Name(s): Period: Date:

Hot Pack Calorimetry

HASPI Medical Chemistry Lab 6b

Background/Introduction

http://www.chemistryland.com/CHM130FieldLab/Lab11/Lab11.html

Many common reactions in chemistry cause temperature changes. ***Endothermic*** reactions absorb heat, making the environment around them feel colder. ***Exothermic*** reactions release heat, making the environment around them feel warmer. We can use these reactions to provide thermotherapy so that treatment can be quick and can be available anywhere.

Thermotherapy is the use of heat or cold to treat an injury or illness. We use chemical reactions to create heat packs and cold packs.

Cold packs are used most often to reduce inflammation and swelling. This is best for a new injury.

- Cold therapy decreases blood flow
- Less blood flow leads to less swelling and inflamation
- Cold therapy is ideal for an area that is swollen or bruised
- Cold treatment is best during the first 24-48 hours after an injury
- Cold therapy should occur for 20 minutes at a time with at least 10 minutes between applications

Hot packs are used to provide pain relief. , Common uses are for back pain or arthritis pain. This treatment is best for recurring pain.

- Blood vessels are dilated so that there is an increased flow of blood to the muscles
- Blood flow brings oxygen and nutrients so that damaged tissue can heal
- The heat allows soft tissues to stretch so that there is decreased stiffness and increased flexibility This can greatly increase the range of motion
- Heat can relax joints, muscles, ligaments and tendons
- Heat therapy should occur in 20 minute increments

Hot Pack Chemistry

There are a few ways that exothermic reactions can be used as hot packs.

The reaction we will study today mixes anhydrous magnesium sulfate ($MgSO_4$) with water. When these mix heat is released. In a hot pack there is powdered $MgSO_4$ along with a small bag of water. When you squeeze the hot pack, the water packet bursts to begin the reaction. This type of hot pack can only be used one time.

Another type of hot pack uses a supersaturated solution of sodium acetate. Sodium acetate has a boiling point of 130°F, but if you melt it by heating it above that temperature, you can cool it back down and it will stay a liquid until something causes just one molecule to freeze. This causes a chain reaction until all of the sodium acetate freezes. In order for the molecules to solidify and freeze, the energy within them must be released, which is why this feels hot in your hands. These hot packs are reusable as long as you boil them to dissolve or melt the sodium acetate again.

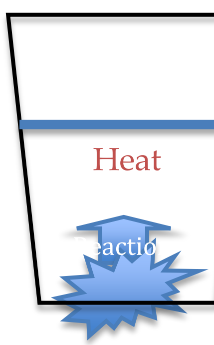
Iron filings can also be used as a hot pack. They react with oxygen in the air, so they are in a mesh packet which is sealed inside of the packaging. Once opened, the iron filings react with the oxygen in the same way they react to create rust, but it happens very quickly because the iron is in very small pieces. Water is able to speed up this reaction, increasing the heat for a moment, but the heat pack will cool more quickly if the reaction speeds up.

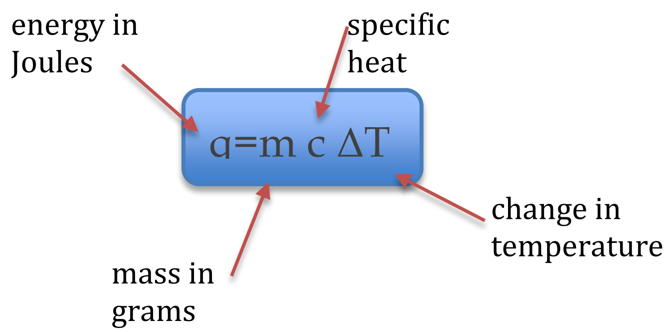
Cold Pack Chemistry
Endothermic reactions can be used as cold packs.

A commonly used cold pack contains ammonium nitrate and water. When the small bag of water inside is put under pressure it bursts to begin the endothermic reaction. This reaction needs energy, so it takes the energy from you, making you *feel* cold. Cold is just the sensation of energy moving away from you.

A New Concept - Calorimetry!
Today we will use calorimetry to study the energy given off by our reactions. Calorimetry allows us to measure the energy given to water by a reaction, or by a hot metal. This is used to find the amount of energy in our food, to identify unknown metals, and in this case, to find the Enthalpy of Hydration for our substance.

In calorimetry we are going to do a reaction in a very well insulated container, which we will make from styrofoam cups. As the reaction proceeds, the water will absorb the energy released. We can use a thermometer to measure the change in temperature of the water, then use the reaction $q=mc\Delta T$ in order to find out the quantity of energy that was released!





Review Questions

1. What is the difference between an exothermic and endothermic reaction?
2. How do you choose which therapy is right for your injury?
3. Why wouldn't a hot pack help with swelling?
4. How do you think you might be able to re-use the chemicals in a Magnesium Sulfate hot pack?
5. If the reaction creating rust (Iron Oxide) is exothermic, why don't rusty cars and nails feel warm?
6. What does the cold sensation mean in chemistry?
7. Why do we need an insulated container for calorimetry?
8. What errors do you predict we will encounter in a calorimetry lab?

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Health and Science Pipeline Initiative

Objectives

Find the amount of energy released when $MgSO_4$ is rehydrated to form $MgSO_4 \bullet 7H_2O$
Specifically:

- Calculate the Enthalpy of Hydration (Δ H) for $MgSO_4$
- Learn Calorimetry techniques
- Apply Hess' Law
- Use the equation $q=mc\Delta T$

Materials

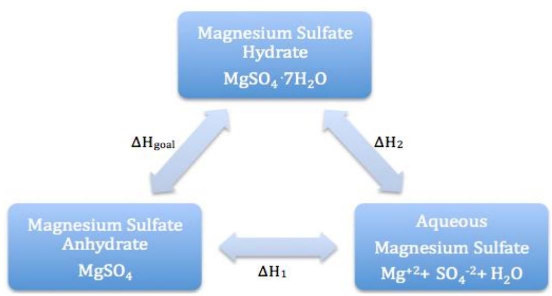
- Foam Cup Calorimeter
- Thermometer
- Anhydrous $MgSO_4$
- $MgSO_4 \bullet 7H_2O$ crystals
- DI or distilled water
- Digital Balance and weigh boats
- 50ml or 100mL Graduated cylinder

Connections to past lab

You prepared the anhydrate by heating the $MgSO_4 \bullet 7H_2O$ (s) last semester in the hydrated crystals lab. It took heat to remove the water from the crystals, which meant it was an endothermic reaction. That energy has been stored in the anhydrate and when the reaction is reversed the energy you put into the MgSO4 with your Bunsen burner will now be released in an exothermic reaction.

Scenario

In this lab we will dissolve Magnesium Sulfate hydrate ($MgSO_4 \bullet 7H_2O$) and magnesium sulfate anhydrate ($MgSO_4$) to find the Δ H values for each process.



A hydrate is a chemical with water trapped in its crystalline structure. When you look at the hydrate you will notice it looks dry, not wet. That is because the water is trapped. We use a dot in the formula of a hydrate to show that the water molecules are trapped.

In order to find the ΔH (enthalpy of hydration) we will perform two separate experiments, and then relate them to find our goal reaction.

Reaction 1: $MgSO_4(s) \xrightarrow{H_2O} Mg^{+2}(aq) + SO_4^{-2}(aq)$
Here we will add water to the anhydrate to dissolve the substance.

Reaction 2: $MgSO_4 \bullet 7H_2O\ (s) \xrightarrow{H_2O} Mg^{+2}(aq) + SO_4^{-2}(aq) + 7H_2O(l)$
Here we will dissolve the hydrated crystal in water.

Goal Reaction: $MgSO_4(s) + 7H_2O(l) \xrightarrow{H_2O} MgSO_4 \bullet 7H_2O\ (s)$
By adding the above reactions together we can determine the ΔH for this reaction.

Using Hess' law you can reverse the second reaction and add it to the first in order to calculate the ΔH of the goal reaction. First we will calculate the ΔH for reaction 1 and reaction 2 using calorimetry.

Procedure and calculations for Reaction 1: $MgSO_4(s) \xrightarrow{H_2O} Mg^{+2}(aq) + SO_4^{-2}(aq)$

1. Obtain a sample of the anhydrous $MgSO_4$. Weigh the sample and record the mass in *Data Table 1*. It should be about 1.50-3.00g.
2. Add exactly 20.0mL of DI water to your calorimeter.
3. Record the initial temperature of the water.
4. Add the $MgSO_4$ anhydrate to your calorimeter and immediately close the lid.
5. Stir until the temperature stops increasing.
6. Record the highest temperature reached as the final temperature.

DATA TABLE 1	
Mass of anhydrous $MgSO_4$ pellet or powder	
Mass of water used (same as mL)	
Initial temperature of the water	
Final Temperature of the water	
ΔT (Final - Initial)	

Calculations for Reaction 1	
1. Use q=mc Δ T to find out how many joules of energy were absorbed by the water. Remember that the specific heat of water is 4.184J/g°C. Be sure to use the mass of water for this calculation. (For water 1mL = 1g)	a) Energy in Joules b) Energy in kJ
2. Find the moles of anhydrate used a. What is the molar mass of $MgSO_4$? b. Convert the grams of $MgSO_4$ in *Data Table 1* into moles	a) molar mass of $MgSO_4$ b) moles of $MgSO_4$ used
3. Find the joules of heat released from the reaction in kJ. Remember, energy absorbed by water is negative the energy released from the reaction.	Energy=
4. Find the Δ H for this reaction in kJ/mol. (divide kJ by moles)	ΔH_1 =
5. Is this reaction endothermic or exothermic?	

Procedure and calculations for Reaction 2:

$MgSO_4 \bullet 7H_2O(s) \xrightarrow{H_2O} Mg^{+2}(aq) + SO_4^{-2}(aq) + H_2O(\ell)$

1. Weigh out about 2.00g of crystalline $MgSO_4 \bullet 7H_2O$
2. Add exactly 20.0mL of DI water to your calorimeter
3. Record the initial temperature of the water
4. Add the $MgSO_4$ hydrated crystals to your calorimeter and immediately close the lid
5. Stir until the temperature stops increasing
6. Record the final temperature reached by the solution after it stops changing

DATA TABLE 2	
Mass of $MgSO_4 \bullet 7H_2O$	
Mass of water used (same as mL)	
Initial temperature of the water	
Final Temperature of the water	
ΔT (Final - Initial)	

Calculations for Reaction 2	
1. Use $q=mc\Delta T$ to find out how many joules of energy were absorbed by the water. Remember that the specific heat of water is 4.184J/g°C. Be sure to use the mass of water for this calculation. (For water 1mL=1g)	a) answer in Joules b) answer in kJ
2. Find the moles of anhydrate used a. What is the molar mass of $MgSO_4 \bullet 7H_2O$? b. Convert the grams of $MgSO_4 \bullet 7H_2O$ in Data Table 2 into moles	a) molar mass of $MgSO \bullet 7H_2O_4$ b) moles of $MgSO_4 \bullet 7H_2O$ used
3. Find the joules of heat absorbed in the reaction in kJ. Remember, energy lost by water is negative the energy gained by the reaction.	Energy=
4. Find the ΔH for this reaction in kJ/mol. (divide kJ by moles)	ΔH_2 =
5. Is this reaction endothermic or exothermic?	

Conclusions: Calculate the ΔH for the goal reaction using the given reactions
1. Label each reaction with the enthalpy of reaction you found in the lab Reaction 1: $MgSO_4(s) \xrightarrow{H_2O} Mg^{+2}(aq) + SO_4^{-2}(aq)$ $\Delta H_1 =$ Reaction 2: $MgSO_4 \bullet 7H_2O\ (s) \xrightarrow{H_2O} Mg^{+2}(aq) + SO_4^{-2}(aq) + 7H_2O(l)$ $\Delta H_2 =$
2. In order to find the goal reaction, the above reactions may be reversed or multiplied as needed so that they add up to make this reaction: $MgSO_4(s) + 7H_2O(l) \rightarrow MgSO_4 \bullet 7H_2O\ (s)$ Remember that a reversed reaction causes the sign on the ΔH value to reverse, and a doubled reaction doubles the ΔH value. Cancel out anything that appears on both sides of the arrows. Show your work as you add these reactions together: Reaction 1: $\Delta H =$ + Reaction 2: $\Delta H =$ Equals Goal Reaction $MgSO_4(s)$ + $7H_2O(l) \rightarrow$ $MgSO_4 \bullet 7H_2O\ (s)$ $\Delta H =$
3. When you add together the ΔH values, what is the ΔH of hydration for $MgSO_4$?
4. If the theoretical value is -104kJ/mol, what is the percent error?
5. List 2 sources of unavoidable error

Applications:
1. Explain the difference between potential energy and kinetic energy in 2-3 sentences.
2. Using the terms *potential* and *kinetic* energy, describe what happened in the first reaction, where you dissolved the anhydrate in water.
3. The law of conservation of energy states that the total energy of an isolated system cannot change. How does this apply to our lab?
4. What could you do to increase the amount of energy released in reaction 1?
5. At the end of the trial, the temperature began to decrease. What happened to the energy that was in the calorimeter?
6. What changes could you have made to the lab apparatus to improve your results, reducing your percent error?

7. Draw a diagram of each experiment, including the calorimeter, the chemical and the water. Use arrows to show the flow of energy between the magnesium sulfate and the water.	
Experiment 1: Dissolving the Anhydrate	Experiment 2: Dissolving the Hydrate

Resources and References

http://www.urmc.rochester.edu/encyclopedia/content.aspx?ContentTypeID=1&ContentID=4483
http://www.spine-health.com/treatment/heat-therapy-cold-therapy/benefits-heat-therapy-lower-back-pain
http://www.middleschoolchemistry.com/lessonplans/chapter5/lesson9
http://home.howstuffworks.com/question290.htm
http://prezi.com/lranljsxv4hx/copy-of-chemistry-hotcold-packs/
Pictures:
http://www.chemistryland.com/CHM130FieldLab/Lab11/Lab11.html