Thermochemistry

Example 1

How much heat is required to raise the temperature of 7.35 g of water from 21.0°C to 98.0°C? (Assume that the specific heat of water 4.18 J/g.°C throughout this temperature range).

T_f T_i q

$q = ?$

$m = 7.35\,g$

$T_i = 21.0°c$

$T_f = 98.0°c$

$C_p = 4.18\ j/g.°c$

$\Delta T = T_f - T_i$

$= 98.0 - 21.0 = 77.0°c$

$$\boxed{q = m \cdot c \cdot \Delta T}$$

$= 7.35 \times 4.18 \times 77$

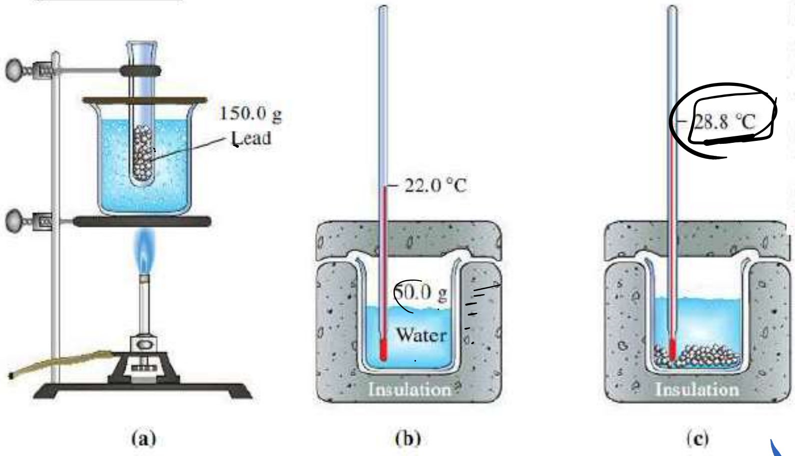
$= 2365.67\ j$

$= 2.36\ kj$

$$C = \frac{q}{m\,\Delta T}$$

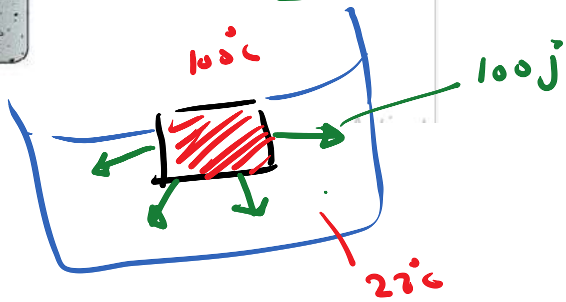
Example 2

- *Calculate the specific heat of lead according to the given data.*



(a) A 150.0 g sample of lead is heated to the temperature of boiling water (100.0 °C). (b) A 50.0 g sample of water is added to a thermally insulated beaker, and its temperature is found to be 22.0 °C. (c) The hot lead is dumped into the cold water, and the temperature of the final lead–water mixture is 28.8 °C.

released
exothermic
-ve



gained
endoth the

Pb

$m = 150\ g$

$T_i = 100°c$

$T_f = 28.8°c$

$C_p = ?$

H_2O

$m = 50\ g$

$T_i = 22.0°c$

$T_f = 28.8°c$

$C_p = 4.18\ j/g°c$

$$q_{lead} = -(+q_{H_2O})$$

$$q_{pb} = -q_{H_2O}$$

$$m_{pb} \cdot C_{pb} \cdot \Delta T = -(m_{H_2O} \cdot C_{H_2O} \cdot \Delta T)$$

$$150 \cdot C_{pb} \cdot (28.8 - 100) = -(50 \cdot 4.18) \times (28.8 - 22.0)$$

$$C_{pb} = \frac{-(50 \cdot 4.18 \times 6.8)}{(28.8 - 100) \times 150} = 0.13\ j/g°c$$

all combustion reaction are exothermic (−ve)

Example 3

$12C + 22H + 11O$
$= 12\times 12 + 22\times 1 + 11\times 16 = 342.3$ g/mol

- The combustion of 1.010 g sucrose, $C_{12}H_{22}O_{11}$, in a bomb calorimeter causes the temperature to rise from 24.92 °C to 28.33 °C. The heat capacity of the calorimeter assembly is 4.90 kJ/°C.

(a) What is the heat of combustion (ΔH) of sucrose expressed in kilojoules per mole of sucrose (kJ/mol)

(b) Verify the claim of sugar producers that one teaspoon of sugar (about 4.8 g) contains only 19 Calories.

Calories

calories

$q \rightarrow$ amount of heat $\rightarrow$ heat transfer between two objects.

$\Delta H \rightarrow$ amount of heat gained or released as a result from a chemical Rx.

$q = m \cdot c \cdot \Delta T$

$\Delta H = \dfrac{q}{n}$ → no. of moles.

$n = \dfrac{m}{Mr}$ → molar mass

(a) $q = m \cdot c \cdot \Delta T$

$= 1.010 \times 4.90 \times (28.33 - 24.92)$

$= 16.87$ kJ

$n = \dfrac{m}{Mr} = \dfrac{1.010}{342.3} = 0.00295$ mol

$\Delta H = \dfrac{-16.87}{0.00295} = -5718.6$ kJ/mol

(b) $q = -\Delta H \cdot n = \Delta H \cdot \dfrac{m}{Mr} = 5718.6 \times \dfrac{4.8}{342.3}$

$= 71.8$ kJ

$= \dfrac{71.8}{4.2} = 17.1$ Cal

$1 \text{ kJ} \rightarrow \dfrac{1}{4.2}$ Cal