

CALORIMETRY LAB PROCEDURES

- ① constant volume calorimetry (BOMB)
- ② constant pressure calorimetry (COFFEE CUP)
 - open to the air

$$\Delta H_{rxn} \quad \Delta H_{NEUT} \quad \Delta H_{soln} \quad \frac{kJ}{mol}$$

BOMB CALORIMETRY (constant V, $\Delta P \neq \Delta H$ but is very close)
 → isolated system

$$q_{rxn} = - \left[q_{H_2O} + q_{BOMB} \right]$$

1.320g SUCROSE $C_{12}H_{22}O_{11}$ burned in a bomb calorimeter
 $C_{BOMB} = 5.430 \text{ kJ/}^\circ\text{C}$; $957.0 \text{ g H}_2\text{O}$
 $T_i = 25.00^\circ\text{C}$, $T_f = 27.31^\circ\text{C}$

$$\Delta H_{rxn} = ? \text{ kJ/mol}$$

$$\rightarrow 3.856 \times 10^{-3} \text{ mol}$$

$$q_{rxn} = - \left[(957.0 \text{ g}) \left(4.184 \frac{\text{J}}{\text{g}^\circ\text{C}} \right) (2.31^\circ\text{C}) + 5.430 \frac{\text{kJ}}{^\circ\text{C}} (2.31^\circ\text{C}) \right]$$

$$= - \left[2.179 \times 10^4 \text{ J} \right]$$

$$\Delta H_{rxn} = \frac{-21.79 \text{ kJ}}{3.856 \times 10^{-3} \text{ mol}}$$

$$q_{rxn} = -21.79 \text{ kJ}$$

$$\Delta H_{rxn} = \ominus 5651 \text{ kJ/mol}$$



$$\Delta H_{comb} = -5651 \text{ kJ/mol}_{rxn}$$

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$$\Delta H_{rxn} \quad \Delta H_{NEUT} \quad \Delta H_{SOLN} \quad \text{kJ/mol}$$

BOMB CALORIMETRY (constant V, ΔP $q \neq \Delta H$ but is very close)
→ isolated system

$$q_{rxn} = - \left[q_{H_2O} + q_{BOMB} \right]$$

msΔT CΔT

1.320g sucrose $C_{12}H_{22}O_{11}$ burned in a bomb calorimeter
 $C_{BOMB} = 5.430 \text{ kJ/}^\circ\text{C}$; 957.0g H_2O
 $T_i = 25.00^\circ\text{C}$, $T_f = 27.31^\circ\text{C}$

$$\Delta H_{rxn} = ? \text{ kJ/mol}$$

$$\rightarrow 3.856 \times 10^{-3} \text{ mol}$$

$$q_{rxn} = - \left[(957.0g)(4.184 \text{ J/g}^\circ\text{C})(2.31^\circ\text{C}) + 5430 \text{ J/}^\circ\text{C}(2.31^\circ\text{C}) \right]$$
$$= - \left[2.179 \times 10^4 \text{ J} \right]$$

$$\Delta H_{rxn} = \frac{-21.79 \text{ kJ}}{3.856 \times 10^{-3} \text{ mol}}$$

$$q_{rxn} = -21.79 \text{ kJ}$$

$$\Delta H_{rxn} = \ominus 5651 \text{ kJ/mol}$$



$$\Delta H_{comb} = -5651 \text{ kJ/mol}_{rxn}$$

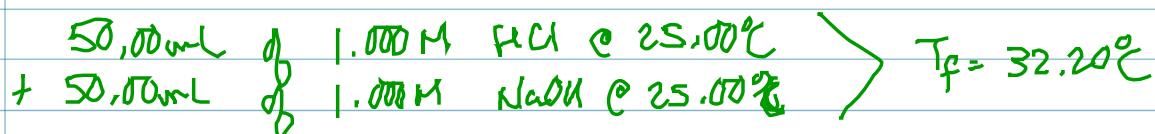
The enthalpy of combustion of benzoic acid C_6H_5COOH is commonly used as a standard for calibrating constant-volume calorimeters; its value has been accurately determined to be -3226.7 kJ/mol . When 1.9862 g of benzoic acid is burned, the temperature of the calorimeter rises from 21.84°C to 25.67°C . What is the heat capacity of the calorimeter? The calorimeter contains 2.000 kg of H_2O .

$$1.9862 \text{ g} \times \frac{1 \text{ mol}}{122.13 \text{ g}} \times \frac{-3226.7 \text{ kJ}}{\text{mol}} = -52.476 \text{ kJ} \quad \text{"q}_{\text{rxn}}"$$

$$-52.476 \text{ J} = -[(2000 \text{ g})(4.184 \text{ J/g}^\circ\text{C})(3.83^\circ\text{C}) + C_{\text{BOMB}}(3.83^\circ\text{C})]$$

$$C_{\text{BOMB}} = 5333 \text{ J/}^\circ\text{C} = 5.333 \text{ kJ/}^\circ\text{C}$$

"COFFEE CUP" CALORIMETRY $\text{constant } P \Rightarrow \Delta H = \frac{q}{n}$



$$q = ? \quad \Delta H = \frac{q}{n} \quad * \text{negligible heat capacity for calorimeter}$$

$$q_{\text{rxn}} = -[q_{\text{H}_2\text{O}}]$$

$$= -[(100 \text{ g})(4.184 \text{ J/g}^\circ\text{C})(7.20^\circ\text{C})] = 3012 \text{ J} = -3.012 \text{ kJ}$$

$$\text{mol}_{\text{HCl}} = M \times L = (1.000 \text{ M})(0.05000 \text{ L}) = 0.05000 \text{ mol}$$

$$\Delta H = \frac{-3.012 \text{ kJ}}{0.05000 \text{ mol}} = -60.24 \text{ kJ/mol}$$