

## Calorimetry Practice Worksheet

Key

- 1) Compound A is burned in a bomb calorimeter that contains 2.50 liters of water. If the combustion of 0.175 moles of this compound causes the temperature of the water to rise 45.0° C, what is the molar heat of combustion of compound A? The heat capacity of water is 4.184 J / g° C.

$$\Delta H = m c_p \Delta T$$

$$m = (2.50 \times 10^3 \text{ mL}) (1.00 \text{ g/mL}) = 2.50 \times 10^3 \text{ g}$$

$$c_p = 4.184 \text{ J/g}^\circ\text{C}$$

$$\Delta T = 45.0^\circ\text{C}$$

$$\Delta H = (2.50 \times 10^3 \text{ g}) (4.184 \text{ J/g}^\circ\text{C}) (45.0^\circ\text{C}) = 471,000 \text{ J} = 471 \text{ kJ}$$

$$\text{molar heat of combustion} = 471 \text{ kJ} / 0.175 \text{ mol} = \boxed{2.70 \times 10^3 \text{ kJ/mol}}$$

- 2) Compound B is burned in a bomb calorimeter that contains 1.50 liters of water. When 1 burned 50.0 grams of compound B in the calorimeter, the temperature rise of the water in the calorimeter was 35.0° C. If the heat of combustion of compound B is 2,150 kJ/mol, what is the molar mass of compound B?

$$\Delta H = m c_p \Delta T$$

$$m_{\text{water}} = 1.50 \times 10^3 \text{ g H}_2\text{O}$$

$$c_p = 4.184 \text{ J/g}^\circ\text{C}$$

$$\Delta T = 35.0^\circ\text{C}$$

$$\Delta H = (1.50 \times 10^3 \text{ g}) (4.184 \text{ J/g}^\circ\text{C}) (35.0^\circ\text{C}) = 2.20 \times 10^5 \text{ J}$$

$$\text{molar heat of combustion} = 2,150 \text{ kJ/mol}$$

$$n_B = \frac{2.20 \times 10^5 \text{ kJ}}{2,150 \text{ kJ}} = 0.102 \text{ mol B}$$

$$m_B = 50.0 \text{ g}$$

$$\text{Molar mass} = \frac{50.0 \text{ g B}}{0.102 \text{ mol}} = \boxed{490. \text{ g/mol}}$$

- 3) The molar heat of combustion of compound C is 1,250 kJ/mol. If I were to burn 0.115 moles of this compound in a bomb calorimeter with a reservoir that holds 2.50 L of water, what would the expected temperature increase be?

$$\Delta H = m c_p \Delta T$$

$$m = 2.50 \times 10^3 \text{ g H}_2\text{O}$$

$$c_p = 4.184 \text{ J/g}^\circ\text{C}$$

$$\Delta T = ?$$

$$\Delta H = (1,250 \text{ kJ/mol}) (0.115 \text{ mol}) = 144,000 \text{ J}$$

$$144,000 \text{ J} = (2.50 \times 10^3 \text{ g}) (4.184 \text{ J/g}^\circ\text{C}) (\Delta T)$$

$$\Delta T = \boxed{13.8^\circ\text{C}}$$

# Calorimetry Fundamentals And Practice

**Joacim Rocklöv**



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probabilistic solutions. 1.4. Road Map ...