Oakland Schools Chemistry Resource Unit

Thermodynamics & Kinetics

Traci Banjanin
Walled Lake Central High School
Walled Lake Consolidated Schools

Thermodynamics and Kinetics

***Content Statement: C3.1x Hess's Law**: For chemical reactions where the state and amounts of reactants and products are known, the amount of energy transferred will be the same regardless of the chemical pathway. This relationship is called Hess's Law.*

Content Expectations:

C3.1a: Calculate the ΔH for a given reactions using Hess's Law.

C3.1b: Draw enthalpy diagrams for exothermic and endothermic reactions.

C3.1c: Calculate the ΔH for a chemical reaction using simple coffee cup calorimetry.

C3.1d: Calculate the amount of heat produced for a given mass of reactant from a balanced chemical equation.

Instructional Background Information:

Hess's Law of Heat Summation: Many chemical reactions are difficult to carry out separately or under standard lab conditions. Therefore, an indirect and alternate method for determining the enthalpy change (heat change) for a specific thermochemical reaction in question can be obtained using Hess's law of heat summation. This law states that the enthalpy change for an overall process may be obtained by adding the enthalpy changes involved in a set of given thermochemical reactions, whose ΔH's are known.

For a given problem, the following steps should be followed:

1. Identify the thermochemical reaction whose ΔH is unknown and note the number of moles of reactants and products.
2. Manipulate the thermochemical equations, whose ΔH values are known, such that the moles of reactants and products are on the correct side. There are two ways in which these given equations may be manipulated.
   - Reverse the equation and change the sign of ΔH.
   - Multiply numbers of moles and ΔH by the same factor.
3. Add the manipulated equations to obtain the thermochemical reaction whose ΔH is desired. NOTE - all substances except those in the desired reaction must cancel.

4. Add the ΔH values to obtain the unknown ΔH.

Students should be able to work with Hess's law problems where two thermochemical equations are given to formulate the ΔH for a third reaction.

Enthalpy Diagrams: An enthalpy diagram may be used to illustrate the difference between the enthalpy of the reactants and products in any given chemical reaction. The enthalpy change of a reaction, aka heat of reaction, refers to the $H_{\text{products (final)}}$ minus the $H_{\text{reactants (initial)}}$. Enthalpy diagrams can be drawn to illustrate the difference between an exothermic and endothermic chemical reaction. Note in the diagrams below that the ΔH is positive for the endothermic reaction and is negative for the exothermic reaction. Students should be able to label the change in enthalpy (ΔH), reactants, products and activation energy (Ea). The activation energy is the minimal amount of reactant collision energy that is required for the chemical reaction to occur.

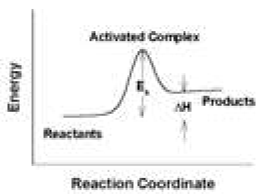

Figure 1 - Endothermic Enthalpy Diagram

This diagram illustrates an endothermic reaction because energy is absorbed and the products possess more energy than the reactants. This gives a positive change in enthalpy and when the energy is absorbed into the products, the temperature will decrease as a result.

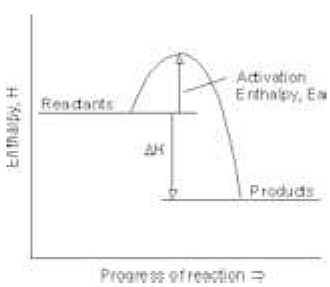

Figure 2 - Exothermic Enthalpy Diagram

This diagram illustrates an exothermic reaction because energy is released and the products possess less energy than the reactants. This gives a negative change in enthalpy and when the heat is released by the system, the released energy will be observed by temperature increasing.

Coffee cup calorimetry: The experimental technique used to measure the heat released or absorbed by a physical or chemical process is known as calorimetry. When a process takes place open to the atmosphere, therefore at a constant pressure, a "coffee-cup" calorimeter is used, and from the gathered data, the change in enthalpy can be calculated. Conceptually, the heat that the surrounding gains is equal to the heat that the system loses, or vice versa. The following mathematical equations are used to determine the change in enthalpy for a reaction:

- $q_{soln} = mass_{soln} \times c_{soln} \times \Delta T_{soln}$
- $q_{rxn} = -(q_{soln})$
- ΔH_{rxn} (kJ/mol) = q_{rxn} / mole reactant

Stoichiometry of Thermochemical Equations: A thermochemical equation is a balanced equation that includes the heat of the reaction. The enthalpy value given is dependent upon the moles of substances and the states of matter in the specific equation.

Grams of substance x Molar mass of substance = Moles of substance

Moles of substance x kJ/ moles of substance from equation = Heat

***C3.4x Enthalpy and Entropy**: All chemical reactions involve rearrangement of the atoms. In an exothermic reaction, the products have less energy than the reactants. There are two natural driving forces: (1) toward minimum energy (enthalpy) and (2) toward maximum disorder (entropy).*

Content Expectations:

C3.4c Write chemical equations including the heat term as part of equation or using ΔH notation.

C3.4d Draw enthalpy diagrams for reactants and products in endothermic and exothermic reactions.

C3.4e Predict if a chemical reaction is spontaneous given the enthalpy (ΔH) and entropy (ΔS) changes for the reaction using Gibb's Free Energy, $\Delta G = \Delta H - T\Delta S$ (Note: mathematical computation of ΔG is NOT required)

C3.4f Explain why some endothermic reactions are spontaneous at room temperature.

Instructional Background Information:

Spontaneous change – When reference is made to a change being *spontaneous*, the implication is that it occurs by itself under the given conditions, without an ongoing input of energy from outside the system. For a *nonspontaneous* change to occur, the system must be supplied with a continuous input of energy from the surrounding. Under a given set of conditions, if a change is spontaneous in one direction, it is nonspontaneous in the other. Also, the term spontaneous doesn't mean instantaneous and in fact, thermodynamic forces are independent of rate. Many spontaneous processes are slow such as ripening, rusting, and aging.

Chemists examine two thermodynamic forces when making predictions about whether a change will occur. **Enthalpy – H** – is basically the internal heat energy locked into the bonds of a structure. **Entropy – S –** is basically the randomness or disorder in a system. Typically reactions that have a ΔH that is negative (reduction in energy as the reaction proceeds from reactant to product) are spontaneous and a ΔS that is positive (increase in randomness/disorganization) are spontaneous.

Gibbs free energy – This constant is one criterion that combines the system's enthalpy and entropy to determine whether a reaction will be spontaneous at a particular temperature. The sign of ΔG tells if a reaction is spontaneous.

If $\Delta G < 0$, the reaction is spontaneous as written,

$\Delta G > 0$, the reaction is nonspontaneous as written,

$\Delta G = 0$, the reaction is at equilibrium.

The table below provides a method for qualitatively determining whether a reaction will be spontaneous or not, when considering all the variables that play a role.

ΔH	ΔS	$-T\Delta S$	ΔG	**Description**
–	+	–	–	Spontaneous at all T
+	–	+	+	Nonspontaneous At all T
+	+	–	+ or _	Spontaneous at higher T; nonspontaneous at lower T
–	–	+	+ or _	Spontaneous at lower T; nonspontaneous at higher T

NOTE - One specific standard requires students to explain why some endothermic reactions (ΔH = +, typically non-spontaneous) are spontaneous at room temperature. This would best be explained by noting that the driving force for this reaction is an increase in entropy (ΔS = +; aka- randomness). It is also important to note that if addressing this concept qualitatively – the product of temperature and change in entropy must be larger than the change in enthalpy for such a reaction to occur spontaneously.

Terms and Concepts

Hess's Law	Enthalpy	Entropy
Exothermic	Endothermic	Gibb's Free Energy
Calorimetry	Spontaneous	

Instructional Resources

Herr, Norman. Hands-On Chemistry Activities with Real-Life Applications, Jossey-Bass, 1999, p.348-372.

nclark.net

Silberberg, Martin, Chemistry: The Molecular Nature of Matter and Change, McGraw-Hill, 2006

www.norton chemistry

Thermodynamics & Kinetics

Activity #1 – Entropy-Driven Reactions

Questions to be investigated

Can an endothermic reaction be spontaneous at room temperature?

Objectives

Upon completion of this activity, students will be able to

- Observe and describe two spontaneous endothermic reactions.
- Qualitatively identify the magnitude of the thermodynamic factors – H, S, G

  & explain whether these variables indicate spontaneity alone and together **C3.4e** and **C3.4f**
- Examine how temperature is impacted in an endothermic reaction.
- Sketch an enthalpy diagram for the reactions observed. **C3.1b** and **3.4d**

Teacher Notes

Two concept building activities are associated with activity #1.

One involves chemicals that are more costly and require proper disposal, which would best be done as a teacher demonstration.

$$Ba(OH)_2 \cdot 8H_2O(s) + 2\ NH_4SCN(s) \rightarrow Ba^{2+}(aq) + 2SCN^{-1}(aq) + 2NH_3(aq) + 10H_2O(l)$$

The other involves inexpensive household chemicals that can be used by ALL students and disposed of down the drain.

$$NaHCO_3(s) + HC_2H_3O_2(aq) \rightarrow Na^{+}(aq) + C_2H_3O_2^{-1}(aq) + H_2O(l) + CO_2(g)$$

Materials

Demonstration: 20 g of barium hydroxide octahydrate – $Ba(OH)_2 \cdot 8\ H_2O$

10 g of ammonium thiocyanate – NH_4SCN

125mL flask

Small wooden block

Stirring rod

Thermometer

Activity: 250mL beaker

Scoopula

Sodium Bicarbonate – $NaHCO_3$ (aka baking soda)

Acetic acid – $HC_2H_3O_2$ (aka vinegar)

50mL graduated cylinder

Safety Concerns

Goggles and gloves should be worn by the teacher when doing the demonstration. Disposal of the demonstration product should be placed into waste just due to the barium ions and disposed of properly. Also, be sure not to sniff the fumes generated by the demonstration as gaseous ammonia is very pungent.

Real-World Connections

Ice cream makers (endothermic reactions) – Old fashioned ice cream makers rely on crushed ice and salt as a refrigerant. Salt has a heat of solution of 4.3 kJ/mole. 4.3kJ of energy are removed from the ice water slurry to dissolve each mole of salt, dropping the temperature of the ice cream maker below 0^oC.

Chemical ice packs – One brand of cold packs contains water separated by ammonium nitrate. When this pack is squeezed, the reactants mix and energy is taken in while the ammonium nitrate dissolves and the packs temperature cools as a result of the energy of the surrounding being absorbed.

Expansion of refrigerant gases (such as Freon) in an air conditioning unit.

Sources

Adapted from: Herr, Norman. Hands-On Chemistry Activities with Real-Life Applications, Jossey-Bass, 1999, p.348-372.

Procedure/Description of Lesson

Demonstration

1. Mass out 20 g of barium hydroxide octahydrate and place into the flask.
2. Mass out 10 g of ammonium thiocyanate and place into the flask of barium hydroxide.
3. Set the flask of reactants on top of a few drops of water on the wooden block.
4. Place a thermometer into the flask and measure the temperature of the reactants and note the state of each of the reactants.
5. Stir the reactants with a stirring rod and have the students observe the products state, measure the temperature of the products, and once the flask causes the water underneath to freeze, lift the flask from its opening and have the students observe that the flask is now frozen to the block.

Activity

1. Obtain a beaker, graduated cylinder, and scoopula.
2. Place 1 scoopula of sodium bicarbonate in the beaker and note the state of the sodium bicarbonate.
3. Obtain 10mL of acetic acid in the graduated cylinder, measure the temperature of the acetic acid, note the state of the acid, and then place it into the beaker of sodium bicarbonate.
4. Observe and record the state of the products and the temperature of the mixture.
5. Pour the products down the drain and clean up.

Assessment Ideas

Inquiry Questions/Activities

One way to assess student understanding would be to have the students do the hands-on activity for learning and then use the demonstration as a formative assessment tool. Included below are questions that may be asked by the instructor while the students are doing the lab or could be placed into a worksheet that the students fill out while doing the activity or upon completion of the activity.

Demonstration Relevant Questions

1. Of the states of matter involved in the demonstration (solid or liquid), which has the highest entropy (most randomness)?
2. Comparing the reactant state and product state, what is the sign of ΔS? Explain whether this thermodynamic factor is spontaneous or not when individually examined.
3. What is the sign of ΔH for the demonstration? Explain how you came to your answer and explain whether this thermodynamic factor is spontaneous or not when individually examined.
4. Explain why the flask freezes to the wooden block.
5. What is the sign of ΔG? Explain whether this thermodynamic factor is spontaneous or not when individually examined.
6. Sketch out an enthalpy diagram for this demonstration and label the reactants H, products H, Ea, and ΔH.
7. What thermodynamic factor (S or H) is the driving force behind this reaction being spontaneous at room temperature?

Activity Relevant Questions

1. Write a complete balanced equation for the reaction between baking soda ($NaHCO_3$) and vinegar ($HC_2H_3O_2$) including phase.
2. Based on your observations, what is the sign of ΔH? Explain using complete sentences.
3. Based on your observations, what is the sign of ΔG? Explain using complete sentences.
4. Based on your observations, what is the sign of ΔS? Explain using complete sentences.
5. Assuming that your values for ΔH and ΔG are correct, what must be the sign of ΔS? Use the Gibbs free energy equation to justify your answer. Show your work.
6. What is the driving force behind this reaction? Explain using complete sentences.
7. Would an enthalpy diagram of this experiment be similar or different to that which you sketched for the demonstration?
8. Explain how the enthalpy diagram for an exothermic reaction would be similar and different to the one you sketched for the demonstration.

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Lesson #2 – Entropy & Rubber Bands

Questions to be investigated

How can thermodynamic factors be used to explain the behavior of a stretched rubber bands length when heated?

Objectives

The purpose of this activity is to examine a simple rubber band system to explore the concepts of enthalpy, entropy, and free-energy in a qualitative method.

Teacher Notes

Here is a short version of the explanation for some of the phenomena we see in this activity.

Most of the work which goes into stretching a rubber band is used to stretch the coiled, cross-lined polymers. Stretching the polymers reduces the entropy of the polymers. In the stretched state the cross-linked polymers of the rubber have fewer conformations that they can assume. In the situation where a rubber band is held to a constant length, increasing the temperature increases the amount of force required to hold it to that length (stress increases). Conversely, if the force is kept constant (constant tension--the hanging weight) and the temperature is raised, the stress increases and the length decreases. In the case of the rubber band stretched across the forehead, the rise in temperature of the rubber band is mostly the result of the work acting on the rubber material to stretch being converted into heat.

The stretching of the rubber band is not a spontaneous process because it doesn't occur without energy from an outside source (positive ΔG). It is observed that stretching the band is exothermic (negative ΔH) since heat is released and felt through the senses in the forehead. An analysis of the Gibbs free energy equation indicates that if ΔG is positive, and ΔH is negative, then the change in entropy (ΔS) must be negative if the equation is to be balanced. Therefore, the arrangement of molecules in a stretched rubber band must be less random (lower entropy) than when relaxed.

Materials

Wide rubber bands
Ring stand
Paper clips
2 ring stand clamp
Meter stick
Weights (1kg masses, washers etc)
Hair dryer

Sources

Young and Lovell, Introduction to Po!ymers, 2nd Ed., Chapman and Hall, 1994

Sperling, Introduction to Physical Polymer Science, 2nd Ed., John Wilely and Sons, 1992

Herr and Cunningham, Hands-On Chemistry Activities with Real-Life Applications, Jossey-Bass, 1999

Procedure / Description of Lesson

1. Set up the ring stand with two clamps: one clamp will support the dangling rubber band and the other clamp will hold up the meter stick.
2. Adjust the meter stick so that zero is at the top of the dangling rubber band. Either the ring stand should be pretty tall or it should be set up near the edge of a table so the rubber band, weight and meter stick can hang below the tabletop. You may need to counter weight the ring stand with some massive object(s) (books?). Whether you use one rubber band or two together will be determined by the amount of weight you actually use. I found that two standard large rubber bands (-0.6 cm wide and -9.5 cm long) worked well with 1 kg mass.
3. Take the rubber band and stretch it while it is in contact with your forehead.

   a) Do you notice any temperature change as the band is stretched or allowed to contract?

   (They should notice that it becomes warmer as it is stretched and relatively cooler as it is allowed to contract).

   b) Based on your observations predict whether the ΔH for stretching the rubber band is positive or negative.

   c) Is stretching the rubber band spontaneous or nonspontaneous?

d) **Based upon your answer in (c),** what is the sign of ΔG for stretching the rubber band?

4. Hang the rubber band on one of the clamps and hang the weight on the rubber band using the paper clip to stretch it.
5. Use the set-up described above and measures the length of the rubber band(s) under tension at room temperature and record the length in cm.

   a) **Predict** how the length of the rubber band will change when heated with a blow dryer.

6. Heat the rubber band with the hair-dryer, observe, and record the length of greatest difference from room temperature.

   a) **Did the rubber band behave as you predicted?**

7. Allow the rubber band to cool back to room temperature and observe (and record) the length to which is returns.

Assessment Ideas

1 . How does the elastomer (the rubber band) differ in its behavior from other polymers and most materials?

2. How are the results you obtained in this experiment consistent with the observations you made of the temperatures of the stretched and relaxed rubber band`?

3. Explain the results of your investigation of the rubber band behavior in terms of entropy (when do you think it increases, when do you think it decreases).

4. Try to give a physical description of what the polymers are like when they are stretched and what they are like when they are relaxed.

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Lesson #3 - Activation Energy

Questions to be investigated

What is activation energy?

Objectives

- The student will define the terms endothermic, exothermic, and activation

  energy.
- The student will construct and interpret an energy diagram showing the progress of an exothermic and endothermic reaction.

Materials

Strike Anywhere Match

Brick, Lasagna Noodle - 1/group

Real-World Connections

In order to burn a log & have a campfire (an exothermic combustion reaction), a minimal amount of energy must be put into the system (Ea – activation energy) to get combustion reaction to occur.

Procedure / Description of Lesson / Assessment

1. Obtain one "strike anywhere match" (aka portable chemical reaction stick).
2. Cause a reaction.

Think, Discuss & Answer

a) What type of energy change was observed (exothermic or endothermic)? Justify your answer.

b) Compare the amount of energy in the match prior to the reaction and after the reaction.

c) Explain the energy changes involved in the activity.

d) Sketch an enthalpy graph to show the relative amounts of energy in the match before and after the reaction.

3. Take the brick & use different positions of the brick to illustrate the two energy states of the match (unburned or burnt match).

   Think, Discuss & Answer

   a) Draw and label the brick positions.

   b) When the brick is standing on one small end, does it have more or less potential energy? Explain.

   c) Does the brick in this position represent the unburned or burnt match?

   d) When the brick is lying flat, does it have more or less potential energy? Explain.

   e) Does the brick in this position represent the unburned or burnt match?

   f) Why doesn't the brick fall down spontaneously? What must be done to make the brick fall?

4. Set the brick standing on one small end and place the lasagna noodle beneath the brick.
5. Tap the brick so that it lightly rocks but doesn't fall over.
6. Now tap the brick briskly so that it falls onto the lasagna noodle.

   Think, Discuss & Answer

   a) Why doesn't the brick fall over when tapped lightly?

   b) Why does the brick to fall over onto the lasagna noodle?

   c) How do steps 5 and 6 correspond to the activation energy concept?

Lastly, use the brick to show a visual representation of an endothermic reaction and explain.

Assessment Ideas

The activity itself is comprised of questions that the students can discuss in the lab. The instructor may assess formatively by circulating around the lab and asking the students to demonstrate various aspects of the activity and explain how it translates to an understanding of endothermic/exothermic/activation energy.

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Lesson #4 – Heat / Calories / Mass

Questions to be investigated

How many calories (joules) of heat are given off when a candle burns?

Objectives

The purpose of this lab is to experimentally determine the heat of combustion per gram of candle wax, write out the thermochemical equation for the combustion of wax (C3.4c**)**, and to practice calculating the amount of heat produced for a given mass of wax combusted (C3.1d**)**

Teacher Notes

When heat is absorbed by liquid water, the temperature of water rises. The amount of heat necessary to raise the temperature of one gram of water by one degree Celsius is reasonably constant between 8 ^{o}C and 80 ^{o}C. Consequently, it provides a simple and reproducible basis for a definition of a standard amount of heat, the calorie. The CALORIE is the amount of heat to raise the temperature of one gram of water one degree Celsius. Conversely, one calorie is released as one gram of water is cooled one degree Celsius. (1 cal= 4.184 J)

The student will mass a candle before burning it. They will then capture the heat from the burning candle with the water. The candle will be re-massed at the end of the experiment. Calculate the amount of heat gained by the water and divide by the amount of candle wax that was burned to obtain the calories per gram of wax.

Materials

1 short candle and a small piece of foil

1 coffee can with holes punched to allow a glass rod through

1 large can with top and bottom removed

1 glass rod, thermometer, balance, match

Safety Concerns

Students should exercise caution to ensure they don't get burned.

Real-World Connections

Cooking/Grilling & Heat Transfer

Exercise, Diet & Calories

Procedure / Description of Lesson

1. Mass a candle & foil/cardboard piece to the nearest 0.01 g. Record this in your data table.
2. Mass the coffee can w/ the glass rod to the nearest 0.01 g and record its mass.
3. Fill the coffee can one-fourth to one-third full of water and mass it again to the nearest 0.01 g.
4. Place the thermometer in the water and take the initial temperature of the water to the nearest 0.1 ^{o}C.
5. Stand the candle on the small piece of foil/cardboard and place the open ended can so the candle is inside of it.
6. Using a ring stand and clamp, hang the water-filled coffee can from the clamp directly above the candle set-up.
7. Light the candle and heat the water until the initial temperature rises 10 to 15 degrees. Record the highest temperature reached to the nearest 0.1 ^{o}C.
8. Mass the candle after burning to the nearest 0.01 gram.

Data

MASS OF CANDLE/FOIL BEFORE BURNING (g)	
MASS OF EMPTY COFFEE CAN (g)	
MASS OF COFFEE CAN AND WATER (g)	
INITIAL TEMPERATURE OF WATER (^{0}C)	
FINAL TEMPERATURE OF WATER (^{o}C)	
MASS OF CANDLE AFTER BURNING (g)	

Calculations (Show Work)

1. Calculate the mass of the candle wax that combusted (burned) (g).
2. If the formula for the candle wax is $C_{25}H_{52}$, what is the molar mass?
3. Calculate the number of moles of was that combusted in the lab.
4. Calculate the mass of water in the coffee can.
5. Calculate the temperature change of the water.
6. Calculate the calories of heat absorbed by the water?

   C_p of water = 1cal/g ^{o}C (Where did this heat come from?)
7. Calculate the heat of combustion of candle wax in calories per gram of wax.
8. Calculate the molar heat of combustion of the candle wax in joules/mole.

Assessment Ideas

Assessment would begin by working around the lab helping the students work through the procedure, data gathering and calculations for the lab.

The lab report would be evaluated.

Analysis Questions – perhaps added to the lab report document

1- Write out the thermochemical combustion reaction for the wax (C25H52).

2- Is the reaction observed, exothermic or endothermic?

3- If you had 8.3g of candle wax to combust, how many joules of heat could be generated?

4- Design a lab that could be conducted to experimentally determine the calories/g in a potato chip.

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Activity #5 – Heat of Reaction - Hess's Law

Questions to be investigated

How can calorimetry be used to measure the change in enthalpy for 3 chemical reactions?

How can the data gathered be used to verify Hess's Law?

Objectives

In this experiment, you will measure and compare the quantity of heat involved in three reactions. These heats of reaction will be measured using a Styrofoam calorimeter. The three reactions are shown below.

Reaction 1: The dissolving of solid sodium hydroxide in water.

$$NaOH(s) \longrightarrow Na^+(aq) + OH^-(aq) + heat$$

Reaction 2: The reaction of solid sodium hydroxide with dilute hydrochloric acid.

$$NaOH(s) + H^+(aq) + Cl^-(aq) \longrightarrow Na^+(aq) + Cl^-(aq) + H_2O$$

Reaction 3: The reaction of sodium hydroxide solution with dilute hydrochloric acid.

$$Na^+(aq) + OH^-(aq) + H^+(aq) + Cl^-(aq) \longrightarrow Na^+(aq) + Cl^-(aq) + H_2O$$

Teacher Notes

The foundation of the study of thermochemistry was laid by the chemist Germain Hess, who investigated heat in chemical reactions during the last century. One statement of the law that bears Hess's name says:

The enthalpy change for any reaction depends on the products and reactants and is independent of the pathway or the number of steps between the reactant and product.

Materials

2 large Styrofoam cups
100 mL graduated cylinder
sodium hydroxide, NaOH
0.50 M sodium hydroxide solution
0.25 M Hydrochloric acid solution
0.50 M Hydrochloric acid solution
thermometers
balance
250 mL beaker
glass stir stick

Safety Concerns

Hydrochloric acid and sodium hydroxide are corrosive. Avoid direct contact. If any touches your skin, wash it off immediately. Solid sodium hydroxide is especially dangerous because it absorbs moisture rapidly from the air, forming an extremely corrosive liquid. Avoid spilling this solid, and if a spill occurs, clean it up immediately. Be sure to close the lids of bottles of sodium hydroxide securely, immediately after using.

Dispose of solutions according to your teacher's instructions. A lab coat or apron is strongly recommended.

Procedure / Description of Lesson

Part One: The Dissolving of Solid Sodium Hydroxide in Water

1. Read through the procedure and look carefully at the data to be gathered.

Based upon the data you will need to **create your own** data table for the lab.

*** Below is a sample data table – should the students not have the background to make their own data table.

2. Put 200 mL of cool distilled water into your nested Styrofoam cup calorimeter. Stir carefully with a thermometer until a constant temperature is reached. Measure and record this temperature.

3. Accurately find and record the mass of about 2 grams of solid sodium hydroxide. Perform this operation as quickly as possible since the solid absorbs moisture from the air very rapidly and forms a very corrosive liquid.

4. Place the solid sodium hydroxide into the water in the cups. Stir gently with the thermometer until the solid is completely dissolved and record the highest temperature reached.

5. Discard the solution safely and rinse the cup thoroughly with water.

Part Two: The Reaction of Solid Sodium Hydroxide with Hydrochloric Acid

1. Repeat steps 2-4, but replace the 200 mL of water with 200 mL of 0.25 M hydrochloric acid.

2. Discard the solution and again rinse the cup thoroughly before proceeding to

Part Three: The Reaction of Sodium Hydroxide Solution with Hydrochloric Acid

1. Accurately measure 100 mL of 0.50 M hydrochloric acid solution into your calorimeter and 100 mL of 0.50 M sodium hydroxide into a 250 mL beaker. Both solutions should be at or near room temperature.

2. Record the temperatures and volumes of each solution.

3. Add the sodium hydroxide solution to the acid solution in the Styrofoam cup. Stir the mixture with the thermometer and record the highest temperature reached.

4. Discard the solution as directed by your teacher.

Data Table

Heat of Reaction - Hess's Law -Data Table

Data for Rxn 1

Mass of Styrofoam cup =	g
Mass of cup and water =	g
Initial temperature =	^{o}C
Final temperature =	^{o}C
Mass of NaOH or KOH =	g

Associated Calculations for Rxn 1

Mass of H_2O used =	g
Temperature change =	^oC
Molecular mass of NaOH or KOH =	g/mole
Moles of NaOH or KOH =	mole
Energy released (E) =	kJ
H dissolution of NaOH or KOH =	kJ/mole
Accepted H dissolution of NaOH or KOH from a reference source =	kJ/mole
% error =	%

Data for Rxn 2

Mass of Styrofoam cup =	g
Mass of cup and 0.25 M HCl solution =	g
Initial temperature =	^oC
Final temperature =	^oC
Mass of NaOH or KOH =	g

Associated Calculations for Rxn 2

Mass of 0.25 M HCl used =	g
Moles of HCl actually in solution =	moles
Temperature change =	^oC
Molecular mass of NaOH or KOH =	g/mole
Moles of NaOH or KOH =	mole
You may assume that the value of Cp for 0.25 M HCl is the same as that	

for water. (4.184 J/g^{o}C)	
Energy released (E) =	kJ
Hrxn =	kJ/mole

Data for Rxn 3

Volume of 0.50 M HCl solution used =	mL
Volume of 0.50 M NaOH or KOH solution used =	mL
Initial temperature of HCl solution =	^{o}C
Initial temperature of the NaOH or KOH solution before mixing =	^{o}C
Highest final temperature of mixture =	^{o}C

Associated Calculations for Rxn 3

Change in temperature =	^{o}C
Moles of HCl used =	moles
Moles of KOH or NaOH used =	moles
Energy released =	kJ
Hrxn =	kJ/mole

Calculations (Show Work)

1. From your data, calculate the following for each part of the experiment.

a) The temperature change of the liquid.
b) The heat absorbed by the solution. (You are to assume that the solution has a $Cp=4.184\ J/g^{o}C$.

c) The number of moles of sodium hydroxide present.

For Parts One and Two this can be found from the grams used.

For Part Three the number of moles of sodium hydroxide is found from the molarity and the volume used.

d) The amount of heat evolved per mole of sodium hydroxide used. This final value is the heat of reaction.

2. The accepted value for the heat of solution of sodium hydroxide can be found in a chemistry handbook. Compare your experimental value with this accepted value and calculate the experimental error.

Assessment Ideas

1- Conclusion questions to be done upon completion of the lab -

1. Add the ionic equations given in the introductions for Parts One and Three. Compare the result with the ionic equation for Part Two.

2. Compare the sum of the heats of reaction for Parts One and Three with that obtained for Part Two. In the light of your answer to Question 1, explain your results here.

3. Discuss Hess's law in terms of the law of conservation of energy and in terms of the three parts of this experiment.

4. Suppose you had used 8 g of sodium hydroxide in Part One.
   a) How would this have affected the change in temperature?
   b) What quantity of heat would have been evolved in your reaction?
   c) What effect would this have had on your calculation of the heat of reaction for Part One?

5. Write the net ionic equation for Part Three of this experiment. Write the net ionic equation for the reaction between solutions of potassium hydroxide and sulfuric acid. Compare the two net ionic equations. What does the heat of reaction for Part Three of this Experiment represent?

2- Based upon doing activity #4 and 5, have the students obtain a flameless ration heater (FRH) and ask them to investigate what the contents of the bag are made of and experimentally determine the amount of heat generated per gram of the pad (kJ/g). An FRH is used by the United States Army for use by soldiers in the field to heat their food rations. The use of FRH's provides an interesting, real-world application of thermochemistry and allows the students to demonstrate mastery of calorimetry **Please note that Activity #6** is that activity to be used as a resource for the teacher to assist in this inquiry-based assessment.

Thermodynamics and Kinetics

Activity #6 – Enthalpy of an FRH

Questions to be investigated

What is the enthalpy change per gram of a FRH?

Objectives

Students will be examining the internal composition of an FRH.

Students will experimentally determine the amount of heat released per g of the FRH using calorimetry.

Teacher Notes

The metallic pad in the FRH is composed of a material containing magnesium-iron alloy and sodium chloride dispersed in a polyethylene matrix. The FRH has the ability to increase the temperature of 151g of food approximately 55 ^{o}C. The chemical reaction of the FRH pad and water involves the oxidation of magnesium to form magnesium hydroxide and hydrogen.

$$Mg(s) + 2\ H_2O(l) \rightarrow Mg(OH)_2(s) + H_2(g) + 352kJ$$

This reaction is thermodynamically favored but has a slow reaction rate therefore the FRH has iron in the composite to act as a promoter (this initiates the reaction at the magnesium surface by producing reaction intermediates) and sodium chloride acts as a catalyst to break down the protective coating on the surface of the magnesium metal. The chloride ions replace the hydroxide ions in the magnesium coating to all water to penetrate the reactive metal surface. The chloride ions redissolve in solution and then function as a true catalyst.

Materials

Flameless Ration Heaters

Balance

Erlenmeyer flask

Thermometer

Safety Concerns

Handle the inner FRH pad using forceps or tongs. Magnesium metal is flammable and burns with an intense flame. Safety goggles should be worn and disposal of the product should be neutralized prior to disposal.

Sources

Flinn – ChemTopic Labs – Experiments & Demonstrations in Chemistry. Thermochemistry Volume 10, Flinn Scientific, 2002, p.75.

Procedure / Description of Lesson

1. Obtain all materials.
2. Cut the FRH bag and remove the inner pad.
3. Measure the mass of the FRH inner pad and record.
4. Cut a 3 cm^2 section of the pad and mass it and record.
5. Add 100mL of water to an Erlenmeyer flask and measure its initial temperature and record.
6. Add the FRH section to the water and stir. Record the highest temperature that the water reaches.

Assessment / Calculations

Post Lab Calculations

1. Calculate the amount of heat gained by the water in the flask from the FRH pad.
2. What is the amount of heat released by the FRH pad in kJ?
3. What is the enthalpy (kJ/g) of the FRH pad sample tested?
4. Calculate the total amount of heat evolved from the entire FRH pad.

Further Inquiry - Investigate and qualitatively verify the composition of the FRH pad

The students could be asked to design an experiment to qualitatively test to confirm the presence of Mg, Fe, and NaCl in the FRH pad. One separation technique could be to use a magnet and separate out the iron filings from the magnesium and sodium chloride. If the FRH pad is reacted in water and filtered, the filtrate will contain sodium chloride solution and magnesium hydroxide. A flame test can be conducted using the filtrate to confirm the presence of Na^+ ions and the presence of hydroxide ion in solution can be confirmed testing the filtrate using phenolphthalein. Confirmation of the presence of Mg can be done by burning the FRH pad and noting the familial bright light emitted by Mg.

Thermodynamics and Kinetics

Hess's Law Norton Website Tutorial / Review / Practice

Questions to be investigated

How can Hess's Law be applied to calculate the ΔH for a given reaction?

Objectives

The purpose of this activity is to serve as a tutorial to reinforce student understanding of Hess's Law and to provide practice and immediate feedback on Hess's practice problems.

Teacher Notes

This is a nice visual 7 section tutorial that could be used in a variety of ways while studying thermochemistry. It could be used after instruction to give students a visual of the process of Hess's Law applied, give practice problems for calculating ΔH using Hess's Law (C3.1a) and provide immediate feedback on calculations. It may also serve as an instructional aid for students who are absent.

This activity also reinforces the ways in which given thermochemical equations enthalpy values are impacted when the reaction is reversed or mole quantities are changed.

Materials

Computer and Internet Access

Real-World Connections

Many reactions are difficult or impossible to carry out separately. Examples of these may be complex biochemical processes, those which take place under extreme environmental conditions or one portion of a multi-stage reaction. If we want to study the enthalpy change for a reaction that can't be run in the lab, Hess's Law may be used to qualitatively or quantitatively determine the change in enthalpy for any reaction that an equation may be written.

Sources

http://www.wwnorton.com/college/chemistry/gilbert/home.htm

Procedure/Description of Lesson

1. Go onto the internet and go to: http://www.wwnorton.com/college/chemistry/gilbert/home.htm
2. Click onto chemtutorials.
3. Go to chapter 11 and click onto the Hess's Law tutorial- sec 11
4. Go through the Hess's Law tutorial.

Assessment Ideas

Students work through the 7 section tutorial and work the problems on the tutorial and get immediate feedback.

http://www.nslc.ucla.edu/STEP/GK12/lessons.htm#Chemistry

Flamin' Foods: Calorie Investigation

Developed by: M. Sam, H. Jonson & H. Kang

Teacher Material

Synopsis

The goal of this inquiry is to have students design their own calorimeter to measure the amount of heat released from burning a food item by measuring the change in temperature of water. They will employ the scientific method to solve this problem. In their design, they would have to consider the concept of heat flow, specifically consider strategically how to transfer all the heat released into the system. They will be familiar with the terms calorie and specific heat and conceptualize what a calorimeter is.

Background

- Proficient in mathematical calculations
- Understand the concept of heat flow
- Understand concept of specific heat and apply the specific heat equation
- Know the density of water

Objectives

- Students design their own calorimeter based on the definition of a calorimeter
- Students utilize scientific method to measure heat flow
- Students understand definition of calorie and nutritional Calorie and basis of Nutrition Facts on labels of food items
- Students will prepare and submit a full scientific report

Suggested Timeline

- Day one. Project proposal, team selection, perform sample calculations, sign agreement form and strategic planning
- Day two. Design, test and fine-tune calorimeter apparatus and discuss and critique design with entire class
- Day three. Begin data collection and data analysis
- Day four. Data analysis continue and report write-up
- Day five. Report due – no exception

Materials (*each group limited to 1of each for certain items*)

Aluminum cans	aluminum foil
Wire hangers	thermometers
Electronic balance	push pins
Rubber stoppers	corks
Beakers	water
Matches	tape
Paper clips	rulers
Rubber bands	tooth picks

Snack foods (cheese puffs, peanuts and marshmallows)
(Optional: can bring other food items if desired)

Student Procedures

Day One- Contract must be done before proceeding to day two!

A. Project proposal (10min)

- Distribute project proposal to each student
- Class discussion of the proposal and outline objectives, expectations and compensations
    a. Volunteer a student to read proposal to class and emphasize key points and ideas
    b. Outline the layout of final report (see student worksheet for detail)
        i. Report must be TYPED and organized as follows:

Page 1: Agreement/Contract Form

Page 2: Purpose – State in your own words the goals of this Investigation
Materials – List all materials used for this investigation
Include an illustration/photo of calorimeter

Page 3: Procedure – List step-by-step how measurements were done
Must be logical and clear
Explain reasoning of how procedures and calorimeter design actually measured the calorie content

Page 4: Data – Include data table of measurements
All calculations must be neat and well presented

Page 5: Results – Summarize results and conclusions from your investigation
Compare results to nutritional information on actual snack food label

Error Analysis – Discuss the accuracy of your procedure
Discuss accuracy of calorimeter
Was all heat from food transferred to water?
How would you improve calorimeter design?

Page 6: Contribution – indicate the role and participation of each team member (each member rate him/herself)

B. Team selection (10min)
    - Students select team members (maximum of five per team)
    - Each group decides on a team name
C. Sample Calculations - Interview Process (25min)
    - Each team must understand and complete the sample calculations
    - Work must be shown on paper
D. Accept and sign proposal
    - Each member of the team must sign the agreement to abide by the rules and regulations set forth in the proposal and accept the investigation
E. Strategic Planning (Optional – only time permits)
    - If team completes all of the above, they may begin their strategic planning (i.e. how to design the calorimeter, etc…)

Day Two

A. Each team begins calorimeter design
    - All materials are available on side benches
    - Allow time for students to struggle and figure out what to do before probing them
    - As calorimeter is being built, ask team about the design
        a. Ask about each part of the design –
            i. What is used to hold water and why choose that?
            ii. How food item is burned?
        b. Consider the heat flow process
        c. Is design efficient in terms of capturing all heat into system?
        d. Enough insulation? Etc…
    - Test calorimeter and take preliminary data
        a. Fine tune calorimeter to perfect design
B. Calorimeter discussion
    - Once all teams are happy with design, share with the class
        a. Discuss all factors associated with calorimeter
        b. Are all factors answered with design?

Day Three

A. Data collection
    - Each team is given a data table (or ask to generate one)
    - Begin experiment and record all measurements and observations
    - Each team must choose at least three food items
        a. For each item, must test at least three times
B. Data analysis
    - After data acquisition, begin data analysis
        a. Follow similar calculations and determine the amount of calories released from each food item
        b. Compare experimental values to nutritional values on food labels

Day Four

A. Report write-up
- Begin writing report as specified
- All members must contribute to work

Day Five

A. Report submission!!

Teacher's Tips

a. Allow time for students to figure out how to solve problem
b. Limit amount of materials per group if supplies are limited
c. Continue to question students about their design to help them understand the basis of this investigation
d. Have students bring in aluminum cans and hangers either make it mandatory or for extra points toward their report
e. Be careful with fire since students are working with matches
f. STRESS SAFETY IN LAB AREA!

Student Handout

Flamin' Foods: Calorie Investigation *FSFIC*

Background

In fall of 2001, a woman sued the makers of Pirate's Booty, a type of puffed rice snack, similar to cheese puffs. The snack was recalled after the Good Housekeeping Institute found that one serving contained 147 Calories and 8.5 grams of fat, while its label said it contained only 120 Calories and 2.5 grams of fat. Many other companies are being investigated as a result of these findings.

Job Opportunity

H. Jonson, head of the Fremont Snack Food Investigation Corporation (FSFIC) has offered you a contract to test the accuracy of the nutritional labels of three snack foods, manufactured by different companies. You and your team must carry out the appropriate experiments to determine the amount of Calories contained in these three snack foods. Specifically, your team must design your own procedure and calorimeter apparatus to determine the amount of heat released by burning various snack foods and compare your findings to values reported on the nutritional labels.

> *A calorimeter is a device used to measure the heat associated with a chemical reaction. The heat released by burning an object will transfer to water in a calorimeter, which can be measured as a temperature change in the water.

Your team will be provided with a variety of materials to build the calorimeter.

Materials

FSFIC will provide the following materials for your investigation: aluminum cans, aluminum foil, wire hangers, thermometers, electronic balance, push pins, rubber stoppers, corks, beakers, water, matches, tape, paper clips, rulers, rubber bands, tooth picks and snack foods (cheese puffs, peanuts and marshmallows). FSFIC has also granted you the choice to choose other food snacks to test; it is to your team's discretion.

Timeline

You and your team may revise the procedure and calorimeter design at any time, but the schedule and deadline must be strictly followed.

Day one. Project proposal, team selection, perform sample calculations, sign agreement form and strategic planning

Day two. Design, test and fine-tune calorimeter apparatus and discuss and critique design with other teams

Day three. Begin data collection and data analysis

Day four. Continue with data analysis and report write-up

Day five. Report due by noon – no exception!

FSFIC

Salary and Expectations

If all expectations are met, each team will receive a stipend of **75 points**. Below are the necessary components, organization and expectations for the final report. **IT MUST BE TYPED!** No handwritten reports will be accepted (except for calculations and illustrations).

Page 1: Agreement/Contract Form

Page 2: Purpose – State in your own words the goals of this investigation
Materials – List all materials used for this investigation
Include an illustration/photo of calorimeter

Page 3: Procedure – List step-by-step how measurements were done
Must be logical and clear
Explain reasoning of how procedures and calorimeter design actually measured the calorie content

Page 4: Data – Include data table of measurements
All calculations must be neat and well presented

Page 5: Results – Summarize results and conclusions from your investigation
Compare results to nutritional information on actual snack food label
Error Analysis – Discuss the accuracy of your procedure
Discuss accuracy of calorimeter
Was all heat from food transferred to water?
How would you improve calorimeter design?

Page 6: Contribution – Indicate the role and participation of each team member

Calorie Investigation Application Form

Names of all persons completing this application form:

1 4

2 5

3 Team Name:

Qualifying Exercise.

There are many calculations involved during the Calorie Investigation that you much be familiar with before you begin your own. To guarantee the success of the investigation, your team must demonstrate your understanding of the necessary calculation process. Complete the questions below:

A candy bar has a total mass of 28.375 grams. In a calorimetry experiment, a 3.5 gram sample of this candy bar was burned in a calorimeter surrounded by 1000g of water. The temperature of the water in contact with the piece of burning candy bar was measured and found to increase from an initial temperature of 21.2 °C to a final temperature of 24.3 °C. The specific heat of water is 1 cal/g°C.

1. Calculate the amount of heat (in calories) released when the 3.5g sample burned.

2. Convert the heat to nutritional Calories (1000 calories = 1 nutritional Calorie).

3. Calculate the energy content of the candy bar (in Calories per gram) by dividing the heat in nutritional Calories (#2) by the mass (in grams) of the snack that was burned.

4. Find the total Calorie content of the candy bar by multiplying the energy content (in Calories per gram) by the total mass of the candy bar.

Agreement Form

We agree to abide by all the safety rules and regulations set forth by the Fremont Snack Food Investigation Corporation (FSFIC). Our team will attempt the investigation to our fullest capability. We also understand that any member of the FSFIC has the right to amend this contract at any time. We are aware of the deadline for this investigation and plan to submit our completed report on or before ***noon on Friday, March 4, 2005****.*

Signature Date

Signature Date

Signature Date

Signature Date

Signature Date

Data Sheet *FSFIC*

Food Sample & Brand Name	Mass of Food Sample Burned (grams)	Mass of Water in Calorimeter (grams)	Initial Water Temperature (°C)	Final Water Temperature (°C)	Measured Calorie Content of Food (**Calories/gram**)	Calorie Content Reported on Label (**Calories/gram**)

Thermochemistry

Developed by: K. Thomas, A. Knudson & J. Ta.

Teacher Material

Synopsis The purpose of this inquiry-based lab/lesson is to have students discover endothermic and exothermic reactions by determining if calcium chloride or ammonium nitrate would be good ingredients for hot packs or cold packs.

Background

From this lab, students should develop an understanding of exothermic and endothermic reactions and know that chemical reactions can be classified as endothermic or exothermic.

Exothermic reactions are reactions that transfer energy to the surroundings. The energy is usually transferred as heat energy, causing the reaction mixture and its surroundings to get hotter. So, a reaction in which heat is given off is called exothermic. In an exothermic reaction the total energy absorbed in bond breaking is less than the total energy released in bond making. Exothermic reactions can be written as: Reactants → Products + Energy. Exothermic processes release heat and will feel hot. In this lesson, we will mix calcium chloride with water to observe an exothermic reaction.

Endothermic reactions are reactions that take in energy from the surroundings. The energy is usually transferred as heat energy, causing the reaction mixture and its surroundings to get colder. So, if heat is absorbed in a reaction, the reaction is said to be endothermic. An endothermic reaction is one that requires more **heat** to break the bonds of the **reactants** than is gained in new bonds of the products. In other words, the reaction absorbs heat, causing its surroundings to get colder. Endothermic reactions can be written as: Reactants + Energy → Products. Endothermic processes absorb heat and will feel cold. In this lesson, we will mix ammonium nitrate with water to observe an endothermic reaction.

Objectives

1. Students will be introduced to hot and cold packs.
2. Students will carry out two reactions, one that is endothermic and one that is exothermic.
3. Students will become more familiar with the scientific method; students will make observations and hypotheses, follow a procedure, collect and record data, graph data, and make conclusions.
4. Students will develop writing skills by completing the required sections for each experiment.
5. Students will develop graphing skills by graphing their data and interpreting the graph.

Suggested Timeline

Day one. Introduction and student experiments. 50-60 minutes.

Day two. As this lesson is designed only to introduce the students to endothermic and exothermic reactions, it should be followed by a detailed discussion of endothermic and exothermic reactions that covers the details of the standard.

Materials (for each group)

Ammonium nitrate (NH_4NO_3)	Goggles	Stopwatch
Aprons (optional)	Paper cups (for chemicals)	Styrofoam cups
Calcium chloride ($CaCl_2$)	Paper towels	Thermometer
Gloves (optional)	Stirring stick	Water

You will also need a beaker (or two) that the students can use to measure 100mL of water and a balance that you can use to weigh the chemicals (see below). Have enough hot and cold packs on hand to show every class. Hot packs will last several hours while cold packs will last only an hour or so. Most of the cold packs you buy in stores consist of ammonium nitrate and water. Hot packs, however, are usually air activated rather than a mixture of CaCl2 and water. Nonetheless, they are an example of an exothermic reaction.

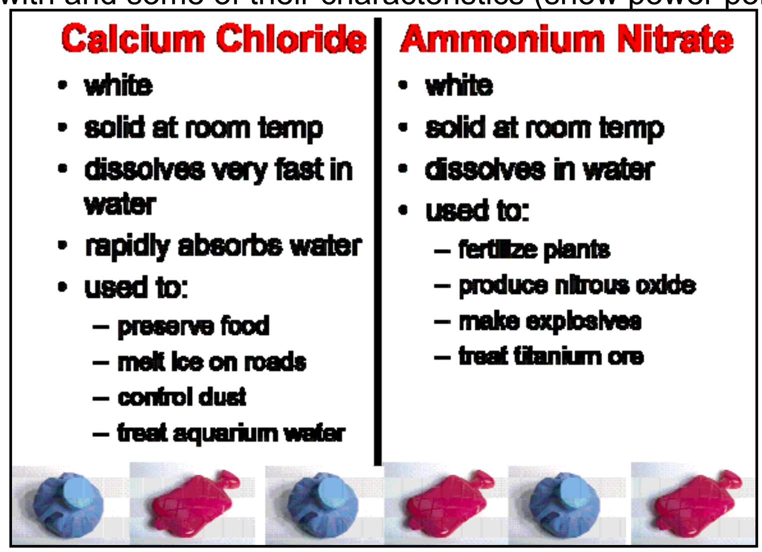
For each reaction, the students will use 100mL of water. For the chemicals, they will need approximately 40g of calcium chloride and 30g of ammonium nitrate (a slightly higher or lower concentration will still produce the desired effect). While we prepared the chemicals for them, you could have the students weigh the chemicals themselves.

Teaching Tips

Outline for activity:

- Hold up and introduce hot and cold packs. Ask the class:
  - What are they used for?
  - What do they do?
  - How do they work?
- Start the reactions and pass the packs around the room- tell students they get three seconds to hold a pack before they pass it along.

- Tell the class: in today's activity, we are going to discover if two chemicals would be good ingredients in a hot or cold pack. Ask the following questions and write their responses to the last three on the board under the titles "materials", "safety", and "data".
    - o How can we find this out? (mix with water)
    - o How will you know if a chemical is a good chemical to use in a hot pack? In a cold pack? (gets hot or cold)
    - o What materials do you need? (water, thermometer, container, …)
    - o What about materials for safety? (goggles, gloves, aprons)
    - o What kind of data should you record? (temperature, time)
- Tell the class: first, let's make some predictions. Here are the two chemicals we have to work with and some of their characteristics (show power point slide):



- Pass out the handouts. Have the students use the information presented on the slide as their observations. Tell them to write down what they think are important observations about these chemicals- observations that will help them make predictions that will answer the questions.
- Tell the class: based on what you know about these chemicals, make some predictions (have them write their hypotheses):
    - o What do you think will happen when you mix Calcium Chloride with water?
    - o What do you think will happen when you mix Ammonium Nitrate with water?
    - o Will either of these chemicals be good candidates for an ingredient in a hot pack? In a cold pack? Why?
- Once they have completed the observations and hypotheses sections, they can start the experiments. Review with them the protocol and handout the role cards. You may need to review Celsius vs. Fahrenheit and show them how to read the thermometers. They should completely finish one reaction and clean up before they start the next.

Roles

We created role cards for each group so that each student had responsibilities and so they could refer to the cards during class. The roles were as follows:

1. Get water and record data:
   Your job is to get 100ml of water in your Styrofoam cup and bring it back to the group. Then, when your group starts recording the temperature, you will write the data down. The last reading will be at 3:00.
2. Get chemicals and timer:
   Your job is to get the chemicals (one at a time) from your teacher (in small paper cup) and take them back to your group. As soon as the chemical is added to the water, start the stop watch and say "read" every 15 seconds. This will let your group know that it is time to read the temperature. When the stop watch reaches 3:00, say "read" and then tell your group that you have reached three minutes.
3. Mixer and cup holder:
   After your group has measured the temperature and everyone is ready, you will add the chemical to the water and stir continuously with the stir stick. It is also your job to hold the Styrofoam cup to make sure it doesn't tip over (very important!).
4. Thermometer reader and holder:
   It is your job to read the initial temperature of the water (in Celsius) and then to read the temperature every 15 seconds when directed by your team. As you do this, make sure you hold the thermometer in the liquid. The last reading will be at 3:00.

Clean up:
After each experiment, they should dump the liquids into separate, marked basins (alternatively, the solutions can be flushed down a sink with plenty of water), throw the cups in the trash, and wipe the thermometer off with the paper towel.

Homework (or in class, if time):

1. Graph how the temperature changed for each chemical. On your graph, the x-axis will be time and the y-axis will be temperature. Make sure you label your axes and give your graph a title.
2. Answer the conclusion questions on your handout.

(STUDENT HANDOUT BEGINS)

THERMOchemistry

STEP 1: QUESTIONS

1. Is calcium chloride ($CaCl_2$) a good candidate (choice) for an ingredient in a cold pack? Is it a good candidate for an ingredient in a hot pack?
2. Is ammonium nitrate (NH_4NO_3) a good candidate (choice) for an ingredient in a cold pack? Is it a good candidate for an ingredient in a hot pack?

STEP 2: OBSERVATIONS

1. Calcium chloride ($CaCl_2$):
   a.______________________________
   b.______________________________
   c.______________________________

2. Ammonium nitrate (NH_4NO_3):
   a.______________________________
   b.______________________________
   c.______________________________

STEP 3: HYPOTHESES

1. What do you think will happen when you mix calcium chloride ($CaCl_2$) with water? Why?______________________________
   ____
   ____
   ____

2. What do you think will happen when you mix ammonium nitrate (NH_4NO_3) with water? Why?
   ____
   ____
   ____
   ____

3. Will either of these chemicals be good candidates for an ingredient in a hot pack? Why or why not? ______________________________
   ____
   ____
   ____

4. Will either of these chemicals be good candidates for an ingredient in a cold pack? Why or why not? ______________________________
   ____
   ____

STEP 4: MATERIALS AND PROCEDURE

MATERIALS:

Ammonium nitrate (NH_4NO_3)
Aprons (optional)
Calcium chloride ($CaCl_2$)
Gloves (optional)
Goggles
Paper cups (for chemicals)
Paper towels
Stirring stick
Stopwatch
Styrofoam cups
Thermometer
Water

PROCEDURE:

1. Using a beaker, measure 100 mL of water into a Styrofoam cup.
2. Measure the temperature of the water and record it in your data table.
3. Using a small paper cup, obtain approximately 40 g of calcium chloride ($CaCl_2$) from your teacher.
4. Add the chemical to the water and start the stop watch immediately. **Be sure to stir.**
5. **Record the temperature every 15 seconds** until you reach 3 minutes.
6. Dispose the solution in the **correct** labeled bin and throw the cups away.
7. Wipe the thermometer off with a paper towel.
8. Repeat steps 1-7 using 30 g of ammonium nitrate (NH_4NO_3).

STEP 5: DATA COLLECTION

1. For **each** reaction, record the temperature of the water **every 15 seconds** in the table below.
2. Graph your data. Label the horizontal x-axis “time” and the vertical y-axis “temperature”. You will plot one line for calcium chloride ($CaCl_2$) and one line for ammonium nitrate (NH_4NO_3). Be sure to title your graph and make a key.

calcium chloride ($CaCl_2$)		ammonium nitrate (NH_4NO_3)	
Time (min:sec)	**Temperature (˚C)**	**Time (min:sec)**	**Temperature (˚C)**
0		0	
0:15		0:15	
0:30		0:30	
0:45		0:45	
1:00		1:00	
1:15		1:15	
1:30		1:30	
1:45		1:45	
2:00		2:00	
2:15		2:15	
2:30		2:30	
2:45		2:45	
3:00		3:00	

STEP 6: CONCLUSIONS

1. Using your graph, summarize what happened when you added calcium chloride ($CaCl_2$) to water. ____
____
____
____

2. Using your graph, summarize what happened when you added ammonium nitrate (NH_4NO_3) to water. ____
____
____
____

3. a. Which of the chemicals would be a good candidate for a hot pack? Why? _____
____
____
____

   b. Does this match your hypothesis? __________

4. a. Which of the chemicals would be a good candidate for a cold pack? Why? _____
____
____
____

   b. Does this match your hypothesis? __________

5. Chemicals can store and release energy in the form of heat. A chemical reaction that releases heat (and becomes hot) is called an ***exothermic*** reaction. But chemical reactions can also absorb heat from the environment and become cold. These reactions are called ***endothermic*** reactions. When chemicals are dissolved in water, sometimes heat is released and sometimes heat is absorbed.

   a. How do you know if a reaction is *exothermic*? __________________________
   ____
   ____

   b. How do you know if a reaction is *endothermic*? _________________________
   ____
   ____

6. Which of these reactions is *exothermic*? ______________________________

7. Which of these reactions is *endothermic*? _____________________________

(STUDENT HANDOUT ENDS)