

# Chemistry 212

## THE ENTHALPY OF FORMATION OF MAGNESIUM OXIDE

### LEARNING OBJECTIVES

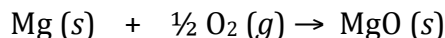
The learning objectives of this experiment are

- A simple coffee cup calorimeter will be used to determine the enthalpy of formation of magnesium oxide.
- Hess's Law will be used in the determination.

### BACKGROUND

The heat (energy) absorbed in the formation of one mole of a compound in its standard state from its elements in their standard states is referred to as the standard enthalpy of formation ( $\Delta H$ ) of the compound.

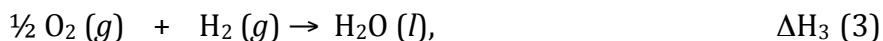
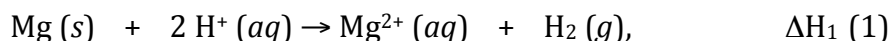
For example, the standard enthalpy of formation of magnesium oxide is the energy transferred in the formation of one mole of magnesium oxide from its elements in their standard states.



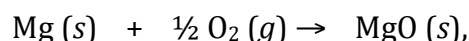
The enthalpy of the above reaction can be determined by measuring the energy transferred when a known quantity of magnesium metal is burned in the presence of excess oxygen. Direct measurement requires the use of complex equipment and is therefore not practical in the first year laboratory.

This difficulty will be circumvented by using Hess's Law, which states that if a reaction can be regarded as the sum of two or more reactions, the change in enthalpy for the overall reaction will be equal to the sum of the enthalpy changes for the component reactions.

Consider the following reactions:



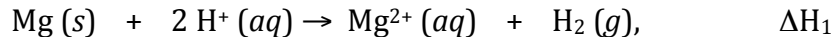
The sum of these three reactions produces:



therefore, the desired enthalpy of formation of magnesium oxide is equal to:

$$\Delta H_{\text{f (MgO)}} = \Delta H_1 + \Delta H_2 + \Delta H_3$$

In this experiment,  $\Delta H_1$  will be determined by measuring the energy released when a known quantity of magnesium metal is reacted with a slight excess of hydrochloric acid.



The reverse of  $\text{H}_2$  will be determined by measuring the energy released when magnesium oxide is dissolved in hydrochloric acid.



The heat of formation of water,  $\Delta H_3$ , is readily available in scientific literature.

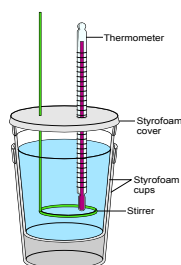
Reactions (1) and (2) will be completed in aqueous solution in the lab. An energy transfer will occur between the reacting species and the solution in which the reaction takes place. This energy transfer will cause a resulting increase (or decrease) in the temperature of the solution. The assumption is made that all energy transfer is in the form of heat. Also, assume that the **density of the HCl solution is 1.00 g/mL**.

The **specific heat capacity of water is 4.184 J C<sup>-1</sup>g<sup>-1</sup>**. It takes 4.184 J (or 1 calorie) of energy to raise the temperature of one gram of water one degree Celsius.

By measuring the change in temperature of a given mass of water as a known quantity of material is reacted, it is possible to calculate the energy transferred during the reaction of that material. The energy which would be released by the reaction of one mole of the material (and thus the change in enthalpy for the reaction) can then be calculated.

One of the problems which is encountered in measuring temperature changes is that while the temperature is still rising, energy is being lost to the surroundings (by radiation and convection), and the observed temperature change will be smaller than the "true" temperature change. The energy loss is minimized by completing the reaction under study in a calorimeter system such as the one illustrated below. The loss of energy to the surroundings cannot be completely eliminated, however, and for this reason a graphical temperature correction is employed.

After the reactants have been mixed, the temperature of the reacting solution is plotted as a function of time. By extrapolating the slope of the temperature/time curve as the solution cools to time zero,  $T_i$ , (the instant of mixing of the reactants) the true maximum temperature,  $T_f$ , and thus the true temperature change,  $T_f - T_i$ , of the reacting solution can



be determined.

Once the true  $\Delta T$  is known, the energy transferred during the reaction can be calculated using the specific heat capacity of water, the mass of solution in the calorimeter, and the temperature change of the solution:

$$\text{Energy transferred (Joules)} = 4.184 \text{ J C}^{-1}\text{g}^{-1} \times \text{g of solution} \times \Delta T$$

## EXPERIMENTAL PROCEDURE

A device which is used to measure the energy evolved or absorbed during a process is called a calorimeter. In this experiment a calorimeter will be constructed from two Styrofoam cups in a fashion similar to that shown here. Stirring can be done occasionally by gently swirling the calorimeter.

### Part A. Determination of $\Delta H_1$

1. Assemble the calorimeter as shown in the diagram (replacing the thermometer with a digital temperature probe). Leave the top and thermistor off of the calorimeter at this point.
2. Measure 60 mL of 1 M HCl into a graduated cylinder and record the volume to the nearest 0.1 mL and add to calorimeter.
4. Dry the probe and place it in the acid solution.
5. Start collecting temperature data.
6. Quickly transfer 0.550 - 0.600 g (make sure to record the exact weight) of magnesium turnings into the calorimeter. It will fizz vigorously, so be careful.
7. When the above procedure is complete, clean and dry the styrofoam cup and repeat the procedure for a second trial.

### Part B. Determination of $\Delta H_2$

1. Using the same procedure as outlined above, determine  $\Delta H_2$ .
2. You should use 0.900 - 1.000 g of MgO and 60 mL of 1 M HCl.
3. Two trials should be completed.
4. The calorimeter should be cleaned and dried between each trial.

## Data Sheet

Part A. **Determination of  $\Delta H_1$**      $\text{Mg} (s) + 2 \text{H}^+ (aq) \rightarrow \text{Mg}^{2+} (aq) + \text{H}_2 (g)$

Quantity	Trial 1	Trial 2
Mass of Mg (g)		
Moles of Mg (mol)		
Mass of solution + Mg (g)		
$\Delta T$ ( $^{\circ}\text{C}$ )		
Energy transferred (kJ)		
$\Delta H_1$ (kJ/mol)		

Part B. **Determination of  $\Delta H_2$**      $\text{Mg}^{2+} (aq) + \text{H}_2\text{O} (l) \rightarrow \text{MgO} (s) + 2 \text{H}^+ (aq)$

Quantity	Trial 1	Trial 2
Mass of MgO (g)		
Moles of MgO (mol)		
Mass of solution + MgO (g)		
$\Delta T$ ( $^{\circ}\text{C}$ )		
Energy transferred (kJ)		
$\Delta H_2$ (kJ/mol)		

**Determination of  $\Delta H_f$**      $\text{Mg} (s) + \frac{1}{2} \text{O}_2 (g) \rightarrow \text{MgO} (s)$

Mean $\Delta H_1$ (Part A)	
Mean $\Delta H_2$ (Part B)	
$\Delta H_3$ ( $\Delta H_f, \text{H}_2\text{O} (l)$ ) (from literature)	
$\Delta H_f, \text{MgO} (s)$ (experimental)	
$\Delta H_f, \text{MgO} (s)$ (from literature)	
% error (of $\Delta H_f, \text{MgO} (s)$ )	

Write out the Hess's Law calculation and discuss the accuracy of your experimental result.

# Chemistry 212

## Pre Laboratory Questions

### THE ENTHALPY OF FORMATION OF MAGNESIUM OXIDE

- For the reaction,  $\text{Mg (s)} + 2 \text{H}^+ (\text{aq}) \rightarrow \text{Mg}^{2+} (\text{aq}) + \text{H}_2 (\text{g})$ , if  $\Delta H_1 < 0$ :
  - Is the reaction endothermic or exothermic?
  - Would the solution in the calorimeter increase or decrease in temperature?  
Explain
- If 60.0 mL of 1 M HCl and 0.560 g of Mg(s) is added to the calorimeter, what would be the mass of the solution in the calorimeter?
- Explain one possible problem with our apparatus that has the potential to cause a systematic error.
  - How would this effect the calculation of energy transferred?