Calculating Energy for Changes of Phase

To Calculate Heat:
Heat to Change the Temperature: \[ Q = m \cdot \Delta T \cdot s \]
Heat to Melt: \[ Q = m \cdot \Delta H_{\text{fus}} \quad \text{Heat to Boil:} \quad Q = m \cdot \Delta H_{\text{vap}} \]

For Water:
\[ \Delta H_{\text{fus}} = 334 \text{ J/g} \quad \Delta H_{\text{vap}} = 2260 \text{ J/g} \]
Specific Heat: ice = 2.06 J/g °C \quad \text{water} = 4.184 \text{ J/g °C} \quad \text{steam} = 2.03 \text{ J/g °C}

Directions: Calculate how much energy it takes to convert 25.0 g of ice at –15 °C to steam at 150 °C.
Give the values for each step. Show your work. Write the formula. Substitute the values.
Use dimensional analysis. Write units. Use sig. figs.

1. Heat the ice from –15 °C to 0 °C

2. Melt the ice at 0 °C

3. Heat the water from 0 °C to 100 °C

4. Evaporate the water at 100 °C

5. Heat the steam from 100 °C to 150 °C

6. Total heat required?

7. Which of the steps required the most heat energy?

Why?
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To Calculate Heat:

Heat to Change the Temperature: \[ Q = m \cdot \Delta T \cdot s \]
Heat to Melt: \[ Q = m \cdot \Delta H_{\text{fus}} \]
Heat to Boil: \[ Q = m \cdot \Delta H_{\text{vap}} \]

For Water:

\[ \Delta H_{\text{fus}} = 334 \text{ J/g} \quad \Delta H_{\text{vap}} = 2260 \text{ J/g} \]

Specific Heat: ice = 2.06 J/g °C \quad water = 4.184 J/g °C \quad steam = 2.03 J/g °C

Directions: Calculate how much energy it takes to convert 25.0 g of ice at −15.0°C to steam at 150.0°C. Give the values for each step. Show your work. Write the formula. Substitute the values. Use dimensional analysis. Write units. Use sig. figs.

1. Heat the ice from −15 °C to 0 °C

\[ Q = m \cdot \Delta T \cdot s \quad Q \boxed{25.0 \text{ g} \quad 15.0^\circ \text{C} \quad 2.06 \text{ J/g °C}} \quad = 773 \text{ J} \]

2. Melt the ice ice at 0 °C

\[ Q = m \cdot \Delta H_{\text{fus}} \quad Q \boxed{25.0 \text{ g} \quad 334 \text{ J}} \quad = 8,350 \text{ J} \]

3. Heat the water from 0 °C to 100 °C

\[ Q = m \cdot \Delta T \cdot s \quad Q \boxed{25.0 \text{ g} \quad 100^\circ \text{C} \quad 4.184 \text{ J/g °C}} \quad = 10,500 \text{ J} \]

4. Evaporate the water at 100 °C

\[ Q = m \cdot \Delta H_{\text{vap}} \quad Q \boxed{25.0 \text{ g} \quad 2260 \text{ J}} \quad = 56,500 \text{ J} \]

5. Heat the steam from 100 °C to 150 °C

\[ Q = m \cdot \Delta T \cdot s \quad Q \boxed{25.0 \text{ g} \quad 50.0^\circ \text{C} \quad 2.03 \text{ J/g °C}} \quad = 2,540 \text{ J} \]

6. Total heat required?

\[ Q = m \cdot \Delta T \cdot s \quad Q \boxed{25.0 \text{ g} \quad 50.0^\circ \text{C} \quad 2.03 \text{ J/g °C}} \quad = 78,700 \text{ J} \]

7. Which of the steps required the most heat energy? Step 4 Evaporating the Water Why? To evaporate water we must add enough heat energy to overcome the strong hydrogen bonds holding the H₂O molecules together.