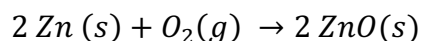


Application of Calorimetry of the Decomposition of Hydrogen Peroxide to Determination of the Enthalpy of a Metallic Oxidation Reaction

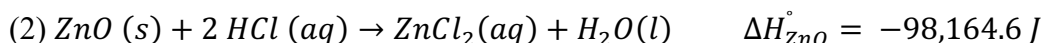
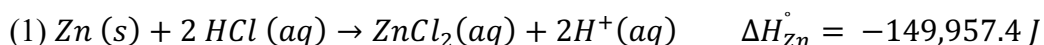
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CHEM 1310L Laboratory

Abstract

A core part of chemistry is the study of how heat plays a role in chemical reactions. Namely, thermochemistry studies the heat exchange between a chemical reaction and its surrounding environment. In the following experiment, we studied the change in enthalpy (ΔH) for the formation of zinc oxide. The traditional chemical reaction for the synthesis of zinc oxide is:



However, conducting this reaction in a laboratory setting is highly impractical. By way of Hess' Law, we can combine the standard enthalpies of reactions involving the individual reactants and products to derive an experimental value for the overall synthesis of zinc oxide. Specifically, we will be measuring the exchange of heat for the reaction of (1) zinc metal with hydrochloric acid and (2) zinc oxide with hydrochloric acid.



When carrying out the experiment, we will utilize the process of calorimetry. In calorimetry, we use a coffee cup with alcohol thermometers to measure initial and final temperature states of the reaction occurring in a bath of 6 M hydrochloric acid. By measuring such states, and from the fact that enthalpy is a state function, we can analyze the thermal changes that occur due to the individual reactions and deduce their overall changes in enthalpy.

Data and Results

Trial	Cold T ($^{\circ}\text{C}$)	Hot T ($^{\circ}\text{C}$)	Final T ($^{\circ}\text{C}$)	C_{cal} ($\text{J}/^{\circ}\text{C}$)	Notes
1	21.5	49.5	34.9	18.7	
2	21.5	61.0	40.0	22.6	
3	21.7	88.0	49.9	73.4	Omitted
Mean C_{cal} :				20.7	
Standard Deviation:				2.8	

Table 1. Calibration data for a coffee-cup calorimeter.

Trial	Initial T ($^{\circ}\text{C}$)	Final T ($^{\circ}\text{C}$)	Mass solid (mL)	Volume HCl (mL)	ΔH (kJ/mol)
1	22.2	23.8	0.102	30.0	-150
2	21.5	24.8	0.400	30.0	-98.2

Table 2. Calorimetry data for the zinc-catalyzed decomposition of aqueous .

Metal	Metal Oxide	ΔH° , measured (kJ/mol)	ΔH° , theoretical (kJ/mol)
Zinc	Zinc Oxide	-337.1	-350.5

Table 3. Theoretical and measured (via Hess's law) enthalpies of oxidation of Zinc by aqueous hydrochloric acid.

Discussion

For part A of this experiment, we calculated the calorimeter constant for our calorimeter setup simply because every calorimeter setup is variable to some degree of error. By including this error term in our overall calculations for the total enthalpy change, we can increase the accuracy of our measurements.

To determine the calorimeter constant, we carried out a simple experiment that mixed hot water with cold water at known temperatures and measured the final temperature of the solution; we were then able to determine the amount of heat lost to the calorimeter walls. We start off with a base equation:

$$q_{hot} = -(q_{cold} + q_{calorimeter})$$

This simplifies to:

$$m * 4.184 * \Delta T_{hot} = -(m * 4.184 * \Delta T_{cold} + C_{cal} \Delta T_{cold})$$

Where we assume that the ΔT of the calorimeter is the same as the cold-water bath since the walls are at the same temperature as the bath water.

$$m * 4.184 * (T_{final} - T_{hot}) = -(m * 4.184 * (T_{final} - T_{cold}) + C_{cal} * (T_{final} - T_{cold}))$$

We are solving for the calorimeter constant; as such, isolating C_{cal} gives us:

$$C_{cal} = \frac{m * 4.184 * (T_{final} - T_{hot}) + m * 4.184 * (T_{final} - T_{cold})}{-(T_{final} - T_{cold})}$$

Substituting values in for trial 1 gives us.

$$C_{cal} = \frac{50.0 * 4.184 * (34.9 - 49.5) + 50.0 * 4.184 * (34.9 - 21.5)}{-(34.9 - 21.5)}$$

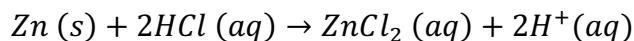
$$C_{cal} = 18.7 \frac{J}{^{\circ}C}$$

We repeat this experiment for three more trials and get our average value of C_{cal} to be $20.7 \frac{J}{^{\circ}C}$.

When using the calorimeter setup to carry out the actual experiment, we use similar equations to model the heat transfer that occurs in the reaction. Instead, we say the heat of the overall reaction is equal to the negative of the heat transferred to the bath and to the walls of the calorimeter.

$$q_{rxn} = -(q_{bath} + q_{cal})$$

For the oxidation of zinc metal in hydrochloric acid, the reaction is modeled as follows:



To determine the over change in enthalpy of the oxidation reaction (ΔH_{ox}) we must determine the amount of heat transferred to the bath of hydrochloric acid and to the calorimeter walls. Substituting values in for the heat equation gives us:

$$q_{rxn} = -(m * c * \Delta T + C_{cal}\Delta T)$$

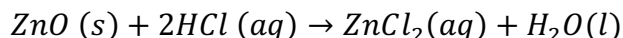
$$q_{rxn} = -(30 * 4.184 * (23.8 - 22.2) + 20.7 * (23.8 - 22.2))$$

$$q_{rxn} = -234.0 J$$

However, the enthalpy for this reaction is just for 0.102 g of zinc. To standardize this, we must find the change in enthalpy for one mole of zinc, also known as the standard enthalpy (ΔH_{ox}°).

$$\Delta H_{ox}^\circ = \frac{\Delta H_{0.102g Zn}}{n_{Zn}} = \frac{-234.0 J}{.102g Zn / 65.38 \frac{g}{mol} Zn} = -149,957.4 \frac{J}{mol} \text{ or } -150.0 \frac{kJ}{mol}$$

Practically the same approach was used to determine the change in enthalpy for the zinc oxide reaction. This time, however, we used 0.4 grams of ZnO in aqueous hydrochloric acid. The equation is as follows:



And calculating the enthalpy of the reaction is similar in that:

$$q_{rxn} = -(m * c * \Delta T + C_{cal}\Delta T)$$

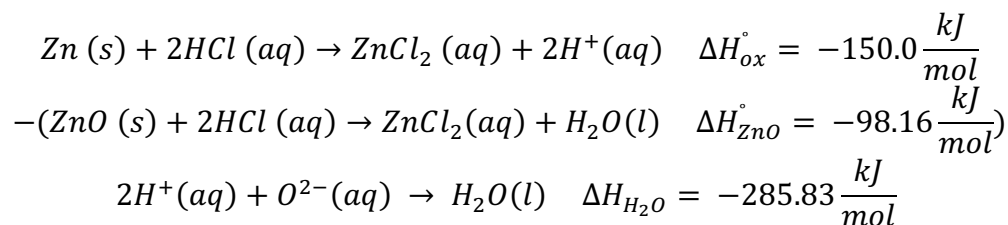
$$q_{rxn} = -(30 * 4.184 * (24.8 - 21.5) + 20.7 * (24.8 - 21.5))$$

$$q_{rxn} = -480.2 J$$

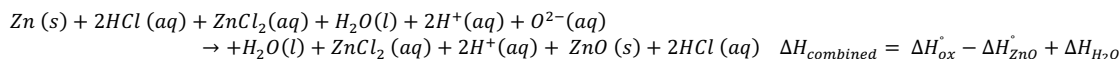
And calculating the enthalpy per mole of reactant was:

$$\Delta H_{ZnO}^\circ = \frac{\Delta H_{0.4g ZnO}}{n_{ZnO}} = \frac{-480.2 J}{.4g ZnO / 81.38 \frac{g}{mol} ZnO} = -98,164.625 \frac{J}{mol} \text{ or } -98.16 \frac{kJ}{mol}$$

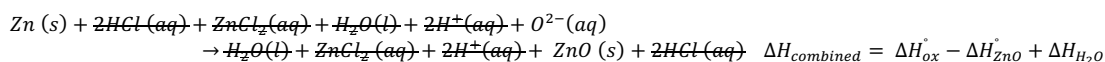
Now, in determining the enthalpy for the synthesis of zinc oxide, we must combine the two reactions by a factor of some scalar. We will take the first reaction of zinc with hydrochloric acid and add it to the reverse of the second reaction, the one with zinc oxide. We also add the reverse reaction for the formation of water in order to cancel elements that do not belong in the intended equation.



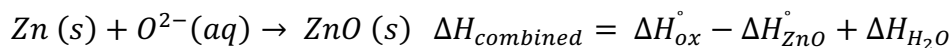
Which equals



We can cancel elements that are repeated on both sides of the equation.



Removing all the canceled molecules reveals:



Where the total change in enthalpy is:

$$\Delta H_{\text{combined}} = -150.0 \frac{\text{kJ}}{\text{mol}} + 98.16 \frac{\text{kJ}}{\text{mol}} - 285.83 \frac{\text{kJ}}{\text{mol}}$$

$$\Delta H_{\text{rxn}} = -337.1 \frac{\text{kJ}}{\text{mol}}$$

In comparison to the theoretical value for the change in enthalpy for the overall reaction

($\Delta H_{\text{theoretical}} = -350.5 \frac{\text{kJ}}{\text{mol}}$), there is a percent error of 3.82% in our own answer of $\Delta H_{\text{rxn}} = -337.1 \frac{\text{kJ}}{\text{mol}}$.