

## SOLVING BOMB CALORIMETER PROBLEMS

A 1.000 g sample of octane ( $C_8H_{18}$ ) is burned in a bomb calorimeter containing 1200 grams of water at an initial temperature of  $25.00^\circ C$ . After the reaction, the final temperature of the water is  $33.20^\circ C$ . The heat capacity of the calorimeter (also known as the "calorimeter constant") is  $837 J/^\circ C$ . The specific heat of water is  $4.184 J/g^\circ C$ . **Calculate the heat of combustion of octane in kJ/mol.**

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Since this is a combustion reaction, heat flows from the system to the surroundings... thus, it is exothermic. The heat released by the reaction will be absorbed by two things: (a) the water in the calorimeter and (b) the calorimeter itself.

- a. Calculate the heat absorbed by the water ( $q_{water}$ )

$$m = 1200 \text{ grams}$$

$$c_{water} = 4.184 J/g^\circ C$$

$$\Delta T = 33.20 - 25.00 = 8.20^\circ C$$

$$q_{water} = (m)(c)(\Delta T), \text{ so}$$

$$q_{water} = (1200 \text{ g})(4.184 J/g^\circ C)(8.20^\circ C) = 41170.56 \text{ J} = \mathbf{41.2 \text{ kJ}}$$

- b. Calculate the heat absorbed by the calorimeter ( $q_{cal}$ )

The temperature change of the calorimeter is the same as the temperature change for water. In this step, however, we must use the **heat capacity** of the calorimeter, which is already known. When using heat capacity, the mass of the calorimeter is not required for the calculation. (It's already incorporated into the heat capacity).

$$C_{cal} = 837 J/^\circ C$$

$$\Delta T = 33.20 - 25.00 = 8.20^\circ C$$

$$q_{cal} = (C_{cal})(\Delta T), \text{ so}$$

$$\text{so, } q_{cal} = (837 J/^\circ C)(8.20^\circ C) = 6863.4 \text{ J} = \mathbf{6.86 \text{ kJ}}$$

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The TOTAL heat absorbed by the water and the calorimeter is the sum of (a) and (b):  $41.2 + 6.86 = \mathbf{+ 48.1 \text{ kJ}}$ . (Remember,  $q$  is positive because the heat is being absorbed).

The amount of heat released by the reaction is equal to the amount of heat absorbed by the water and the calorimeter. We just need to change the sign. So,

$$\mathbf{q_{reaction} = - 48.1 \text{ kJ}}$$

Since 1.000 gram of octane was burned, the **heat of combustion** for octane is equal to  $- 48.1 \text{ kJ/gram}$ . In other words, when one mole of octane is burned,  $48.1 \text{ kJ}$  of heat is released from the reaction. *What is the heat of combustion in kJ/mol?*

$$1 \text{ mol of octane weighs } 114 \text{ grams, so } (-48.1 \text{ kJ/g})(114 \text{ g/mol}) = \mathbf{- 5483 \text{ kJ/mol.}}$$

## BOMB CALORIMETRY PRACTICE PROBLEMS

Note: the specific heat of water is 4.184 J/g°C

1. A 0.500 g sample of naphthalene ( $C_{10}H_8$ ) is burned in a bomb calorimeter containing 650 grams of water at an initial temperature of 20.00°C. After the reaction, the final temperature of the water is 26.4°C. The heat capacity of the calorimeter is 420 J/°C. Using these data, calculate the heat of combustion of naphthalene in kJ/mol.
2. A bomb calorimeter containing 900 grams of water was calibrated by burning a sample of benzoic acid ( $C_6H_5COOH$ ), whose heat of combustion is  $-3227$  kJ/mol. When 1.890 g of benzoic acid is burned in the calorimeter, the temperature of the water and the calorimeter increases by 8.32°C. Using these data, calculate the calorimeter constant for this calorimeter in J/°C.
3. A 1.000 gram sample of the rocket fuel hydrazine ( $N_2H_4$ ) is burned in a bomb calorimeter. The temperature rises from 24.62°C to 28.16°C. The heat capacity of the calorimeter (including the water) is 5860 J/°C. Calculate the molar heat of combustion of hydrazine, in kJ/mole.
4. A 0.500 g sample of TNT ( $C_7H_5N_2O_6$ ) is burned in a bomb calorimeter containing 610 grams of water at an initial temperature of 20.00°C. The heat capacity of the calorimeter is 420 J/°C and the heat of combustion of TNT is 3374 kJ/mol. Using these data, calculate the final temperature of the water and calorimeter once the reaction is complete.

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### ANSWERS

1.  $\Delta H_{\text{combustion}} = -5140$  kJ/mol
2.  $C_{\text{calorimeter}} = 2243$  J/°C
3.  $\Delta H_{\text{combustion}} = -664$  kJ/mol
4.  $\Delta T = 2.50^\circ\text{C}$ , so the final temperature = 22.5°C