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).

$q = ?$

$m = 7.35\ g$

$T_i = 21.0\ °C$

$T_f = 98.0\ °C$

$C = 4.18\ J/g.°C$

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

$= 7.35 \times 4.18 \times (98.0 - 21.0)$

$= 2365.67\ J = 2.365\ kJ$

$q = m \cdot c \cdot \Delta T$ — m: mass (g); q: amount of heat (J); c → specific heat

$\Delta T = T_f - T_i$

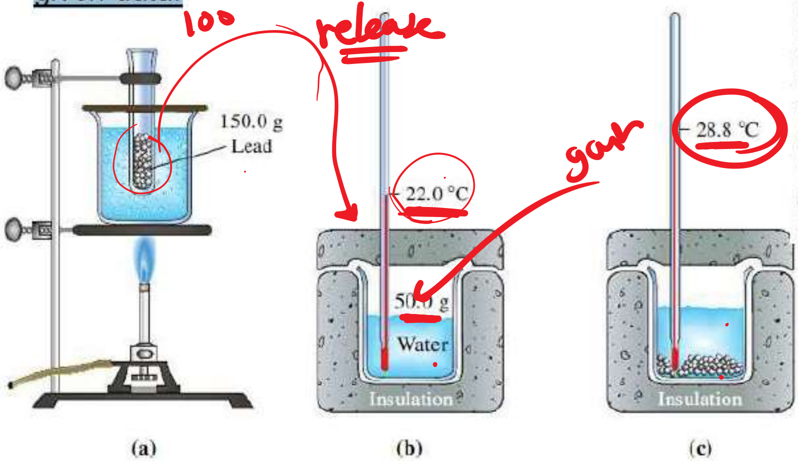
$\Delta H = \frac{q}{n}$

$q_{released} = -q_{absorb}$

$q = C\,\Delta T$ — C: heat capacity.

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.

Pb

$m = 150\ g$

$T_i = 100^\circ c$

$T_f = 28.8^\circ c$

$c = ?$

H_2O

$m = 50.0\ g$

$T_i = 22.0^\circ c$

$T_f = 28.8^\circ c$

$c = 4.18$

$$q_{lead} = -q_{H_2O}$$

$$q_{H_2O} = m \cdot c \cdot \Delta T$$

$$= 50 \times 4.18 \times (28.8 - 22.0) = 1.14 \times 10^{-3}\ j$$

$$q_{lead} = -q_{H_2O} = -1.14 \times 10^{-3}$$

$$m \cdot c \cdot \Delta T = -1.14 \times 10^{-3}$$

$$c = \frac{-1.14 \times 10^{-3}}{m\,\Delta T}$$

$$= \frac{-1.14 \times 10^{-3}}{150 \times (28.8 - 100)}$$

$$= 0.13\ j/g\,^\circ C$$

Example 3

- 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.

t_i t_f

(a) What is the heat of combustion of sucrose expressed in kilojoules per mole of sucrose

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

$m_{sucrose} = 1.010 g$

$T_i = 24.92 °c$

$T_f = 28.33 °c$

$C = 4.90 kj/°c$

$q = ?$ kj/mol

$q = m.c.\Delta T$

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

$= 16.87$ kj.

$n_{sucrose} = \frac{m}{Mr} = \frac{0.1010g}{342.3} = 0.00295$ mol.

$\Delta H = -\frac{Q}{n} = \frac{-16.87}{0.00295} = -5718.6$ kj/mol.

(2) $Q = -\Delta H_{comb} \times \frac{m}{Mr} = 5718.6 \times \frac{4.8}{342.3}$

$= 71.8$ kj → Cal

$$1\ kj \rightarrow \frac{1}{4.2}\ Cal$$

$$Q = \frac{71.8}{4.2} = 17.1\ Cal$$