

# Activity 11 - Energy Changes of Solution Formation

## Goals

- Measure and record energy changes in the dissolution of a salt in water.
- Describe and explain energy changes in the dissolution process in terms of changes in entropy and enthalpy.
- Observe the change in freezing point of water upon addition of a solute.

**Pre-Lab Lecture Questions.** *Answer these questions on a separate sheet using complete sentences.*

1. What is the relationship between temperature and heat? Explain your answer.
2. Chemical and physical changes involve changes in energy. Give an example of a change that requires an input of energy and an example of a change that produces energy.
3. After reading the introduction section, define the properties represented by the symbols of T, G, H and S. What does “ $\Delta$ ” mean?
4. What does it mean if a reaction is spontaneous? How does spontaneity relate to Gibbs free energy?
5. Which of the following changes include increases in entropy?
  - a.  $\text{ATP} + \text{H}_2\text{O} \rightarrow \text{ADP} + \text{phosphate ion} + \text{hydronium ion}$
  - b. One starch molecule  $\rightarrow$  101,000 glucose molecules
  - c.  $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$
6. Is the making of ice cream increasing or decreasing entropy? If the ice cream ingredients mixed together is the chemical system, is it an exothermic or endothermic process to convert them to ice cream?

## Introduction

Why are some changes spontaneous while others are not? Spontaneous reactions occur because this change will lead to an increase in the entropy (S) of the universe. A helpful picture is to think of entropy as a measure of disorder or chaos. In the context of any process under consideration, the universe can be separated into two components. One is the **system**, the portion of the universe that is the focus of attention (such as a beaker containing 100 mL of water and 2.0 g of NaCl). The other is the **thermal surroundings**, which is simply a source or sink of thermal energy (heat). **Note that the net entropy change ( $\Delta S$ ) in any process is the sum of the entropy changes in the system and the surroundings.**

There are two aspects to consider when assessing whether or not a process under consideration is spontaneous. One has to do with the organization of the materials under consideration. If the materials have become more “disorganized”, **entropy** has increased and this aspect of the process favors spontaneity. Examples are a solid or liquid dissolving in a solvent, or a chemical reaction in which the number of product molecules is larger than the number of reactant molecules. In both these cases the chemical species under consideration have become more disorganized. (Another way to think of this is that the number of ways to arrange the components of the matter under consideration has increased without changing the appearance of the collective. For an apt example, the number of ways to arrange the components of an aqueous solution of NaCl without changing the appearance of the solution in any way far exceed the number of ways to arrange the components of solid NaCl and liquid water as separated phases.)

The other aspect of a process that has an effect on spontaneity is the transfer of thermal energy (heat) into or out of the system. When this transfer occurs under a condition of constant pressure, the quantity of heat transferred is called the **enthalpy change ( $\Delta H$ )** associated with the process. Recall that the thermal surroundings are either a source or sink of heat. The entropy of the surroundings always increases when heat is transferred to it from the system (such processes are termed **exothermic**;  $\Delta H$  is negative for these), and always decreases when the transfer of heat goes in the opposite direction (such processes are termed **endothermic**;  $\Delta H$  is positive for these). This means that if a spontaneous reaction occurs in the system that is exothermic, the entropy increases in both the system and the surroundings. However when such a reaction is endothermic, whether or not it is spontaneous depends on the relative magnitudes of the entropy increase in

the system and the entropy decrease in the surroundings, with the greater magnitude determining the outcome. Thus it is possible for an endothermic reaction to be spontaneous, but only if the increase in the entropy of the system is greater than the decrease in the entropy of the surroundings.

Any process for which  $\Delta S$  is positive and  $\Delta H$  is negative will lead to a spontaneous process, since the entropy of both the system and the surroundings has increased. However if only one of these driving forces is present, the spontaneity of the process depends on the relative magnitudes of each change. An American mathematician named J.W. Gibbs developed a relationship that considers both changes in the entropy of the system and that of the surroundings, but uses only quantities related to the system. It makes use of a term known as the *Gibbs free energy* ( $G$ ). The relationship developed by Gibbs for a chemical process is as follows:

$$\Delta G = \Delta H - T\Delta S$$

In the above expression  $\Delta G = G_{\text{prod}} - G_{\text{react}}$ ,  $\Delta H = H_{\text{prod}} - H_{\text{react}}$ , and  $\Delta S = S_{\text{prod}} - S_{\text{react}}$ . A negative Gibbs free energy change ( $\Delta G < 0$ ) means that the products are lower in overall free energy than the reactants. A reaction with a negative  $\Delta G$  is called *exergonic* and is spontaneous. Both terms in the above relationship have units of energy and the overall change in  $G$  is considered the amount of energy available to do work (hence the term “free energy”).

Many ionic compounds (salts) are soluble in water. For these the dissolution process is spontaneous and therefore  $\Delta G$  is negative. To observe the contributions of enthalpy and entropy to  $\Delta G$ , different salts will be dissolved in water and the final temperature of the solution will be measured. The typical ice/table salt mixture used in ice cream freezers will also be investigated.

### Experimental Concerns

Use only small amount of quantities as described.

Use comparable amounts of salts.

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### Safety

*Wear safety glasses at all times!*

Act in accordance with the laboratory safety rules of Cabrillo College.

Avoid contact with all chemical reagents and dispose of reactions using appropriate waste containers.

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### Materials

Rock salt, sodium chloride

ammonium chloride

Calcium chloride

crushed ice

Equipment: Lab top reaction surface, plastic spoon, 4 (1 oz) cups, alcohol thermometer

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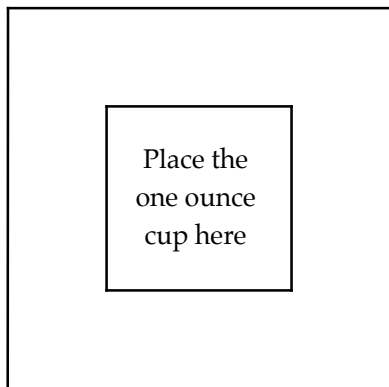
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**Reaction Template:** Insert this page into the labtop.

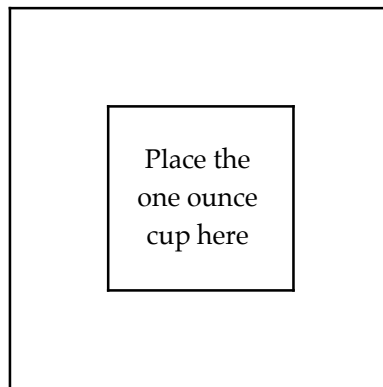
Arrange four plastic cups according to the diagram. Do each experiment one at a time in the following way:

1. Transfer two level spoonfuls of deionized water into the cup. (*Use crushed ice in part (d).*)
  2. Record the initial temperature of the water. Make sure you read the temperature to the limit of the device. (Hint: The last decimal place in the measurement is an estimated one.)
  3. Add one level spoonful of the indicated solid chemical.
  4. Stir gently and measure the highest or lowest temperature reached. Record this temperature as the final temperature.
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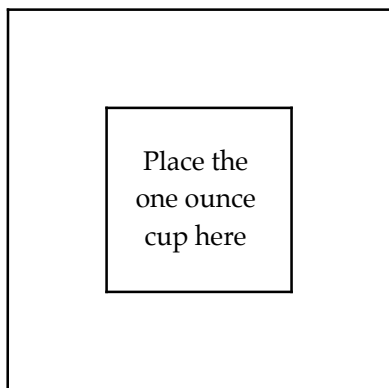
a.  $\text{NaCl}_{(s)}$



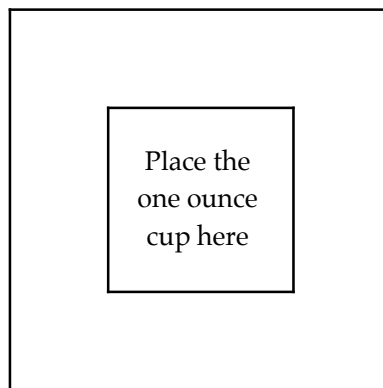
b.  $\text{NH}_4\text{Cl}_{(s)}$



c.  $\text{CaCl}_2_{(s)}$



d.  $\text{NaCl}_{(s)}$



### Cleaning up

**Avoid contamination by cleaning up in a way that protects you and your environment. Pour the contents of the cups down the drain, rinsing them with plenty of water. Rinse and dry the thermometer**

**before returning. Dispose of all paper towels in the waste bin. Wash your hands with soap and water before leaving.**

## Activity 11 - Energy Changes in Solution Formation

Name \_\_\_\_\_

Section \_\_\_\_\_ Date \_\_\_\_\_

**Experimental Data and Calculations:** Record your initial and final temperatures and then determine the overall change in temperature. Include the appropriate sign for the change in temperature.

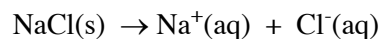
Experimental Mixture	Final Temperature (°C)	Initial Temperature (°C)	$\Delta T = T_f - T_i$ (°C)
a. NaCl(s) + H <sub>2</sub> O(l)			
b. NH <sub>4</sub> Cl(s) + H <sub>2</sub> O(l)			
c. CaCl <sub>2</sub> (s) + H <sub>2</sub> O(l)			
d. NaCl(s) + H <sub>2</sub> O(l)			

### Questions and Problems

Use your results in the table above to answer the following questions.

1. Upon mixing, which mixture(s) cooled off?
2. Which mixture(s) warmed up?
3. Which mixture had the least temperature change?

When sodium chloride dissolves in water, the ions dissociate or come apart in solution:



4. Write ionic equations like the one above to show how ammonium chloride **and** calcium chloride dissociate when they dissolve in water.

5. Is the dissolving of ammonium chloride an exothermic or endothermic process? Explain. Rewrite the ionic equation from Question 4, but this time include "heat:" on the appropriate side of the equation.
  
6. Is the dissolving of ammonium chloride a spontaneous process? Does it occur with an increase or decrease in entropy? (Consider both the system and the surroundings.) What is the driving force for spontaneous change here?
  
7. Is the dissolving of calcium chloride an exothermic or endothermic process? Explain. Rewrite the ionic equation from Question 4, showing heat as a reactant or product.
  
8. Is the dissolving of calcium chloride a spontaneous process? Does it occur with an increase or decrease in entropy? (Consider both the system and the surroundings.) What is the driving force here?
  
9. Is the dissolving of sodium chloride an exothermic or endothermic process? Explain. Is it spontaneous? What is the driving force here?
  
10. Describe your observations when you mix sodium chloride with liquid water on the one hand, and solid water (ice) on the other. Any differences between the two results?
  
11. The principle behind an ice cream freezer is that when salt is added to the ice, the melting ice absorbs heat from the ice cream, and the ice cream freezes. What drives the spontaneous melting of ice when it is mixed with salt?