Chapter 5: Thermochemistry

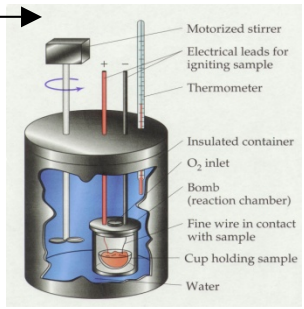
Section 5.5: Calorimetry

Calorimetry: measurement of heat flow Calorimeter: apparatus that measures heat flow

Heat capacity: amount of heat required to raise an object's temperature by 1 K

Molar heat capacity: heat capacity of 1 mol of a substance Specific heat: heat capacity of 1 g of a substance

Bomb calorimeter: →



$$\text{Specific heat} = \frac{q}{m\Delta T}$$

q = quantity of heat transferred
m = grams of substance
ΔT = change in temperature

Example 1: Specific Heat

How many kJ of heat are needed to raise 10.00 kg of water from 24.6°C to 46.2°C? (Specific heat for water is 4.18 J/g-K)

(4.18 J/g-K) = q/(10,000 g)(21.6 K)

$q = 9.03 \times 10^5$ J = 903 kJ

Constant-Pressure Calorimetry:

$q_{soln} = -q_{rxn}$

q_{rxn} = - (specific heat)mΔT

q_{soln} = heat transfer in the solution
q_{rxn} = heat transfer in the reaction

Bomb Calorimetry (Constant-Volume Calorimetry):

$q_{rxn} = -C_{cal}\Delta T$

C_{cal} = heat capacity of calorimeter

Example 2: Constant-Pressure Calorimetry

When a 6.50 g sample of NaOH dissolves in 100.0 g of water in a coffee-cup calorimeter, the temperature increases from 21.6°C to 37.8°C. Calculate the enthalpy change in kJ/mol NaOH assuming the specific heat is 4.18 J/g-K because it is an aqueous solution.

q_{rxn} = - (4.18 J/g-K)(106.5 g)(16.2 K) = -7212 J = -7.21 kJ

(6.50 g NaOH)(1 mol/40.0 g) = 0.1625 mol NaOH ΔH = q_{rxn} = (-7.21 kJ)/(0.1625 mol) = -44.4 kJ/mol NaOH

Example 3: Constant-Volume Calorimetry

A 2.200 g sample of quinine ($C_6H_4O_2$) is burned in a bomb calorimeter whose heat capacity is 7.854 kJ/°C. The temperature raises from 23.44°C to 30.57°C. What is the heat of combustion per gram and per mol $C_4H_6O_2$?

q_{rxn} = - (7.854 kJ/°C)(7.13°C) = -56.0 kJ (-56.0 kJ/2.200 g) = -25.5 kJ/g (25.5 kJ/g)(108.1 g/1 mol) = -2756 kJ/mol

Section 5.6: Hess's Law

Hess's Law: if a reaction is carried out in a series of steps, ΔH for the reaction will be equal to the sum of the enthalpy changes for the individual steps

Example 4: Hess's Law

Using the equations: $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$ ΔH = -393.5 kJ $CO_{(g)} + \frac{1}{2}O_{2(g)} \rightarrow CO_{2(g)}$ ΔH = -283.0 kJ

Calculate the enthalpy change of: $C_{(s)} + \frac{1}{2}O_{2(g)} \rightarrow CO_{(g)}$

$C_{(s)} + \frac{1}{2}O_{2(g)} \rightarrow CO_{2(g)}$ ΔH = -393.5 kJ

$CO_{2(g)} \rightarrow CO_{(g)} + \frac{1}{2}O_{2(g)}$ -(ΔH = -283.0 kJ)

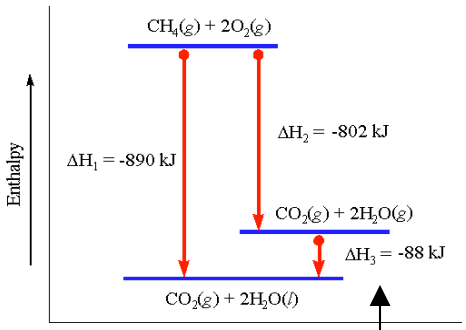
$C_{(s)} + \frac{1}{2}O_{2(g)} \rightarrow CO_{(g)}$ ΔH = -110.5 kJ

Section 5.7: Enthalpies of Formation

Enthalpy of formation (ΔH_f): enthalpy change associated with the formation of a compound from its constituent elements

Standard enthalpy (ΔH°): enthalpy change when all reactants and products are in their standard states

Standard enthalpy of formation ($\Delta H°_f$): change in enthalpy that forms 1 mol of the compound from its elements all in their standard states



Example of an enthalpy diagram

Using Enthalpies of Formation to Calculate Enthalpies of Reaction:

$\Delta H°_{rxn} = \Sigma n\Delta H°_f(\text{products}) - \Sigma m\Delta H°_f(\text{reactants})$

Example 5: Calculating Enthalpies of Reactions

Calculate the standard enthalpy change for the combustion of 1 mol of benzene, $C_6H_{6(l)}$, to $CO_{2(g)}$ and $H_2O_{(l)}$

$$C_6H_{6(l)} + (15/2)O_{2(g)} \rightarrow 6CO_{2(g)} + 3H_2O_{(l)}$$

$\Delta H°_{rxn} = [6\Delta H°_f(CO_2) + 3\Delta H°_f(H_2O)] - [\Delta H°_f(C_6H_6) + (15/2)\Delta H°_f(O_2)]$

= [6(-393.5 kJ) + 3(-285.8 kJ)] – [(49.0 kJ) + (15/2)(0 kJ)]

= (-2361 – 857.4 – 49.0) kJ

= -3267 kJ