# Calorimetry - a method for measuring the energy in the form of heat from a chemical process.

When some chemical reactions occur the temperature of the *surroundings* increases. This chemical process is *Exothermic*, releasing energy in the form of heat because the products of the reaction have less energy than the reactants did. *Endothermic* processes require energy so they take it from the surroundings causing the temperature to drop. Energy is absorbed to break bonds in the reactants and energy is released when new bonds form in the products. The difference between bond energies of products and reactants accounts for the energy change in the chemical process.

*Enthalpy* change for a chemical process corresponds to the difference in heat energy of the products and reactants. Watch the Crash Course Chemistry #18 episode on YouTube to learn about enthalpy *<https://youtu.be/SV7U4yAXL5I>*.

A *calorimeter* is a device that isolates the heat involved in the chemical reaction (which is our system being studied) and the immediate surroundings of the reaction (in our case this will be the solvent water) from the rest of the universe. If our calorimeter provided perfect insulation, no heat energy could leak out of it or into it. This means that all of the energy produced in an exothermic reaction would be absorbed by the water solvent causing its temperature to go up. Since the system is losing the energy this is assigned a negative sign. An endothermic reaction absorbs all of the energy it needs to proceed from the surrounding water solvent causing the temperature of the water to go down. Since the system gains the energy it is given a positive sign.

You can make a decent calorimeter from two nested insulated coffee cups with a sponge or corrugated cardboard lid (from a box) with a hole cut in for the thermometer and stirrer (wire or plastic coffee stirrer). I made a very nice one with an insulated coffee cup fit in a coozie (insulated sleeve for soda cans). I used the a sponge lid that fits loosely over the coffee cup and stir by swirling. Do not use an air tight lid because gas is generated in two of the reactions.

This experiment is designed for high school science students. Younger students can do the measurement with adult help and skip the calculations.

Ready to begin the experiment?

Equipment required:

“Coffee Cup” Calorimeter, see Figure 7.7 for a diagram <http://www.brainkart.com/article/Measurement-of---H-using-coffee-cup-calorimeter_34794/>

Thermometer (should be filled with red liquid not mercury for safety)

Graduated cylinder (50 or 100 mL size)

Balance measuring to nearest 0.1 gram at least.

Second timer

Supplies needed:

Safety equipment: safety eye goggles and gloves (nitrile or latex)

Water

Vinegar

Baking soda

Yeast (the kind used for baking, active dry yeast) FYI for use as catalyst

3% Hydrogen peroxide solution

Milk of Magnesia (liquid form containing magnesium hydroxide)

computer for graphing and keeping experiment file

Do BLUE - when you see a blue instruction be sure to put it in your lab notebook.

1. Draw a labeled diagram of your calorimeter.
2. Produce a table entitled “Heat Loss Data for Calorimeter - temperature versus time”. Time every 10 seconds to 120 seconds, then 150 s, 180 s, 210 s, 240 s, 300 s, and 360 s.

# Part 1: Rate of Heat Loss for your calorimeter

1. Heat 100 mL of water to between 80 and 100 degrees Celsius. If the water is boiling, wait for it to stop bubbling before using.
2. Pour the hot water in your calorimeter, insert the thermometer so that it touches the liquid, and cover with the lid.
3. Take an initial temperature (time equals zero) reading as you begin the timer. Then record data in your data table that you created in your laboratory notebook.

3. Produce a graph entitled “Heat Loss Rate Graph for Calorimeter, temperature in degrees Celsius versus time in seconds”. Determine the best straight line for your data and show the calculation for the slope, which equals the rate in degrees Celsius per second.

# Part 2: Measuring the Enthalpy Change for the decomposition of hydrogen peroxide.

Hydrogen peroxide (dihydrogen dioxide) decomposes very slowly under normal temperature and pressure conditions to produce water liquid and oxygen gas. It would be impossible to measure the enthalpy change at the slow pace of the reaction because the rate of heat loss from your calorimeter would be faster than the rate of the reaction. Luckily we can speed up the rate of the reaction by adding a catalyst, activated dry yeast. The catalyst changes the activation energy for the reaction, but not the enthalpy change. <https://chem.libretexts.org/Courses/Valley_City_State_University/Chem_122/Chapter_4%3A_Chemical_Kinetics/4.6%3A_Catalysis>

4. Produce a table entitled “Enthalpy Change Data using Calorimeter, temperature versus time”. Time every 10 seconds to 120 seconds, then 150 s, 180 s, 210 s, 240 s, 270 s, and 300 s.

1. Weigh the calorimeter cup empty. Place 100 mL of 3.0% hydrogen peroxide solution in the calorimeter cup and reweigh. Determine the mass of 3% hydrogen peroxide solution used.
2. Insert the thermometer so that it touches the liquid, and cover with the lid. Measure the initial temperature.
3. Add packet or one tablespoon of dried activated yeast and begin timing. Take temperature readings then record data in your data table that you created in your laboratory notebook. Keeping the bottom of the calorimeter on the counter top you can gently swirl in circles to mix or use the stirrer. Do not splash!

5. Produce a graph entitled “Enthalpy Change Data using Calorimeter, temperature in degrees Celsius versus time in seconds”. Draw the best smooth curve with your data points. What is the peak temperature? Why is that temperature not as high as it should be? Read off your Heat Loss graph how much heat would have been lost in the time it took to get your peak temperature.

final temperature = peak temperature + |degrees Celsius heat lost|

6. Calculate the Enthalpy Change in kilojoules per mole of hydrogen peroxide reacted\*. \*See directions for calculations at the end of the experimental section if needed.

# Part 3: Measuring the Enthalpy Change for the reaction of acetic acid with sodium bicarbonate.

Acetic acid in a 5% solution in water is vinegar. Baking soda is pure sodium bicarbonate. They react to produce aqueous sodium acetate, water liquid and carbon dioxide gas. The enthalpy change for the reaction will be calculated in Kilojoules per mole of sodium bicarbonate, our limiting reagent. The rate of heat exchange through your calorimeter will be very small because the rate of the reaction is very fast. In this experiment, you will skip the graphing step and just measure the temperature at the time in seconds that the temperature reaches the lowest point.

1. Weigh the calorimeter cup empty. Place 100 mL of 5% acetic acid solution (aka vinegar) in the calorimeter cup and reweigh. Determine the mass of vinegar used. This should contain about 0.08 moles of acetic acid. Can you do the calculation?
2. Insert the thermometer so that it touches the liquid, and cover with the lid. Measure the initial temperature.
3. Add between 0.300 and 0.400 grams of baking soda (be sure to record the actual grams used in your notebook) and begin timing. Take the minimum temperature reading and time that occurred, then record the data in your laboratory notebook or computer file. Keeping the bottom of the calorimeter on the counter top you can gently swirl in circles to mix or use the stirrer. Do not splash!

7. Calculate the Enthalpy Change in kilojoules per mole of sodium bicarbonate reacted\*. \*See directions for calculations at the end of the experimental section if needed.

# Part 4: Measuring the Enthalpy Change for the reaction of acetic acid with magnesium hydroxide.

Acetic acid in a 5% solution in water is vinegar. Milk of Magnesia is a suspension of magnesium hydroxide in water and ethylene glycol and some non reactive flavoring/coloring compounds. Acetic acid and magnesium hydroxide react to produce aqueous magnesium acetate and water liquid. The enthalpy change for the reaction will be calculated in Kilojoules per mole magnesium hydroxide, our limiting reagent. Is the rate of heat transfer from the reaction to the water fast enough that the heat loss is less than one degree Celsius? No, the rate of the reaction is slowed by the low water solubility of magnesium hydroxide in water. The correction for heat loss is needed to find the final temperature. In this experiment, you will skip the graphing step and just measure the highest temperature and the time in seconds that the temperature decreases after it reaches the highest point.

1. Weigh the calorimeter cup empty. Place 100 mL of 5% acetic acid solution (vinegar) in the calorimeter cup and reweigh. Determine the temperature of the vinegar, this is the initial temperature.

2. Shake the Milk of Magnesia bottle very well, inverting at least 5 times.

3. Measure out 30 mL of the Milk of Magnesia (2.4 g magnesium hydroxide should be in this, check the label).

4. Add Milk of Magnesia to the vinegar and begin timing. Take temperature readings then record data in your data table that you created in your laboratory notebook. Keeping the bottom of the calorimeter on the counter top you can gently swirl in circles to mix or use the stirrer. Do not splash! Keep collecting data until the temperature goes down from the peak temperature.

8. Use your Heat Loss slope to correct for how much heat would have been lost in the time it took for the temperature to decrease from your peak temperature.

final temperature = peak temperature + |degrees Celsius heat lost|

Calculate the Enthalpy Change in kilojoules per mole of magnesium hydroxide reacted\*. \*See directions for calculations at the end of the experimental section if needed.

# Calculations

The basic premise behind calorimetry calculations is that you have isolated your system, which is the chemical reaction, and surroundings, which is the water solvent, from the rest of the universe. Obviously, with our coffee cup calorimeter we do not have perfect blocking of energy flow to and from the universe. Still our equipment is good enough to measure values to 2 or 3 significant figures of accuracy. Since all heat transfer is assumed to occur between the system and the surroundings, q is used as the symbol for heat.

q for water = (4.18 J/g degrees C) x g of water x temperature change

Where: Specific Heat for water solution = (4.18 J/g degrees C)

Mass of water in grams\*

Temperature Change = final - initial temperature

TIP The sign of the temperature change matters!

\*in a 3% peroxide solution there are 3 g peroxide and 97 grams of water to make 100 grams of solution.

q for reaction = - q for the water

Because the same amount of energy is transferred, but one is losing and the other is gaining.

The heat from the reaction is in Joules/gram. This need to be converted to kJ/ mole of the limiting reagent. Conversion factor: 1 kJ = 1000 J

The limiting reagent in each reaction is given in the calculation question. Convert the grams to moles, then divide the kJ by the moles. Delta H for the reaction is the Enthalpy change in kJ per mole times the coefficient of this reagent in the balanced chemical equation.

If you a high school or college chemistry student try these:

9. Using the terms *system and surroundings,* describe the energy flow inside the calorimeter for: a) an endothermic reaction and b) an exothermic reaction. Explain the sign of delta H for each.

10. Write the balanced chemical equation for the decomposition of hydrogen peroxide (a catalyst is not included in the reaction equation). Write your value of delta H at the end. Compare this to what you calculate using the Heat of formation data under standard conditions...for aqueous hydrogen peroxide -191.0 kJ/mole, water liquid -285.8 kJ/mole, and oxygen gas 0.0 kJ/mole.