

CHAPTER 10	Investigation 10.A: Hess's Law and the Enthalpy of Combustion of Magnesium Answer Key	BLM 10.1.4A
ANSWER KEY		

Answers to Prediction Questions

The molar enthalpy of formation of magnesium oxide (which is also the molar enthalpy of combustion of magnesium) as listed in the appendix is -601.6 kJ/mol . In predicting the enthalpies of reactions (2) and (3), having the “right” answer is not as important as having a reasonable answer with a logical explanation. You could predict that since equation (1) is highly exothermic, equation (3) will probably be exothermic as well. You could predict that equation (2) is exothermic because it is spontaneous (which is not necessarily the case, but a reasonable statement in terms of your current knowledge). You might also predict that since equation (1) is highly exothermic, equation (2) will probably be endothermic (the equation is reversed when equations (2), (3), and (4) are added to give equation (1), thus when reversed an endothermic reaction would be exothermic, increasing the ΔH for equation (1).) In the investigation, you will find that the reactions are both exothermic.

Answers to Analysis Questions

1. You will use your temperature data to determine Δt . Remember, $\Delta t = t_f - t_i$. Since both reactions are exothermic, the value of Δt for both will be positive. You should use $4.19 \text{ J/g}\cdot^\circ\text{C}$ for c . If you did not measure the mass of the solutions, you should use $m = d \cdot V$ to determine mass, where $d = 1.00 \text{ g/mL}$ (making the assumption that the density of the solution is close to the density of water) and $V = 100.0 \text{ mL}$. You will then determine the enthalpy change using the equations $\Delta H = n\Delta_x H$ and $Q = mc\Delta t$. You should use the equation $n = m/M$ to determine the amounts (in mol) of MgO(s) and Mg(s) used, if you do your calculation in two steps. For example, if you used 0.80 g of MgO(s) in Part 1, the calculation is as follows:

$$\begin{aligned}
 n &= \frac{m}{M} \\
 &= \frac{0.80 \text{ g MgO}}{40.31 \frac{\text{g}}{\text{mol}} \text{ MgO}} \\
 &= 0.020 \text{ mol MgO(s)}
 \end{aligned}$$

Similarly, if you used 0.50 g of Mg(s) in Part 2, the calculation is:

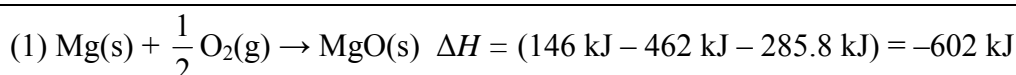
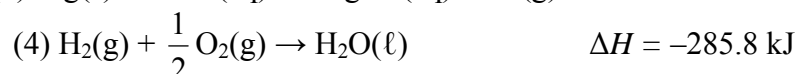
$$\begin{aligned}
 n &= \frac{m}{M} \\
 &= \frac{0.50 \text{ g Mg}}{24.31 \frac{\text{g}}{\text{mol}} \text{ Mg}} \\
 &= 0.021 \text{ mol Mg(s)}
 \end{aligned}$$

Using $\Delta H = n\Delta_x H$ and $Q = mc\Delta t$ and your answers for mole calculations, you can determine the enthalpy change for 1 mol of MgO(s) (equation (2)) or Mg(s) (equation (3)). For equation (2), the literature value for ΔH° is $-146 \text{ kJ/mol MgO(s)}$. For equation (3), the literature value for ΔH° is $-462 \text{ kJ/mol Mg(s)}$.

2. (2) $\text{MgO(s)} + 2\text{HCl(aq)} \rightarrow \text{MgCl}_2\text{(aq)} + \text{H}_2\text{O(l)}$ $\Delta H = -146 \text{ kJ}$
 (3) $\text{Mg(s)} + 2\text{HCl(aq)} \rightarrow \text{MgCl}_2\text{(aq)} + \text{H}_2\text{(g)}$ $\Delta H = -462 \text{ kJ}$

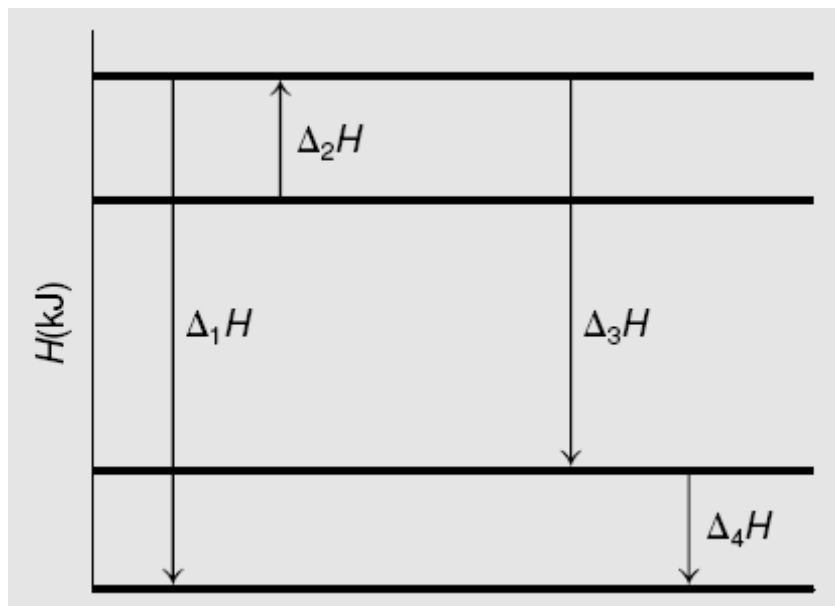
CHAPTER 10	Investigation 10.A: Hess's Law and the Enthalpy of Combustion of Magnesium Answer Key (cont'd)	BLM 10.1.4A
ANSWER KEY		

3. (a) You should show that the reverse of equation (2), plus equations (3) and (4), add to give equation (1). You should use their ΔH values for equations (2) and (3), and ΔH for equation (4) provided in the introduction to the experiment. Ensure that you multiply ΔH for equation (2) by -1 when you reverse the equation.



Your results should have only two significant digits, since the mass of magnesium has only two significant digits.

(b)



4. (a) The accepted value of ΔH° for the combustion of magnesium (also the formation of magnesium oxide) is -601.6 kJ/mol .

$$\text{percent difference} = \left| \frac{\text{experimental value} - \left(-601.6 \frac{\text{kJ}}{\text{mol}} \right)}{-601.6 \frac{\text{kJ}}{\text{mol}}} \right| \times 100\%$$

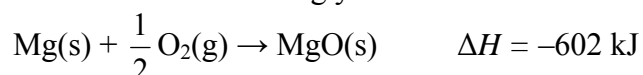
Your results should have only two significant digits, since the mass of magnesium has only two significant digits.

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(b) The major source of error is heat lost from the calorimeter to the surroundings. The heat absorbed by the calorimeter should also be mentioned, and measurement errors in reading the thermometer, the graduated cylinder, and measuring the mass of solids. Relatively small errors include assuming the density of the solution is 1.00 g/mL, and the specific heat capacity of the solution is 4.19 J/g•°C. The fact that conditions are not at SATP is a trivial source of error. To improve the experiment, more trials should be undertaken. As well, now that you know that the reactions are exothermic and the approximate temperature change, you can minimize heat loss or gain to and from the surroundings by starting the investigation as far below room temperature as possible, since the temperature will increase above room temperature.

Answers to Conclusion Questions

5. You should briefly explain that you carried out two reactions and used a calorimeter to determine enthalpy values for thermochemical equations. Using the equation of a reaction with a known ΔH , you were able to manipulate and add those three equations to obtain the target equation and find its ΔH . This was possible because of Hess's Law, which states that the enthalpy change of a chemical process depends only on the beginning and end conditions; ΔH is independent of the pathway. You may reproduce here the solution shown for Analysis Question 3(a). The thermochemical equation should be written using your results for ΔH :



Answers to Application Questions

6. You should design an investigation that is similar to Investigation 10.A. Using a simple calorimeter, you would be able to determine ΔH for equations (1), (2), and (3), and check whether the calculated ΔH values from equations (1) and (3) add up to the empirically determined ΔH for equation (2), verifying Hess's law. For equation (1), you would find the mass of some NaOH(s) pellets, then determine the enthalpy of dissolving of the pellets in an accurately measured volume of water. For equation (2), you would add an accurately measured sample of NaOH(s) pellets to neutralize 0.500 mol/L HCl(aq). The mass of NaOH(s) used should be the stoichiometric quantity to just neutralize the HCl(aq). For equation (3), you would use an accurately measured sample of NaOH(aq) solution to neutralize the HCl(aq). Ensure that you include appropriate safety precautions and detailed procedure steps.