Using Calorimetry to Calculate Molar Enthalpies

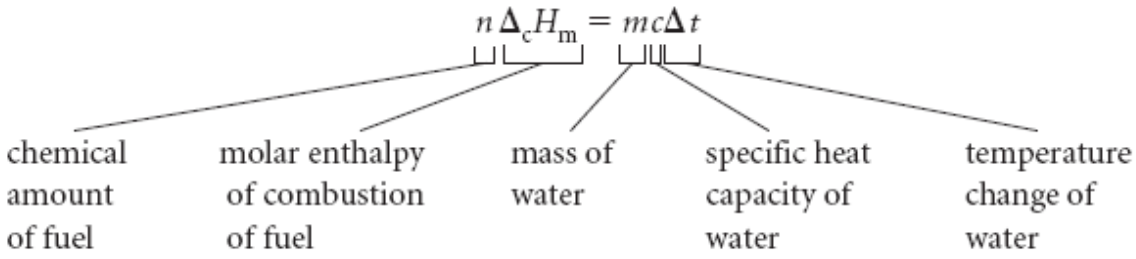
The molar enthalpy of reaction can be calculated from or used in calorimetry investigations.

Again, the law of conservation of energy is used to produce an equality.

The change in enthalpy (ΔH) of the chemical system is equal to the change in thermal energy(Q) of the calorimeter.

$$\Delta H = Q$$

$$n\,\Delta_c H_m = mc\Delta t$$



Ethanol is often added to gasoline as a renewable component that reduces harmful emissions. The mixture is known as gasohol. In a research laboratory, the combustion of 3.50 g of ethanol in a sophisticated calorimeter causes the temperature of 3.63 L of water to increase from 19.88 °C to 26.18 °C. Use this evidence to determine the molar enthalpy of combustion of ethanol.

burning of ethanol → $\Delta_c H = Q$ ← H_2O

$$n\,\Delta_c H_m = mc\Delta t$$

$$\left(\frac{3.50\,g}{46.08\,g\,mol^{-1}}\right)\Delta_c H_m = (3630\,g)(4.19\,J\,g^{-1}\,^\circ C^{-1})(6.30\,^\circ C)$$

$$\Delta_c H_m \; C_2H_5OH = 1261553.357\,J\,mol^{-1}$$

$$= -1.26\,MJ\,mol^{-1}$$

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LAB EXERCISE 11.A

Molar Enthalpy of Neutralization

The molar enthalpy of neutralization, $\Delta_n H_m$, for hydrochloric acid when it reacts with a base can be determined by calorimetry. When two aqueous solutions are mixed in a calorimeter, the final aqueous mixture is considered to be the water in the calorimeter when calculating Q.

Purpose

The purpose of this exercise is to use calorimetry to determine the molar enthalpy of neutralization for hydrochloric acid in its reaction with aqueous ammonia.

Report Checklist

○ Purpose	○ Design	● Analysis
● Problem	● Materials	○ Evaluation
○ Hypothesis	○ Procedure	
○ Prediction	○ Evidence	

Design

1.00 L of 0.100 mol/L hydrochloric acid and 1.10 L of 0.100 mol/L aqueous ammonia are combined in a calorimeter.

Evidence

initial temperature of each solution = 23.20 °C
final temperature of mixture = 23.82 °C
volume of mixture = 2.10 L

0.62°C

$n = cv$

$\Delta_n H = Q$

$n\Delta_n H_m = mc\Delta t$

$(1.00\text{ L} \times 0.100\ \frac{\text{mol}}{\text{L}})\Delta_n H_m = (2100\text{ g})(4.19\ \text{J/g}\cdot^\circ\text{C})(0.62^\circ\text{C})$

$\Delta_n H_m = -5415538\ \text{J/mol}$

$\Delta_n H_{m\ HCl} = -55\ \text{kJ/mol}$

Propane has a molar enthalpy of combustion of -2.04 MJ/mol. 10.0 g of propane is burned in a calorimeter that contains 9.50 L of water at 18.0 °C. What is the final temperature of the water?

$$\Delta_c H = Q$$

$$n\Delta_c H_m = mc\Delta t$$

$$\left(\frac{10.0\text{ g}}{44.11\text{ g mol}^{-1}}\right)\left(2.04\times10^{6}\text{ J mol}^{-1}\right) = (9500\text{ g})(4.19\text{ J g}^{-1}\,^{\circ}\text{C}^{-1})(\Delta t)$$

$$\Delta t = 11.6186\ldots\,^{\circ}\text{C}$$

$$18.0\,^{\circ}\text{C} + 11.61\ldots\,^{\circ}\text{C}$$

$$= \boxed{29.6\,^{\circ}\text{C}}$$

What mass of butane, which has a molar enthalpy of combustion of -4.04×10^3 kJ/mol, would need to burn to heat up 2.00 L of water in a calorimeter by 15.0°C.

$\Delta_c H = Q$

$n\Delta_c H_m = mc\Delta t$

$n(4040000\ J/mol) = (2000\ g)(4.19\ J/g°C)(15.0°C)$

$n = 0.031113\ldots\ mol$

$m = nM = (\ \downarrow\ )(58.14\ g/mol)$

$= 1.808\ldots\ g = \boxed{1.81\ g}$