(H) Chemistry **Name:**
Thermochemistry **Blk:** **Date:**
Heat of Solution Lab **Lab#__________**

Introduction:

When salts are dissolved in water, there is often a temperature change associated with the process, sometimes quite noticeable. Some salts dissolve releasing heat in the process. Others dissolve while absorbing heat. Recall that processes that proceed with a release of heat energy are 'exothermic processes', while those that absorb heat are called 'endothermic processes'. When energy is released to the surroundings in an exothermic process, the heat of solution would be given a negative value because the energy of the system is decreasing.

$$NaOH(s) \rightarrow Na^{+}(aq) + OH^{-}(aq) \qquad \Delta H_{soln} = -44.51 \ kJ$$

When energy is absorbed from the surroundings in an endothermic process, the heat of solution is given a positive value because the energy of the system is increasing.

$$KNO_3(s) \rightarrow K^{+}(aq) + NO_3^{-}(aq) \qquad \Delta H_{soln} = +34.89 \ kJ$$

The energy or enthalpy change associated with the process of a solute dissolving in a solvent is called the heat of solution (ΔH_{soln}). In the case of an ionic compound dissolving in water, the overall energy change is the net result of two processes – the energy required to break the attractive forces (ionic bonds) *between* the ions in the crystal lattice, and the energy released when the dissociated (free) ions form ion-dipole attractive forces with the water molecules.

Heats of solution and other enthalpy changes are generally measured in an insulated vessel called a *calorimeter* that reduces or prevents heat loss to the atmosphere outside the reaction vessel. The process of a solute dissolving in water will either release heat into the aqueous solution or absorb heat from the solution, but the amount of heat exchange between the calorimeter and the outside surroundings should be minimal. When using a calorimeter, the regents being studies are mixed directly in the calorimeter and the temperature is recorded both before and after the reaction has occurred. The amount of heat change occurring in the calorimeter may be calculated using the basic heat equation: $q = mC_p\Delta T$. The specific heat of the solution is generally assumed to be the same as that of water, 4.184 J/g℃ , due to the significant component in the solution being water.

Background:

Instant cold packs are familiar first aid devices used to treat injuries when ice is unavailable. Most commercial cold packs consist of a plastic package containing a white solid and an inner pouch of water. Firmly squeezing the pack causes the inner pouch to break. The solid then dissolves in the water producing a change in temperature. In this lab we will measure the temperature change that occurs in this reaction and determine the heat change for this process.

Safety precautions:

The cold pack solid is slightly toxic by ingestion and is a body tissue irritant. Avoid contact of all chemicals with eyes and skin. Wear chemical goggles and apron.

Procedure:
Complete the following chart to become familiar with the nature and amounts of materials in a commercial cold pack. (Part A)

1. Review a label of commercial cold pack and write the name of the solid used in the pack.

2. Read the warning information on the label and Record any hazards associated with this product.

3. Using the known charges of ions, write the formula of the solid.

4. Calculate the molar mass of the solid

5. Determine the total mass of the solid

6. Calculate the number of moles of solid in the Pack.

7. Determine the volume of water contained In the inner pouch.

Name of solid	
Warning	
Formula of solid	
Molar mass	
Mass of solid	
Moles of solid	
Volume of water	
Mass of water	

8. Calculate the mass of water in the commercial cold pack. (assume the density of water is 1.0 g/mL)

(Reaction procedure – Part B)

9. Obtain a student calorimeter and a thermometer. Set the calorimeter into a 250 mL beaker for stability.

10. Measure 30.0 mL distilled water in a graduated cylinder and add it to the calorimeter. Record the volume of water used in the data table.

11. Place a thermometer in the water in the cup and allow the thermometer and the water to adjust to room temperature for about 2-3 minutes.

12. Tare a weigh boat on the balance. Add about 5 g of cold pack solid crystal to the weigh boat. Measure and record this exact mass of solid to the nearest centigram.

13. Measure the initial temperature of the water in the cup to the nearest 0.1℃ and record the value in the data tab le.

14. Stir the water in the cup with a gentle swirling motion (be careful NOT to swirl any solution out of the calorimeter!!). Make sure all of the solid dissolves in the water.

15. Monitor the temperature of the solution while continuing to gently swirl the solution. Record the lowest temperature that is reached at equilibration.

16. Rinse the cup contents down the drain with plenty of excess water and dry the inside of the cup with a towel.

17. Repeat steps 9-15 for 2 trials (3 trials if time permits).

<u>Data Table:</u>

Trial #	Mass of cold pack solid (g)	Volume of water (mL)	Mass of solution (g)	Initial temperature (°C)	Final temperature (°C)	Temperature change, ? T (°C)
1						
2						
3						

<u>Post-Lab Calculations & Analysis:</u>

1. Calculate the *heat energy in joules* when the cold pack solid dissolved in water in your experiment (show calculations for all trials – use the basic heat equation).

2. Calculate the energy change in *joules per gram of solid* for the cold pack solid dissolving in water (show calculation for all trials).

3. Calculate the energy change in units of *kilojoules per mole* of cold pack solid dissolving in water. To do this:

   (a) convert the heat energy change found in #1 to kilojoules

(b) convert the grams of solid used in the experiment into moles

(c) divide the energy change in kJ by the number of moles of solid to determine the energy change in units of kJ/mol (common thermodynamic units). **Calculate the average value of the heat of solution as well.**

4. Using the result from question #3c and the information obtained in Part A, calculate the kJ involved when the entire cold pack is activated.

5. Circle the correct choices in the following sentence to summarize the heat change that occurs when the commercial cold pack is activated:

> "When the white solid in the commercial cold pack dissolves in water, the pack feels ***(hot/cold)*** because the temperature of the solution ***(increases/decreases)***. Energy is ***(absorbed/released)*** from the surroundings during this reaction and the reaction is classified as ***(endothermic/exothermic)***. The sign of ?H for the heat of solution is ***(positive/negative)***.

6. Calculate $\%_{error}$ for your average value for the heat of solution of the cold pack solid. Assume 25.4 kJ/mol

7. Compare the value in #6 with the calculation for dissociation obtained when strictly employing the $?H_{rxn} = ?H_{prod} - ?H_{rxt}$ equation (use values from your text)

8. Using internet resources, describe the Thermochemistry behind the MRE ("meal ready-to eat") technology used by the military and other industries. (cite your references!!)