Exploring Heat Transfers and Enthalpy of Solution: A Two-step Lab

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INTRODUCTION

The movement, transfer, and measurement of heat is a common theme in both chemistry and physics curricula. In these labs, students will begin by exploring how dissolving different chemical compounds (three ionic and one organic) changes the temperature of the solution. Next, students will use what they've learned to engineer a "thermal pack" device that uses the thermal activity of the salt to control the temperature (either hot or cold) of the solution.

BACKGROUND

The heat of solution of a substance is the amount of energy absorbed or released when the substance is dissolved in solvent. The steps of the heat of solution are:

1. Breaking the solute-solute bonds (the cation and anion in the salt, for instance). This is an endothermic step, requiring energy.

2. Breaking the solvent-solvent bonds (in the case of the water in this lab, the hydrogen bonds in the water molecules). This is also an endothermic step, requiring energy.

3. Forming the solute-solvent bonds. This is an exothermic step, releasing energy.

As described above, the first two steps require energy. This energy comes from the motion of the water molecules which interact with the salt molecules. The amount of energy this requires depends on the type of solute used and is called the *enthalpy of solution.* It can be calculated using the equation below:

$$q_{salt} = m\Delta H_{salt}$$

Where m is the mass of the salt, and ΔH_{salt} is the salt's enthalpy of solution. I've included the enthalpies of solution for the four salts used in this lab below in the *Teacher Guide* section, in *kJ/mol*, *J/g,* and *cal/g*.

As a quick note, I know that urea is an organic amide and not a salt, technically. However, for simplicity's sake, I'm going to use the word "salt" to describe all the compounds in this lab.

The energy that was used to break the bonds of the salt must have come from somewhere, since energy is conserved. This is a topic our students often confuse—they struggle to see that the q that went in to the salt needed to have come from the water, and since the energy is conserved, the amount of energy **leaving** the water must be exactly equal to the energy **breaking** the bonds of the salt:

$$q_{salt} = -q_{soln}$$

Adding up the energy requirements of all three steps gives us the total enthalpy of solution (hereafter ΔH_{soln}). If the energy changes of steps 1 and 2 exceed the energy change of step 3, then the ΔH_{soln} of the salt will be endothermic, and will take in energy, causing the solution to feel cold when the salt and water mix. If the opposite is true, and the energy change in step 3 is greater than the sum of steps 1 and 2, the ΔH_{soln} will be exothermic, and the solution will feel warm or hot when the water and salt mix.

We can use the example below (from the lab) to demonstrate these principles.

$$\boldsymbol{CaCl_{2\,(s)} \rightarrow Ca^{2+}_{(aq)} + 2\,Cl^{-}_{(aq)} + heat}$$

The calcium chloride salt dissolves in water (an endothermic step) and the hydrogen bonds in the water break (also an endothermic step). However, the interaction between the calcium and chloride ions and the water molecules is so exothermic that it "overcomes" the endothermic first two steps, so the process overall releases heat, and the solution feels warm.

The opposite is true with ammonium chloride, as listed below:

$$\boldsymbol{NH_4Cl_{(s)} + heat \rightarrow NH^{+}_{4\,(aq)} + Cl^{-}_{(aq)}}$$

The first two steps are endothermic, and the third is exothermic. However, the heat required from the first two steps is greater than the heat created from the third, resulting in the net absorption of energy, and the solution feels cold as the salt dissolves.

The units of heat most commonly taught include joules and calories. In this experiment, we will use calories, but if your curriculum is designed around the joule, it's easy to adapt. The calorie is defined as the amount of energy needed to raise the temperature of one gram of water by one degree Celsius. Converting between joules and calories is easy, as there are approximately 4.184 joules per calorie. In the data tables below, I often use kilojoules. There are 1,000 joules per kilojoule.

In this two-part lab, students will explore the enthalpy of solution of four different salts—two with a positive ΔH_{soln} (making them endothermic and cooling the solution) and two with a negative ΔH_{soln} (making them exothermic and heating the solution) Students will begin by exploring the salts, discovering which heats and which cools, and how to calculate the amount of heat released with a set mass of salt. The second part of the lab applies that knowledge, and requires the students to design, build, assess, and test a thermal pack device that uses the enthalpy of solution reactions to heat up or cool down.

APPLICATIONS

These labs can be easily adapted for high school and undergraduate chemistry labs, high school and undergraduate physics labs, and middle school labs. In chemistry labs, the emphasis can be on enthalpy and heat of solution. Students can find the molar masses of the chemical salts and can find the enthalpy of solution by dividing q from the moles of salt the students added. In physics and middle school physical science labs, the emphasis doesn't necessarily have to focus on enthalpy, but rather the engineering aspect of moving heat from system to surroundings. Using "calories" and "grams" in place of "joules" and "moles" works pretty well to adapt this way.

For more on the science behind these labs, check out:

[TED-Ed] (2014, Sept. 11) *The chemistry of cold packs - John Pollard*. [video file]. Retrieved from https://youtu.be/hVh-bpAv4_E

Lab 1: Heat Transfer of Salt

Salt may seem like a boring, common powder, but it can give and receive energy from its surroundings! In this lab, we're going to use different chemical salts and we're going to measure the quantity of heat it exchanges with its surroundings.

As a salt dissolves in water, it either takes in energy (in the form of heat) from the water to dissolve, or it releases energy to the water as it dissolves. This energy changes the kinetic energy of the water molecules, and the temperature of the solution changes accordingly.

Two big exchanges are happening in this process: heat is moving to/from the salt, and heat is moving to/from the water as the solution forms. The formulae that govern this process are listed here:

$$q_{salt} = m_{salt}\Delta H_{salt}$$

Where m_{salt} is the mass of the salt crystals, and ΔH_{salt} is the *enthalpy of solution* of the salt, or how much heat we get from each gram of salt. You'll notice that we can rearrange the equation to solve for ΔH_{salt}:

$$\Delta H_{salt} = \frac{q_{salt}}{m_{salt}}$$

This gives us a unit of energy divided by a unit of mass—a ratio of how much heat we get from each gram of salt.

Recall that energy can't be created or destroyed—it must come from and go somewhere. So in our cup, the heat has to move from the salt to the water, and vice-versa. We can easily determine that the heat from the salt must have gone into the water, and our solution. But since the heat is **leaving** one system and **entering** another, one will be positive and the other negative—they are **equal and opposite** to each other!

$$q_{salt} = -q_{soln}$$

Using our other heat transfer formula, we can find q_{soln}.

$$q_{soln} = m_{soln}c\Delta T$$

Where m_{soln} is the mass of the solution (mass of water plus mass of salt), c is the specific heat of the solution (we'll use 1 *cal/g°C*) and ΔT is the change in temperature of the solution, in °C.

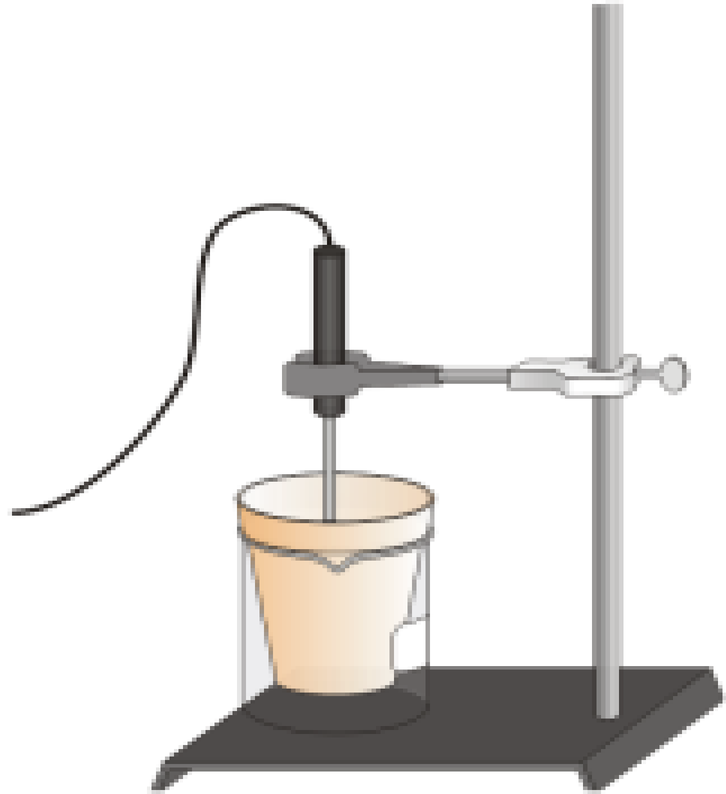
In part one of this lab **your goal is to find ΔH_{salt} for all four salts**, using the information above and your own research.

PURPOSE

To explore the enthalpy of dissolution of different salts, and to calculate the amount of heat transferred.

MATERIALS

Chromebook
Graphical Analysis 4 app
Temperature Probe
Coffee Cup calorimeter
A beaker or 2nd cup (optional)
Ammonium Chloride (NH_4Cl)
Calcium Chloride ($CaCl_2$)
Magnesium Sulfate ($MgSO_4$) Anhydrous
Urea ($CO(NH_2)_2$)



SAFETY

Wear gloves and use splash goggles throughout this lab. Ammonium Chloride and Calcium Chloride are irritants to the skin and eyes. If any gets on your skin, wash thoroughly with soap and water. If any gets in your eyes, alert your teacher. Clean up any spilled salts and wash your hands thoroughly after the lab.

PROCEDURE

1. We're going to use **50 ml** of water and **5 g** of each salt.

2. Prepare a coffee cup calorimeter, with a lid and a temperature probe or thermometer. Take the initial temperature of the water in °C and record that in data Table A.

3. Carefully add the salt crystals to the water and swirl your cup to dissolve the salt. Some might take more swirling than others to dissolve.

4. As you start to swirl, start collecting data with your temperature probe, or **carefully watch your thermometer.** When your temperature reaches the greatest difference from your initial temperature, record that final temperature on Table A.

5. Using your mass of water plus the mass of your salt, and the specific heat of water, calculate the amount of heat the solution gained from the salt.

$$q_{soln} = m_{soln} c \Delta T$$

$$\Delta T = T_{final} - T_{initial}$$

$$c_{soln} = \frac{1.0\ cal.}{g * °C}$$

Divide q by the mass of salt you used to find the heat per gram of salt (this is the enthalpy of solution of the salt).

CONCLUSION

Write a one-paragraph conclusion analyzing and interpreting your data: what salt releases the most energy per gram? Which salts, if any, absorbed energy from the water as they dissolved?

Write an argument from your evidence: how do you know which salt is best?

EXTENSION

Now choose a salt from your table above. We're going to work **backwards,** and have you use what you've learned to predict the ΔT of the water with some settings that you decide. First, decide on a mass of water, between 40 and 80 ml, and a mass of the salt, between 8 and 15 g. Using those settings, use the formulas above to solve for ΔT. This is your prediction of the change in temperature of the solution.

Salt	Mass of Salt (g)	Mass of Water (g)	Total mass of Solution (g)	Initial Temp. of Water (°C)	Final Temp. of Water (°C)	Change in Water Temp. (°C)	Heat gained by Solution (cal.)	Heat/Gram Salt (Enthalpy) (cal./g)	% Error

Measure out that mass of water in your calorimeter and get your initial temperature. Then repeat the process with the mass of salt you selected and get your final temperature. Subtract your initial temperature, and calculate how close you got to your prediction using the formula below:

$$\%\,Error = \frac{Prediction\ Temp. - Experimental\ Temp.}{Prediction\ Temp.} * 100\%.$$

TEACHER GUIDE

Some hints:

- Don't use Epsom salt for your magnesium sulfate. Epsom salt is the heptahydrate form, and is not exothermic enough to provide a good test for the students. Order the anhydrous magnesium sulfate from Flinn or another chemical supplier.

- The chemical salts were selected for their thermochemical activity, and because they are relatively safe compared to more endo/exothermic salts, such as sodium hydroxide and ammonium nitrate. They are also more easily disposable and less hazardous to the environment. However, proper safety measures still need to be taken. Eye protection and gloves should be used. Lab coats or aprons are also recommended. Safety data sheets for these chemicals are available at https://www.flinnsci.com/sds/.

- I used Vernier temperature probes from Vernier's website (www.vernier.com, order code: GO-TEMP). Their interface software Graphical Analysis is free to use on Chromebooks and other devices. Vernier sells a temperature probe that connects via Bluetooth to mobile devices/iPads if that is more convenient for your lab. Of course, analog thermometers work fine as well.

- Polystyrene (StyroFoam©) cups work well in this experiment. Make sure the depth of the solution is adequate for your temperature collecting probes or thermometer. Smaller cups work best, about 8 ounces.

- Because of heat loss to the environment, students are unlikely to measure the full extent of the heat loss/gain according to the physical property table below. However, their data should accurately predict temperature changes.

- According to Flinn, all four salts can safely and cleanly be disposed in a landfill (see https://www.flinnsci.com/disposal-methods/, method #26A). The solutions should also be safe to dispose down most drains but check with local/LEA protocols to ensure you are compliant with all local laws.

The salts and other equipment can be purchased from Flinn Scientific at www.flinnsci.com using the purchase codes in the table below:

Item	P/N	Price
Magnesium Sulfate, Anhyrdous, Lab Grade, 500 g	M0020	$36.50
Ammonium Chloride, Lab Grade, 2kg	A0046	$20.95
Calcium Chloride, Anhydrous, 2kg	C0017	$17.05
Urea, 500 g	U0003	$10.15

The physical data for the salts can be found below:

Salt	Solubility in room temp. water (g/100 ml)	Heat of Solution (kJ/ mol)	Heat of Solution (J/g)	Heat of Solution (cal/g)
$CaCl_2$	74.5	-82.9	-747	-179
$MgSO_4$	35.4	-91.2	-757	-181
NH_4Cl	38.3	14.6	273	65.2
Urea	108	15.0	250	59.8

STUDENT PAGE

This page is designed to be a guided printout of the lab activity for students. This page can be adapted to an Excel or Google Sheets spreadsheet if you are integrating technology into your curriculum, or the table can be copied by students in individual lab journals, if that is your preference.

DATA

Salt	Mass of Salt (g)	Mass of Water used (g)	Total Mass of Solution (g)	Initial Temp. of Water (°C)	Final Temp. of Water (°C)	Change in Water Temp. (°C)	Heat gained by Solution (cal.)	Heat/Gram Salt (Enthalpy) (cal./g)
NH_4Cl								
$CaCl_2$								
$MgSO_4$								
$CO(NH_2)_2$								

CONCLUSION

DISCUSSION

Salt	Mass of Salt (g)	Mass of Water (g)	Total mass of Solution (g)	Initial Temp. of Water (°C)	Final Temp. of Water (°C)	Change in Water Temp. (°C)	Heat gained by Solution (cal.)	Heat/Gram Salt (Enthalpy) (cal./g)	% Error

Lab 2: Homemade Hot/Cold Packs

INTRODUCTION

Now that we've learned about how dissolving chemical salts can absorb or release energy, we can use that knowledge to design a device to control temperature.

Chemical hot and cold packs are designed to provide heat energy or remove heat energy for therapeutic purposes (such as sore or pulled muscles). These devices use these chemical reactions in a controlled system to achieve this measure. For these devices to be effective and safe, it is important they are designed and constructed in a specific way. You now have the skills to achieve this!

PURPOSE

To use knowledge of thermodynamic salts to design and engineer a device to control the rate of temperature in a thermal pack.

SUPPLIES/EQUIPMENT

Chromebook
Graphical Analysis 4 app
Temperature Probe
The chemical salts used in the last activity: $CaCl_2$, $MgSO_4$, NH_4Cl, $CO(NH_2)_2$
Different thermal pack materials—larger plastic bags, smaller plastic bags, rubber bands, hot glue guns, zip-ties, twist ties, string, plastic bottles, etc.
Balance
Weigh boats
Graduated Cylinder
Distilled Water

SAFETY

Wear gloves and use splash goggles throughout this lab. Ammonium Chloride and Calcium Chloride are irritants to the skin and eyes. If any gets on your skin, wash thoroughly with soap and water. If any gets in your eyes, alert your teacher. Clean up any spilled salts and wash your hands thoroughly after the lab.

PROCEDURE

1. Consulting your data from the first part of the lab, you (and your partner) are going to design a hot pack or cold pack that works just like the cold packs in first aid kits and the hot packs found in hand warmers. Start by writing down ideas on how much salt you will need in your hot/cold pack. The pack will need to fit some certain criteria:

a. **It must use 100 ml of water as a starting solvent.**

b. **The hot packs must reach a temperature of 52-60°C and the cold packs must reach between 9-5°C. Temperatures beyond these would be too dangerous!**

c. **The bag must be well-constructed; leaking solution is dangerous!**

d. **The reaction of your hot/cold pack must start when the user wants it to. It must be room temperature until you show it to me, and then a starting mechanism must trigger the reaction.**

2. You have access to different sizes of plastic bags, weigh boats, rubber bands, zip-ties and many other things to design your thermal pack. Design **3 different versions** of a thermal pack, drawing them in your journals or the space provided. You can test your designs using the sand provided in place of the salt. Choose your most promising design to refine, and test it again using the sand. Before you build your final device, have your teacher approve and sign with his/her initials.

3. Once you feel you've constructed a hot pack that fits **all** the above criteria, show it to your teacher. Be prepared to start it and record its change in temperature. If it fits all the criteria, your teacher will sign your journal with his/her initials, indicating you were successful.

4. Use these formulae to help figure out how much salt you need for your pack. Remember, the amount of heat transferred **from the salt** equals the amount of heat transferred **to the solution.**

$$q_{salt} = -q_{soln}$$

As the salt dissolves into the solution, it will change temperature. Using your goal temperature for your thermal pack, you can find the amount of heat that needs to transfer from the salt to the solution.

$$q_{soln} = m_{soln} c \Delta T$$

m_{soln} is the total mass of the solution—the mass of the water plus the mass of the salt. c is the specific heat of the solution—since there's much more water than salt, we'll just use the specific heat of water—1 cal/g°C.

If you were paying attention in the first part of the lab, you should have the enthalpy of solution for each of the salts. This will help you figure out how many grams of salt you need to hit your target temperature.

CONCLUSION

Write a one-paragraph conclusion of your design. Discuss how your group engineered your thermal pack. How did you choose which of your three designs you chose? Which material (twist ties, string, balloon, etc.) best contributed to your success?

DISCUSSION

1. Compare your device to another groups. What worked with their device that didn't work with yours? What worked with your device but not theirs? If you had to purchase a thermal pack, whose design would you trust more, and why?

2. One salt we didn’t use was potassium nitrate (KNO_3). But it would make a pretty good cold pack. Suppose we took 20 grams of KNO_3 and mixed it in 100 grams of water. What would the final temperature of the mixture be, if we started with water at 21°C and KNO_3 took in energy at a rate of 83 calories per gram?

3. Another salt we didn’t use was sodium hydroxide (NaOH). It is tricky to use since it absorbs water from the air very quickly which ruins its effectiveness. But it makes a great chemical hot pack. If you mix 20 grams of NaOH in 100 grams of 21°C, and after all of the NaOH has dissolved the water shoots up to 74°C (too hot to use safely in a hot pack), how many calories of heat energy did you get from each gram of NaOH?

TEACHER GUIDE

Some hints:

- Don't use Epsom salt for your magnesium sulfate. Epsom salt is the heptahydrate form, and is not exothermic enough to provide a good test for the students. Order the anhydrous magnesium sulfate from Flinn or another chemical supplier.

- To save your chemicals and to allow the students more freedom to experiment, I had them use sand in place of the salts until they were confident they had a good design. Large bags of sand are inexpensive and can easily be thrown away, thus saving you your salt until the students are ready to test.

- The chemical salts were selected for their thermochemical activity, and because they are relatively safe compared to more endo/exothermic salts, such as sodium hydroxide and ammonium nitrate. They are also more easily disposable and less hazardous to the environment. However, proper safety measures still need to be taken. Eye protection and gloves should be used. Lab coats or aprons are also recommended. Safety data sheets for these chemicals are available at https://www.flinnsci.com/sds/.

- I used Vernier temperature probes from Vernier's website (www.vernier.com, order code: GO-TEMP). Their interface software Graphical Analysis is free to use on Chromebooks and other devices. Vernier sells a temperature probe that connects via Bluetooth to mobile devices/iPads if that is more convenient for your lab. Of course, analog thermometers work fine as well.

- Most students will figure out that they need to place a smaller bag inside a larger one to make the thermal pack. Despite this obviousness, this lab works best as a true design challenge: give the students all sorts of things they can try to fit the parameters listed above. Some questions I pose to the students are listed below:

  - "How much water should you start with? How much salt?"

  - "How are you going to control when the salt and water mix?"

  - "Are you going to keep the salt in the small bag and the water in the large bag, or vice-versa?"

- Depending on your access to resources and budget, you can use truly numberless combinations of equipment to manufacture the thermal pack. For instance, I used not only twist-tie bags (listed below), but I provide for the students water-soluble bags (made of polyvinyl alcohol), zip-ties, hot glue, tape, heat sealing equipment, easy open Eppendorf tubes, balloons, simple polyester string, and binder clips. There is no one "right answer", if the rules above are met, and the students are staying safe.

- Once source of confusion is how the mass of the salt affects the amount of heat released, and the mass of the solution. Make sure the students are adding the mass of salt to the mass of water to get the mass of solution (which is your m in $mc\Delta T$). Also, make sure that they understand the heat lost by the salt is **equal and opposite** to the heat gained by the solution, and vice-versa, per the first law of thermodynamics.

- Because of heat loss to the environment, students are unlikely to measure the full extent of the heat loss/gain according to the physical property table below. However, their data should accurately predict temperature changes.

- According to Flinn, all four salts can safely and cleanly be disposed in a landfill (see https://www.flinnsci.com/disposal-methods/, method #26A). The solutions should also be safe to dispose down most drains but check with local/LEA protocols to ensure you are compliant with all local laws.

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Calcium Chloride, Anhydrous, 2kg	C0017	$17.05
Urea, 500 g	U0003	$10.15

- Nasco products work well for this lab, especially their “Whirl-Pak ®” bags. They include twist-ties on the top that the students can use to prevent spilling chemicals while working (although they don’t prevent liquids from leaking). These bags can easily be heat sealed, using commercial heat sealers.

The physical data for the salts can be found below:

Salt	Solubility in room temp. water (g/100 ml)	Heat of Solution (kJ/ mol)	Heat of Solution (J/g)	Heat of Solution (cal/g)
$CaCl_2$	74. 5	-82.9	-747	-179
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DESIGNS

CONCLUSION:

DISCUSSION QUESTIONS: