Enthalpy of Solution of Ionic Compounds

Introduction

When a solid dissolves in water, the process always has an energy change associated with it; because the positive and negative ions of the salt *interact* with water molecules. Examples of dissolving of ionic compounds exist for both endothermic and exothermic heats of solution.

However, the dissolving process itself is really a two-step process. The first step, that of breaking down the solid crystal, is endothermic while the second step, that of *hydrating* (that is to say – the surrounding of an ion with a layer of water molecules) the individual particles released into the solvent, is exothermic.

For an ionic compound, ionic bonds are broken, i.e. *crystal lattice*, while *ion-dipole* forces are formed. Ion-dipole forces are formed when the slightly positive hydrogen atoms (of the highly polar water molecules) are attracted to the negative ions, i.e. the anions of the ionic compound in order to maximize electrostatic attractive forces. Whilst the slightly negative oxygen atoms of the polar water molecules are attracted to the positive ions, the cations of the ionic compound – the *hydration* of the ions.

The overall heat of solution depends on the relative amounts of energy involved in these two individual steps – the lattice energy and the hydration energy. In this experiment, you will determine the heats of solution for the dissolving of some ionic compounds in water, using a simple styrofoam cup as a calorimeter.

Prelab Assignment

1. For each compound used in this experiment:
   a) Calculate the number of moles present in a 3.00g sample of each of the ionic compounds to be used in this experiment, (see Reagents).
   b) Write the chemical equation for each compound's dissociation in water.
2. Describe two factors that affect the strength of an ionic bond in a compound.
3. a) Define crystal lattice energy and hydration energy.
   b) Which of the properties in part (a) reduces the solubility of an ionic compound? Which property, when increased, increases the solubility?

Objectives

To find the molar heat of solution for a group of ionic compounds.
To relate the heat of solution of an ionic compound to the compound's crystal lattice energy and hydration energy.

Apparatus

balance
nested Styrofoam cup calorimeter
thermometer
safety goggles
50 cm^3 graduated cylinder

Reagents

potassium nitrate, KNO_3
sodium acetate trihydrate, $CH_3COO.Na..3H_2O$
ammonium chloride, NH_4Cl
sodium hydroxide, NaOH

Safety

1. Sodium hydroxide is extremely corrosive. Eye protection must be worn at all times.
2. Ammonium chloride is a very fine powder that is potentially dangerous if it comes in contact with your eyes.

Procedure

Dissolve ~ 3.00g of each ionic compound, (*note the mass of the ionic compound used)*, in a nested Styrofoam cup calorimeter using a precisely measured volume of distilled water, (25 cm^3 is a suitable amount). Note the initial temperature of the water and note the maximum temperature change. Rinse the calorimeter well and repeat for the other compounds.

Data Collection

Make a suitable table for the collection of all your data.

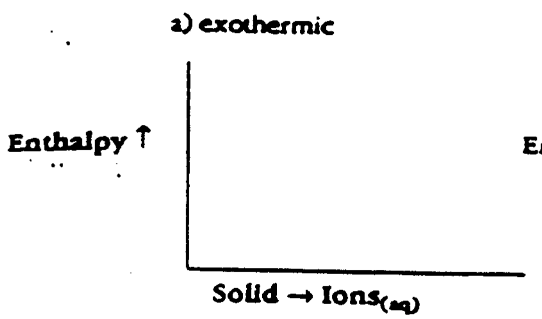
Data Processing and Presentation

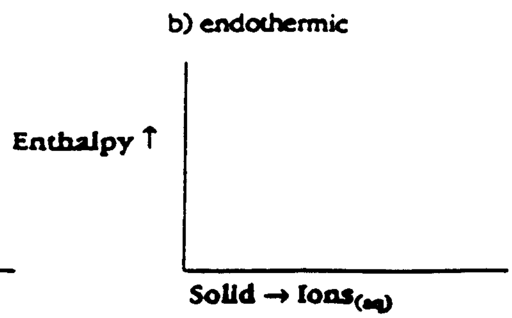
1. From your data, calculate the following for each part of the experiment.
   a) the temperature change of the water.
   b) the mass of the water.
   c) the quantity of heat absorbed (or given off) by the water during the dissolving, given that the specific heat of water is $4.18\ J.g^{-1}.^{O}C^{-1}$
   d) the number of moles of solid used (see Prelab Assignment).
   e) the quantity of heat involved per mole of solid dissolved. This is called the molar heat of solution.
2. Write a thermochemical equation for the dissolving process for each solid, (i.e. include the heat term in each equation).
3. Find the accepted values for the molar heats of solution for these solids from your Data Book, and calculate the percent error of your experimental values.

Developing the Idea

1. a) In your notebook, draw a table to summarize your data and calculations.
   b) For each compound, state whether the dissolving was an endothermic or exothermic process.
2. When a substance dissolves its heat content, or enthalpy, will either increase or decrease. This change in enthalpy is observed as a change in temperature of the water in the calorimeter.

Sketch a graph showing the change in enthalpy for the dissolving of an ionic compound which is:





3. Consider the fact that dissolving is actually a two-step process. How does the nature of these individual steps combine to determine whether the overall process will be endothermic or exothermic?
4. For an ionic solid to melt, external energy must be applied to overcome the attractive forces holding its ions together. For example, sodium hydroxide must be heated to temperatures exceeding 318 ºC for it to melt. Despite such a high melting point, sodium hydroxide readily dissolves and dissociates into its ions in water at room temperature.
   a) What is the source of the energy required to separate the ions from the solid sodium hydroxide?
   b) Why does the resulting sodium hydroxide solution feel hot?
5. Ammonium chloride is also a soluble despite having a melting point of 320 ºC. Why does an ammonium chloride solution feel cold to the touch?

6. Cold packs and hot packs are available for a variety of medical uses. From the reagents used in this lab state, which of these chemicals would be suitable for which type of pack.

7. Magnesium oxide, MgO, and magnesium chloride, $MgCl_2$, are very similar, white ionic solids with the following properties:

Compound	Melting Point (^{O}C)	Solubility
MgO	**2800**	**Insoluble**
$MgCl_2$	**1412**	**very Soluble**

a) Give the formula of the ions of each compound.
b) Account for the drastic difference in physical properties.

8. Consider the following data from Group I chlorides. These compounds are similar in structure, each being formed from a +1 cation and the –1 chloride ion. However, they are quite different with respect to lattice energy, hydration energy and enthalpy of solution. Explain these differences, the trend in these differences, and the relationship between the three energies listed:

Compound	Lattice Energy * (kJ mol^{-1})	Hydration Energy (kJ mol^{-1})	Enthalpy of Solution (kJ mol^{-1})
LiCl	833	- 833	- 50
NaCl	766	- 770	- 4
KCl	690	- 686	4

* These values are the true lattice energies given opposite signs. True lattice energies have negative values, since they relate to the exothermic formation of crystalline lattices from their constituent, gaseous ions.

9. The energies of hydrations of individual ions add up to the hydration energy of a salt. Which would be expected to have the larger hydration energy, Al^{+3} or Li^{+}? Explain

10. The value of ΔH_{soln} for the formation of an acetone – water solution is negative. Explain this in general terms that discuss intermolecular forces of attraction.

Conclusion and Evaluation

Make a valid conclusion, based on the correct interpretation of the results, with an explanation relating to the process that will determine whether a reaction will be endothermic or exothermic.