

LAD C.2 (pg 1 of 2) Enthalpy Change of a Reaction Using Hess' Law

Introduction:

Thermochemistry and thermodynamics is the branch of chemistry that studies energy and its transformations. The First Law of Thermodynamics states that the energy of the universe is a finite quantity and that energy can be neither created nor destroyed, only transformed. An object can contain energy in two forms

- kinetic energy - energy of motion (directly proportional to temperature)
- potential energy - energy stored because of position (electrostatic attractions and repulsions, bond energy)

Energy can be transferred in two different forms

- work (w) - defined as force over distance. In chemistry, the work performed is usually the work done by or done on gases in the form of volume changes, gases expanding or compressing. $w = P\Delta V$. But in many cases as with this experiment, the reaction occurs in solution, and there are no pressure or Δ in gas volume considerations and thus no measurable work is done.
- heat (q) - measured temperature changes used in combination with $q = c * m * \Delta T$ will be used to calculate the amount of heat transferred during a reaction. (c = specific heat capacity, m = mass, ΔT = temp change)
- $\Delta E_{\text{total}} = q + w$

Heat measured under constant pressure conditions (as is the case in this experiment - reactions that occur in solution in an open container are under the constant pressure of the open atmosphere) is called enthalpy, H . In many reaction situations the ΔE is nearly equal to ΔH and these quantities are often used interchangeably.

The sign of ΔH is important

- ΔH will be negative (exothermic) when heat is lost by a system
- ΔH will be positive (endothermic) when heat is gained by a system.

PreLAD:

Data Table - On a separate piece of paper, not on this LAD sheet.

Read the Procedure Overview, Procedure and Processing the Data and set up a data/results table. Put your data entries in rows and make two columns for two trials (in the event, a second trial is needed). Leave a third column for calculations.

Write balanced net ionic equations for the following reactions

1. One mole of magnesium oxide is formed by burning a strip of magnesium in oxygen.
2. A strip of magnesium reacts with hydrochloric acid
3. Magnesium oxide powder is reacted with hydrochloric acid
4. Write the formation reaction for liquid water, and look up the ΔH°_f value.

Answer the following questions, showing work as appropriate

1. What are the typical units for c , specific heat capacity?
2. What are the typical units for ΔH_{rx} ?

Procedure Overview:

The enthalpy of the first reaction written above is very difficult to measure directly with any accuracy. Three different reactions can be manipulated using Hess' Law to determine the enthalpy for the first reaction, the combustion of magnesium. Two of enthalpy values for the reactions needed in the Hess' Law manipulation will be measured during the experiment, the third value for the formation of water will be acquired from the formation tables.

Materials on tray for two groups

- 2x foam cups
- 2x thermometer
- 2x Mg strip
- 2x vial of magnesium oxide powder
- 2x 100 ml graduated cylinder
- 3M hydrochloric acid solution
- 2x stirring bar and stirrer
- 500 ml 3 M HCl

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Procedure - *Protective eyewear is not optional.*

Reaction 1 - All waste solutions can go down the sink. Hang on to the stirring bar.

- Measure the mass of ~100 ml of 3 M HCl into a foam cup. Place the stirring bar in and measure an initial temperature.
- Measure the mass of magnesium strip. Fold it loosely into a shape that will completely submerge and fit on the bottom of the cup. Turn the stirrer on slow speed.
- Drop in the magnesium and record the highest temperature.
- Dispose of the solution down the sink, and rinse the calorimeter with plenty of tap water.

Reaction 2 - All waste solutions can go down the sink. Hang on to the stirring bar.

- Measure the mass of a fresh ~100 ml of 3 M HCl into a foam cup. Place the stirring bar in and measure an initial temperature.
- Measure the mass of the vial with MgO powder (without cover). Turn the stirrer on slow speed.
- Pour the MgO into the acid and record the highest temperature. Measure the mass of the empty vial.
- Dispose of the solution down the sink, and rinse the calorimeter with plenty of tap water.

Processing the data:

Consider the specific heat capacity (c) of 3 M HCl to be 4.05 J/g°C

- Using your lab data, calculate q for reaction 1 & 2
- Calculate the number of moles of Mg (MM 24.31 g/mole) and MgO (MM 40.31g/mole) used.
- Using your q from #1 and mole values in # 2, calculate experimental ΔH for reaction 1 & 2
- With ΔH values for reactions 1 & 2 in combination with the ΔH_f° value for the formation of water (which you should look up on your thermodynamic table), you can use Hess' law to calculate the experimental ΔH_f° for MgO. Show clearly how the reactions add and their enthalpy values add to result in the desired reaction and ΔH value.
- Look up the theoretical ΔH_f° for MgO and calculate a % error for your experimental value.

POST LAD QUESTIONS - To be completed on the back of your Data/Results Table. Show all work as necessary. This will be turned in on the second day following the completion of the LAD. Staple your own lab papers and then paperclip with partners papers with with the clearest presentation on the top. (Put all partners names on all labs - circle the author.) If you have questions about this LAD they should be asked before the LAD is due (it would be wise to ask during the LAD class or in class on the day following the LAD.)

- For reaction 1, if 6 M HCl had been used, how would the measurements and the resulting calculation of q have been different? How would ΔH for reaction 1 be affected?
- For reaction 1, if twice the mass of Mg had been used, how would the measurements and the resulting calculation of q have been different? How would ΔH for reaction 1 be affected?
- You saw a demonstration of the combustion of Mg in class. What problems would be encountered trying to directly determine the ΔH for this reaction by doing the reaction itself (instead of employing Hess' Law.)?
- What major assumption(s) is(are) made in this experiment?
- We did not cover the calorimeter. How would the calculation of the ΔH_f° of MgO be affected if a large amount of heat were lost in reactions 1 & 2