Heats of Solution

Description: A qualitative comparison of $\Delta H_{\text{soln}}$ for LiCl and KCl is demonstrated by dissolving LiCl and KCl in water separately and observing the temperature change.

Materials:

- LiCl
- KCl
- 2 beakers/glass stirring rods
- Deionized water

Procedure:

Have two volunteers perform this demonstration.

In one beaker, dissolve a spatula full of LiCl in 50 mL of water. In the other beaker, dissolve a spatula full of KCl in 50 mL of water. Have students hold the beakers and observe the change in temperature. The beaker containing KCl will become noticeably colder while the beaker containing LiCl will become noticeably warmer. Alternatively, a mole of each salt can be dissolved in separate beakers and the temperature measure with a thermometer. *Dissolving one mole of LiCl in water may result in temperature increase close to 60°C.

Discussion: This demonstration illustrates that salvation can be an exothermic or an endothermic process, although in many instances, the process is endothermic. For the following process:

$$\text{LiCl (s) + H}_2\text{O (l) \rightarrow Li}^{+} (\text{aq}) + \text{Cl}^{-} (\text{aq});$$

the $\Delta H_{\text{soln}}^{\circ} = -37.1$ kJ/mol as described by Shakhashiri. Other Li salts produce similar results (with the exception of LiF) although most evolve less heat than LiCl. Lithium salts tend to have higher energies of solution because of the large value for the hydration energy of Li$^+$. The following heats of solution for other chloride salts show this to be true: $\text{NH}_4\text{Cl} (+15.2 \text{ kJ/mol})$, $\text{NaCl} (+3.9 \text{ kJ/mol})$, $\text{KCl} (+17.24 \text{ kJ/mol})$, $\text{RbCl} (+16.7 \text{ kJ/mol})$. 
Safety: The dissolution of LiCl can create enough heat to cause burns

Disposal: Solutions can be flushed down the drain with plenty of water.

References:
