If you combine 410.0 mL of water at 25.00 °C and 130.0 mL of water at 95.00 °C, what is the final temperature of the mixture? Use 1.00 g/mL as the density of water.

Solution.

The heat lost by the hot water sample will be equal to the heat absorbed by the room-temperature water sample.

$$-\mathrm{q}_{\mathrm{lost}} = \mathrm{q}_{\mathrm{absorbed}}, \text{‘(1)}$$

q_{lost} - heat lost by the hot water;

$\mathrm{q}_{\mathrm{absorbed}}$- heat absorbed by the room-temperature water sample.

$$\mathrm{q} = \mathrm{m} \cdot \mathrm{c}_{water} \cdot \Delta\mathrm{T}, \text{(2)}$$

q – the amount of heat;

m – the mass of the sample;

c_{sp} – the specific heat of the water;

$\Delta\mathrm{T}$ – the change in temperature.

$$\Delta\mathrm{T}_{\mathrm{sample1}} = \mathrm{T}_{\mathrm{final}} - 25.00\ °\mathrm{C}$$

$$\Delta\mathrm{T}_{\mathrm{sample2}} = \mathrm{T}_{\mathrm{final}} - 95.00\ °\mathrm{C}$$

Use equations (1) and (2) top write:

$$-\mathrm{m}_{\mathrm{sample1}} \cdot \mathrm{c}_{sp} \cdot \Delta\mathrm{T}_{\mathrm{sample1}} = \mathrm{m}_{\mathrm{sample2}} \cdot \mathrm{c}_{sp} \cdot \Delta\mathrm{T}_{\mathrm{sample2}}$$

$$-\mathrm{m}_{\mathrm{sample1}} \cdot \Delta\mathrm{T}_{\mathrm{sample1}} = \mathrm{m}_{\mathrm{sample2}} \cdot \Delta\mathrm{T}_{\mathrm{sample2}}$$

$$-410.00\ \mathrm{g} \cdot (\mathrm{T}_{\mathrm{final}} - 25.00\ °\mathrm{C}) = 130.00\ \mathrm{g} \cdot (\mathrm{T}_{\mathrm{final}} - 95.00\ °\mathrm{C})$$

$$-410.00 \cdot \mathrm{T}_{\mathrm{final}} + 10250 = 130.00 \cdot \mathrm{T}_{\mathrm{final}} - 12350$$

$$540.00 \cdot \mathrm{T}_{\mathrm{final}} = 22600$$

$$\mathrm{T}_{\mathrm{final}} = 41{,}85\ °\mathrm{C}$$

Answer.

The final temperature of the mixture

$$T_{final} = 41,85\ °C$$