   solution _____________________________
   solvent _____________________________
   solute _____________________________

   Give an example of a common solution found in everyday life in each of the following phases.
   gas: _________________________
   liquid: _______________________
   solid:  _______________________

2. Although solutions can occur for any solvent and in any phase of matter, the most commonly
   encountered solutions are aqueous, i.e., in water. Therefore, we will consider the solution process
   more closely for aqueous solutions.

a. In the beakers shown, use symbols to represent the appropriate compounds.
   i. pure water, □
   ii. pure ethanol, ○
   iii. a solution of ethanol in water, approximately 10% ethanol, 90% water

3. During this process, significant changes in the Intermolecular Forces of Attraction occur. List
   the IMF’s present in each of the pure compounds.
   i. ___________________________________
   ii. ___________________________________

4. In the solution, what IMF’s are present. List what compounds are interacting and how.
   iii. ___________________________________
       ___________________________________
5. One way to view the changes in energy associated with formation of solutions is to break the solution process down into a series of steps. The various steps are listed below. For our example of ethanol in water, circle the appropriate enthalpy change. Base your answer on the IMF’s present and how these will be affected.

i. Separating some solvent molecules to make space available for the solute. 
\[ \Delta H_1 \] is **exothermic, endothermic or neither**. 
Explain your answer. ________________________

ii. Separating all solute molecules from each other. 
\[ \Delta H_2 \] is **exothermic, endothermic or neither** 
Explain your answer. ________________________

iii. Placing solute molecules in the available spaces in the solvent after step i. 
\[ \Delta H_3 \] is **exothermic, endothermic or neither** 
Explain your answer. ________________________

The overall enthalpy change is the sum of the enthalpy changes for these three steps, 
\[ \Delta H_{\text{soln}} = \Delta H_1 + \Delta H_2 + \Delta H_3 \]
While each combination of solute and solvent will have a different value for \( \Delta H_{\text{soln}} \), as a general trend, a solution will form if \( \Delta H_{\text{soln}} \) is exothermic and will not form if \( \Delta H_{\text{soln}} \) is highly endothermic. Since \( \Delta H_1 \) and \( \Delta H_2 \) are ALWAYS endothermic, the magnitude of \( \Delta H_3 \) is critical in determining solubility. If the IMF’s formed between solute and solvent are of comparable magnitude to those broken, the solute will usually be soluble in that solvent. Exactly how much solute will dissolve before the solution becomes saturated is also determined by these relative strengths of IMF’s. This type of analysis is often summarized in the common statement: 
"Like dissolves like."

Based upon the ideas illustrated above, complete the following tables of solubilities.

<table>
<thead>
<tr>
<th>Solvent</th>
<th>Compound</th>
<th>Solute</th>
<th>Soluble or Insoluble</th>
</tr>
</thead>
<tbody>
<tr>
<td>compound</td>
<td>polar or nonpolar?</td>
<td>compound</td>
<td>Ionic? Molecular? (if yes, polar or nonpolar)?</td>
</tr>
<tr>
<td>water</td>
<td>Polar</td>
<td>KCl</td>
<td>ionic</td>
</tr>
<tr>
<td>water</td>
<td></td>
<td>NH₄NO₃</td>
<td></td>
</tr>
<tr>
<td>water</td>
<td></td>
<td>glucose</td>
<td></td>
</tr>
<tr>
<td>water</td>
<td></td>
<td>C₁₀H₈</td>
<td></td>
</tr>
<tr>
<td>hexane, C₆H₁₄</td>
<td></td>
<td>water</td>
<td></td>
</tr>
<tr>
<td>hexane</td>
<td></td>
<td>C₁₀H₈</td>
<td></td>
</tr>
<tr>
<td>gasoline</td>
<td></td>
<td>sucrose (a sugar)</td>
<td></td>
</tr>
</tbody>
</table>

* the structure of glucose is
6. Is the dissolving of CaCl₂(s) into water endothermic or exothermic? (i.e., is ΔHsoln positive or negative?)

Is the dissolving of NH₄Cl(s) into water endothermic or exothermic? (i.e., is ΔHsoln positive or negative?)

Having a negative ΔHsoln is one factor that favors formation of solutions. However, many solutions are known to form spontaneously when ΔHsoln is zero or even positive. Another factor must also be involved in determining whether a solution will form. That factor is disorder or entropy.

Examine your three pictures on page 1 of this handout. Is disorder INCREASING OR DECREASING as the ethanol dissolves in the water? Explain your answer.

Does entropy increase or decrease when a solid lattice breaks apart into ions?

Does entropy increase or decrease when solutes are mixed into solvents?

Ions with high charges (+2, +3) strongly attract a shell of hydrated water molecules. Would locking a large number of water molecules in place this way increase or decrease the entropy of the system?
Temperature, Pressure Effects on Solubility

**Pressure Effects on Solubility:**

<table>
<thead>
<tr>
<th>Solid and Liquid Solutes</th>
<th>Gaseous solutes (picture at left)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Little effect of pressure on solubility. Why?</td>
<td>Solubility depends on partial pressure of gas</td>
</tr>
<tr>
<td></td>
<td>( S_{\text{gas}} = kP_{\text{gas}} ) (Henry’s Law)</td>
</tr>
</tbody>
</table>

**Problem:** Calculate the concentration of \( \text{CO}_2 \) in a soft drink that is bottled with a partial pressure of \( \text{CO}_2 \) of 4.0 atm over the liquid at 25°C. The Henry’s Law constant for \( \text{CO}_2 \) in water is \( k_{\text{CO}_2} = 3.1 \times 10^{-2} \text{ M/atm} \)

Temperature Effects on Solubility

Is a solution that contains 50 g KCl and 250 g of water at 25°C saturated or unsaturated? (support your choice by showing your work)

If the solution is cooled to close to 0°C, will solid KCl precipitate out? (Do the calculations required to support your answer)