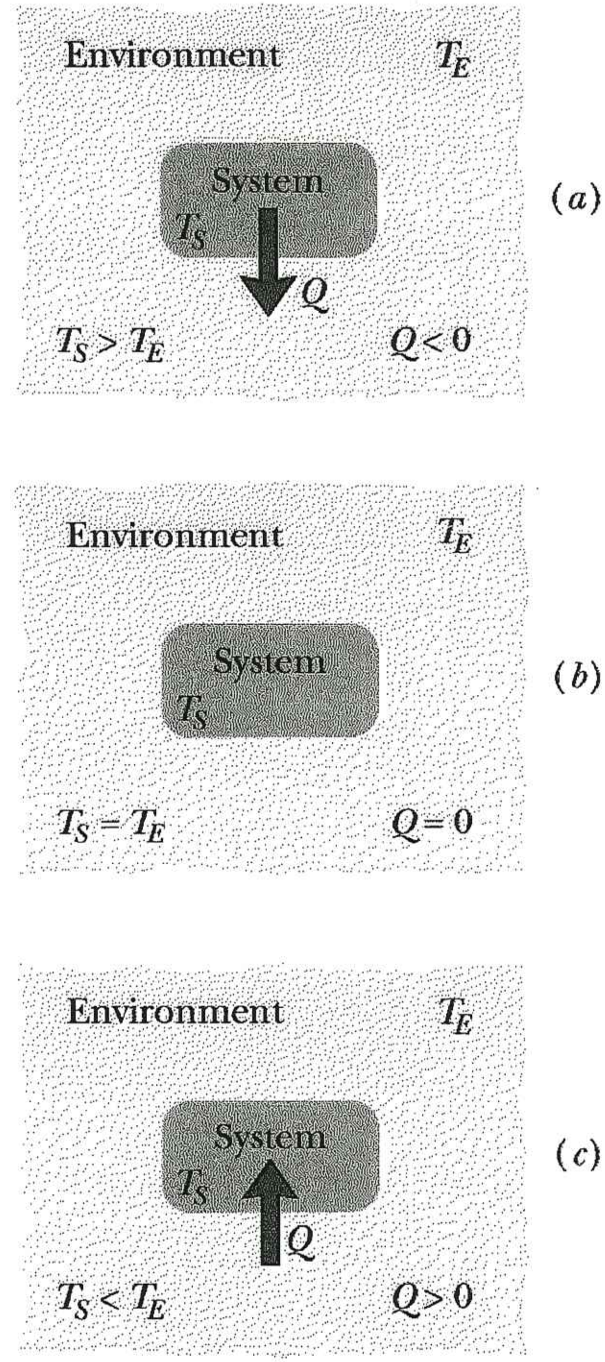
Lecture 36

(Ch. 11: 2)

Temperature and Heat (Q)

- System vs Environment
- Change in temperature: transfer of internal energy, called ***heat***
- (see Fig 18-12 =>)
- ***Heat*** is the energy that is transferred between a system and its environment because of the temperature difference that exists between them



Absorption of Heat

- **Heat Capacity**
  - capacity of a body to absorb heat
  - specific to one body

$$Q = C(T_f - T_i)$$

$$C = \frac{Q}{T_f - T_i}$$

  - Units: cal/K, Btu/K, J/K

- **Specific Heat**
  - specific to units of mass

$$Q = c\,m(T_f - T_i)$$

$$c = \frac{Q}{m\Delta T} = \frac{Q}{m(T_f - T_i)}$$

  - specific heat of water (for other Table 18-3)

$$c_w = 1 cal / g \cdot K = 1 Btu / lb \cdot F = 4190 J / Kg \cdot K$$

Checkpoint 18-3

A certain amount of heat Q will warm 1 g of material A by 3° C and 1 g of material B by 4° C. Which material has the greatest specific heat?

Table 13-1

TABLE 13.1 Specific Heat of Selected Materials (at $T = 20°C$ unless indicated)

Material	Specific heat c, $J/(kg \cdot °C)$	Specific heat c, $cal/(g \cdot °C)$
Aluminum	900	0.215
Beryllium	1970	0.471
Copper	385	0.092
Ethanol	2430	0.581
Human body (average, $T = 37°C$)	3500	0.840
Ice (0°C)	2090	0.499
Iron	449	0.107
Lead	128	0.031
Mercury	140	0.033
Silver	235	0.056
Water	4186	1.000
Wood (typical)	1400	0.33
Steel (typical)	500	0.12

An iron rod of mass 0.5 kg is at temperature of 20 °C. How much heat, Q, in Joules must it absorb so that its temperature raises to 80 °C ?

By definition $Q = m.c.\Delta T$ where specific heat of iron $c_{Fe} = 449$ J/(kg. °C)

Therefore, $Q = (0.5\ \text{kg}).\ (449).(80 - 20) = 13470$ J

9. Lake Erie contains roughly $4.00 \times 10^{11}\ \text{m}^3$ of water.

   a. How much energy is required to raise the temperature of that volume of water from 11.0°C to 12.0°C?

   Answer ↓

   b. How many years would it take to supply this amount of energy by using the 1.00×10^4-MW exhaust energy of an electric power plant?

11.9 The mass of water involved is

$$m = \rho V = \left(10^3 \frac{\text{kg}}{\text{m}^3}\right)\left(4.00\times 10^{11}\ \text{m}^3\right) = 4.00\times 10^{14}\ \text{kg}$$

(a) $Q = mc(\Delta T) = (4.00 \times 10^{14}\ \text{kg})(4\,186\ \text{J/kg}\cdot{}^\circ\text{C})(1.00^\circ\text{C}) = \boxed{1.67 \times 10^{15}\ \text{J}}$

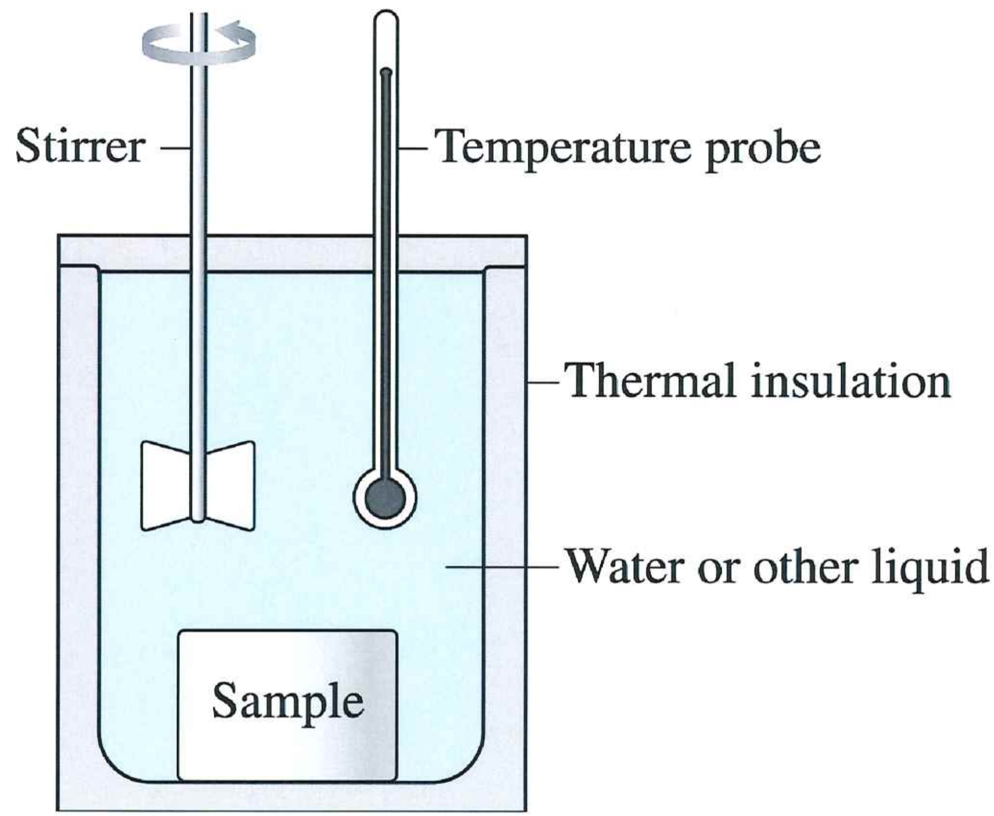
(b) The power input is $P = 1\,000\ \text{MW} = 1.00 \times 10^9\ \text{J/s}$,

so $$t = \frac{Q}{P} = \frac{1.67\times 10^{15}\ \text{J}}{1.00\times 10^9\ \text{J/s}}\left(\frac{1\ \text{yr}}{3.156\times 10^7\ \text{s}}\right) = \boxed{52.9\ \text{yr}}$$

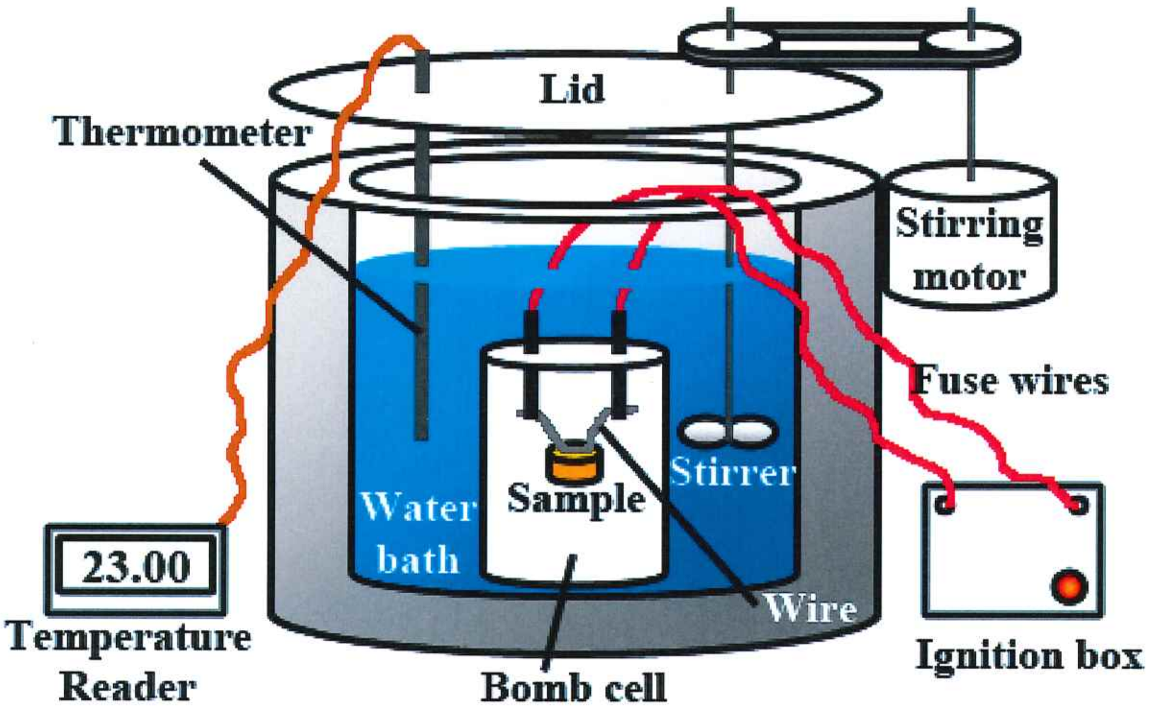
Chapter 11: Heat - Calorimetry

Calorimetry is a technique for types of measurements that involve transferring thermal energy from one object with a measured initial temperature to a known quantity of water at a known initial temperature.

The change in temperature of the water will indicate the heat capacity of the unknown object that is cooled. Knowing the mass of the object, the specific heat is obtained.



Calorimetry



Example:

$$T_{\text{beaker}} = 25°$$

$$T_{\text{water}} = 40°$$

$$T_{\text{Al}} = 37°$$

$$Q = mc\left(T_f - T_i\right)$$

$$\Sigma Q_k = 0$$

$$Q_{\text{cold}} = -Q_{\text{hot}}$$

$$Q_w + Q_{\text{al}} + Q_g = 0$$

,. You mix 18 kg of water at 25 °C with 6 kg of water at 2 °C, what is the final temperature ?

All we can say is that the hotter water changes temperature by: $\Delta T_{\text{hot}} = T_f - 25^\circ\text{C}$, while the colder water changes temperature by: $\Delta T_{\text{cold}} = T_f - 2.0^\circ\text{C}$. We will be able to solve for T_f using Equation 13.2,

$$\text{i.e. } Q = \text{c.m. } \Delta\text{T}$$

and the fact that the heat lost by the hot water is gained by the cold water: $Q_{\text{hot}} = -Q_{\text{cold}}$, assuming of course that no heat is lost to the surroundings.

Known: $m_{\text{hot}} = 18$ kg, $m_{\text{cold}} = 6$ kg. and c_{water}= 4186 J/(kg.°C) (Table 13.1)

SOLVE The equal but opposite heat exchange implies:

$$Q_{\text{hot}} = -Q_{\text{cold}} \Rightarrow m_{\text{hot}} c \Delta T_{\text{hot}} = -m_{\text{cold}} c \Delta T_{\text{cold}}$$

Solving for the final temperature:

$$T_f - 25^\circ\text{C} = -\frac{6 \text{ kg}}{18 \text{ kg}}(T_f - 2.0^\circ\text{C}) \Rightarrow T_f = 19^\circ\text{C}$$

20. A large room in a house holds 975 kg of dry air at 30.0° C. A woman opens a window briefly and a cool breeze brings in an additional 50.0 kg of dry air at 18.0° C. At what temperature will the two air masses come into thermal equilibrium, assuming they form a closed system? (The specific heat of dry air is 1 006 J/kg · °C, although that value will cancel out of the calorimetry equation.)

11.20 By conservation of energy, the two air masses will exchange energy such that $Q_{warm} + Q_{cool} = 0$. Substitute $Q = mc\Delta T$ to find the final temperature:

$$Q_{warm} = -Q_{cool}$$

$$m_{warm}c_{air}(T - T_{warm}) = -m_{cool}c_{air}(T - T_{cool})$$

$$T(m_{warm} + m_{cool}) = m_{warm}T_{warm} + m_{cool}T_{cool}$$

$$T = \frac{m_{warm}T_{warm} + m_{cool}T_{cool}}{m_{warm} + m_{cool}} = \frac{(975\text{ kg})(30.0°\text{C}) + (50.0\text{ kg})(18.0°\text{C})}{975\text{ kg} + 50.0\text{ kg}}$$

$$= \boxed{29.4°\text{C}}$$

17. What mass of water at 25.0°C must be allowed to come to thermal equilibrium with a 1.85-kg cube of aluminum initially at 1.50×10^{2}°C to lower the temperature of the aluminum to 65.0°C? Assume any water turned to steam subsequently recondenses.

11.17 When thermal equilibrium is reached, the water and aluminum will have a common temperature of $T_f = 65.0°C$. Assuming that the water-aluminum system is thermally isolated from the environment, $Q_{cold} = -Q_{hot}$, so $m_w c_w (T_f - T_{i,w}) = -m_{Al} c_{Al} (T_f - T_{i,Al})$, or

$$m_w = \frac{-m_{Al} c_{Al} (T_f - T_{i,Al})}{c_w (T_f - T_{i,w})} = \frac{-(1.85 \text{ kg})(900 \text{ J/kg}\cdot°\text{C})(65.0°\text{C} - 150°\text{C})}{(4\,186 \text{ J/kg}\cdot°\text{C})(65.0°\text{C} - 25.0°\text{C})}$$

$$= \boxed{0.845 \text{ kg}}$$

Sample Problem

A copper slug whose mass m_c is 75 g is heated in a laboratory oven to a temperature T of 312°C. The slug is then dropped in a glass beaker containing a mass $m_w = 220$ g of water. The heat capacity C_b of the beaker is 45 cal/K. The initial temperature T_i of the water and the beaker is 12°C. Assuming that the slug, beaker, and water are an isolated system and the water does not vaporize, find the temperature T_f of the system at thermal equilibrium.

$$\left.\begin{aligned} Q_w &= c_w m_w (T_f - T_i) \\ Q_b &= C_b (T_f - T_i) \\ Q_c &= c_c m_c (T_f - T_i) \\ Q_w + Q_b + Q_c &= 0 \end{aligned}\right\} c_w m_w (T_f - T_i) + C_b (T_f - T_i) + c_c m_c (T_f - T_i) = 0$$

$$c_w m_w T_f + C_b T_f + c_c m_c T_f = c_w m_w T_i + C_b T_i + c_c m_c T_i$$

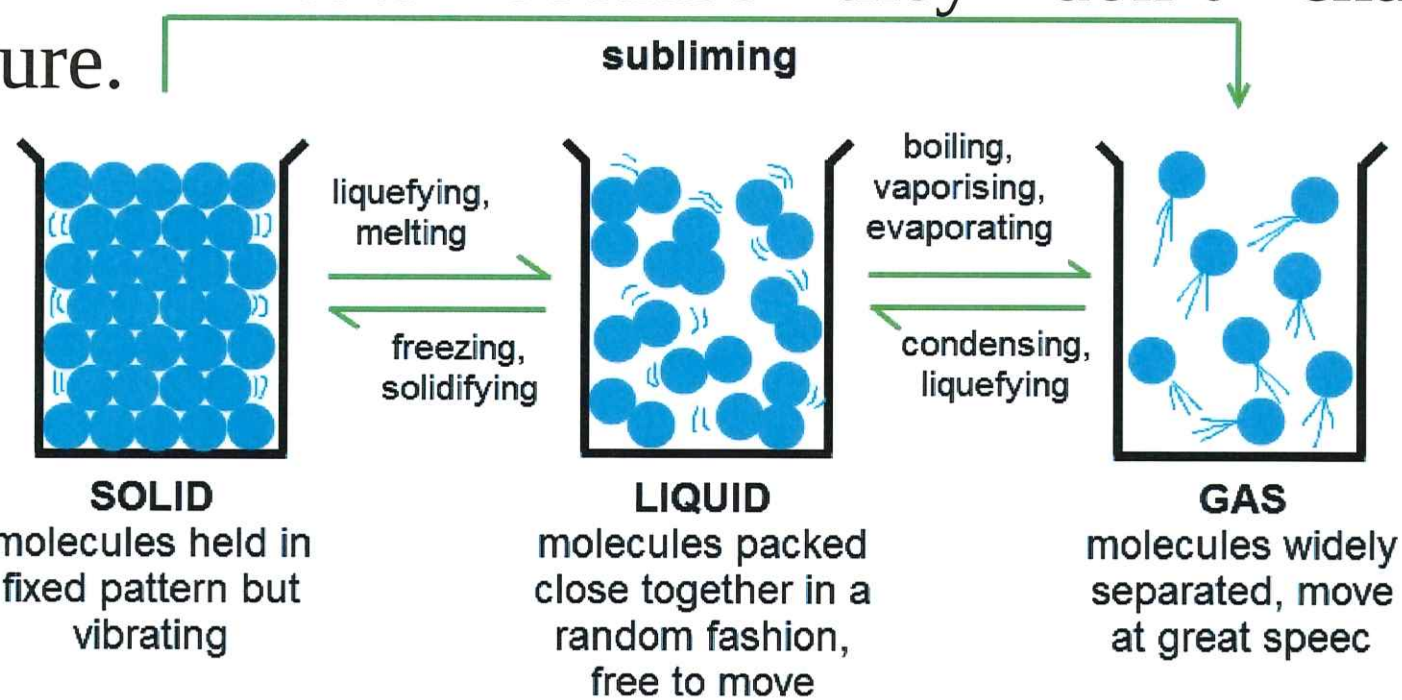
$$T_f = \frac{c_w m_w T_i + C_b T_i + c_c m_c T_i}{c_w m_w + C_b + c_c m_c} = \frac{(1.0)(220)(12^o) + (45)(12^o) + (0.0923)(75)(312^o)}{(1.0 cal/g\cdot K)(220g) + (45 cal/K) + (0.0923 cal/g\cdot K)(75g)}$$

$$T_f = 19.64^o C \qquad Q_w \approx 1680 cal \quad Q_b \approx 344 cal \quad Q_c \approx -2024 cal$$

Chapter 11: Heat – Phase Changes

Heats of transformation

For a **phase transition** to occur, some energy needs to be exchanged, and this is more energy than just changing the temperature by 1 °C. At a fundamental level, gaining or "activating" degrees of freedom requires an energy intake, while losing them will release energy. These energies are called ***latent heats*** because they don't change the temperature.



Chapter 11: Heat – Phase Changes

Heat of fusion

Melting a solid into a liquid requires energy that can separate the "fused" molecules.

$Q_f = m \cdot L_f$, where L_f is the heat of fusion, in SI: J/kg

Heat of vaporization

Turning a liquid into gas requires further energy that can separate the closely packed molecules.

$Q_v = m \cdot L_v$, where L_v is the heat of vaporization, in SI: J/kg

Reversing the process, the equations are valid, but the heat will have a negative sign, as energy is released in condensing (liquefying) or freezing (solidifying).

Latent Heat and Phase Change

$$Q = \pm mL$$

for water:

$$L_f = 3.33\times10^5 \text{ J/kg}$$

$$L_v = 2.26\times10^6 \text{ J/kg}$$

Table 11.2 Latent Heats of Fusion and Vaporization

		Latent Heat of Fusion			Latent Heat of Vaporization	
Substance	**Melting Point (°C)**	**(J/kg)**	**cal/g**	**Boiling Point (°C)**	**(J/kg)**	**cal/g**
Helium	−269.65	5.23×10^3	1.25	−268.93	2.09×10^4	4.99
Nitrogen	−209.97	2.55×10^4	6.09	−195.81	2.01×10^5	48.0
Oxygen	−218.79	1.38×10^4	3.30	−182.97	2.13×10^5	50.9
Ethyl alcohol	−114	1.04×10^5	24.9	78	8.54×10^5	204
Water	0.00	3.33×10^5	79.7	100.00	2.26×10^6	540
Sulfur	119	3.81×10^4	9.10	444.60	3.26×10^5	77.9
Lead	327.3	2.45×10^4	5.85	1 750	8.70×10^5	208
Aluminum	660	3.97×10^5	94.8	2 450	1.14×10^7	2 720
Silver	960.80	8.82×10^4	21.1	2 193	2.33×10^6	558
Gold	1 063.00	6.44×10^4	15.4	2 660	1.58×10^6	377
Copper	1 083	1.34×10^5	32.0	1 187	5.06×10^6	1 210

How much heat, Q, is required to melt 500 g of ice at 0 °C ?

By definition $Q = mL_f$ where the heat of fusion of ice is $L_{ice} = 3.33 \times 10^5$ J/kg

Therefore, $Q = (0.5\ \text{kg}).\ (3.33 \times 10^5) = 166.5\ \text{kJ}$

Chapter 11: Heat – Phase Changes

Example

A 75 g cube of ice at 10.0 C is placed in 0.500 kg of water at 50.0 °C in an insulating container so that no heat is lost to the environment. Will the ice melt completely?

The heat required to completely melt the ice is:

$$Q_{ice} = m_{ice} c_{ice} \Delta T_{ice} + m_{ice} L_f$$

$$Q_{ice} = 0.075\,\text{kg} \cdot 2.1\,\text{kJ/kg °C} \cdot 10\,\text{°C} + 0.075\,kg \cdot 333.7\,\text{kJ/kg} = 27\,\text{kJ}$$

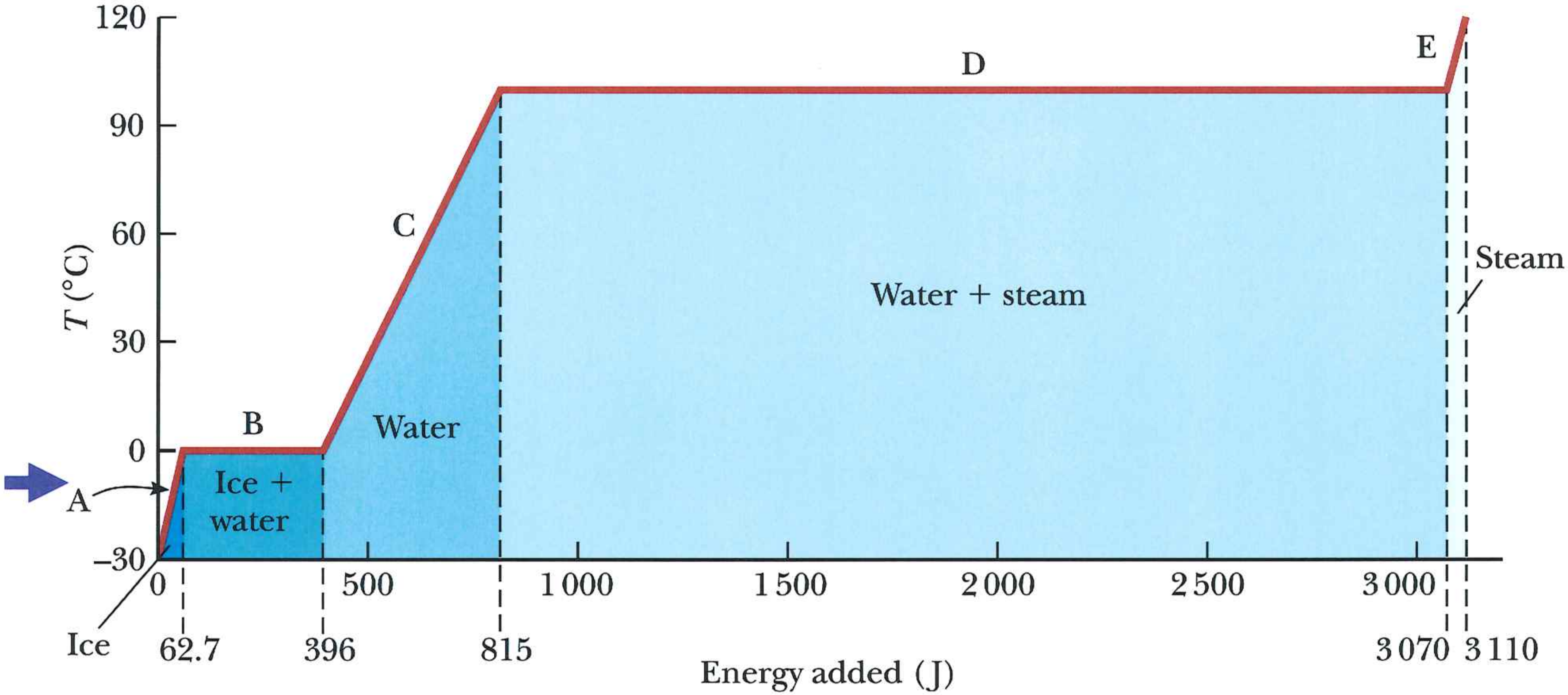
The heat required to cool the water to the freezing point is:

$$Q_{H_20} = m_{H_20} c_{H_20} \Delta T_{H_20}$$

$$Q_{H_20} = 0.5\,\text{kg} \cdot 4.186\,\text{kJ/kg°C} \cdot 50\,\text{°C} = 105\,\text{kJ}$$

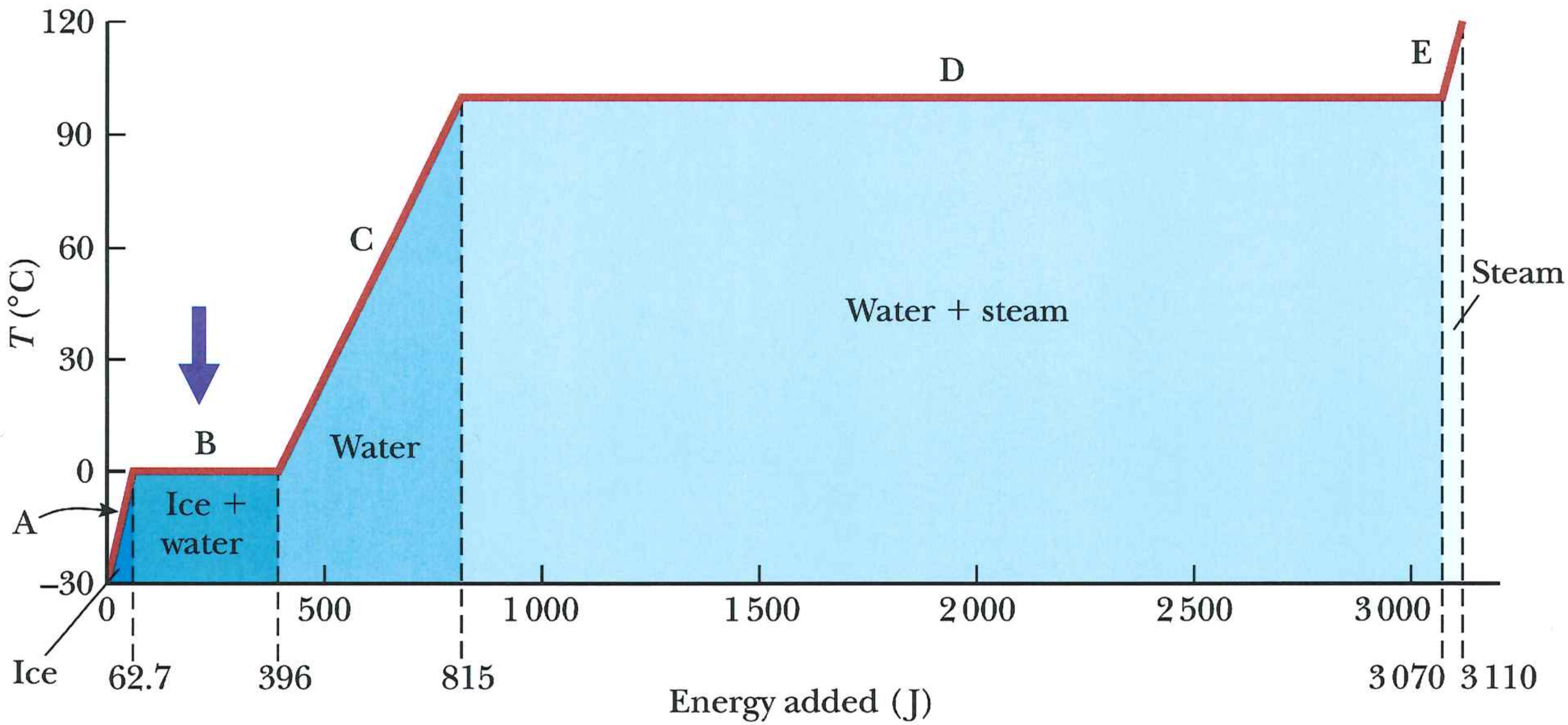
Since $Q_{ice} < Q_{water}$ the ice will completely melt.

Heat and Internal Energy



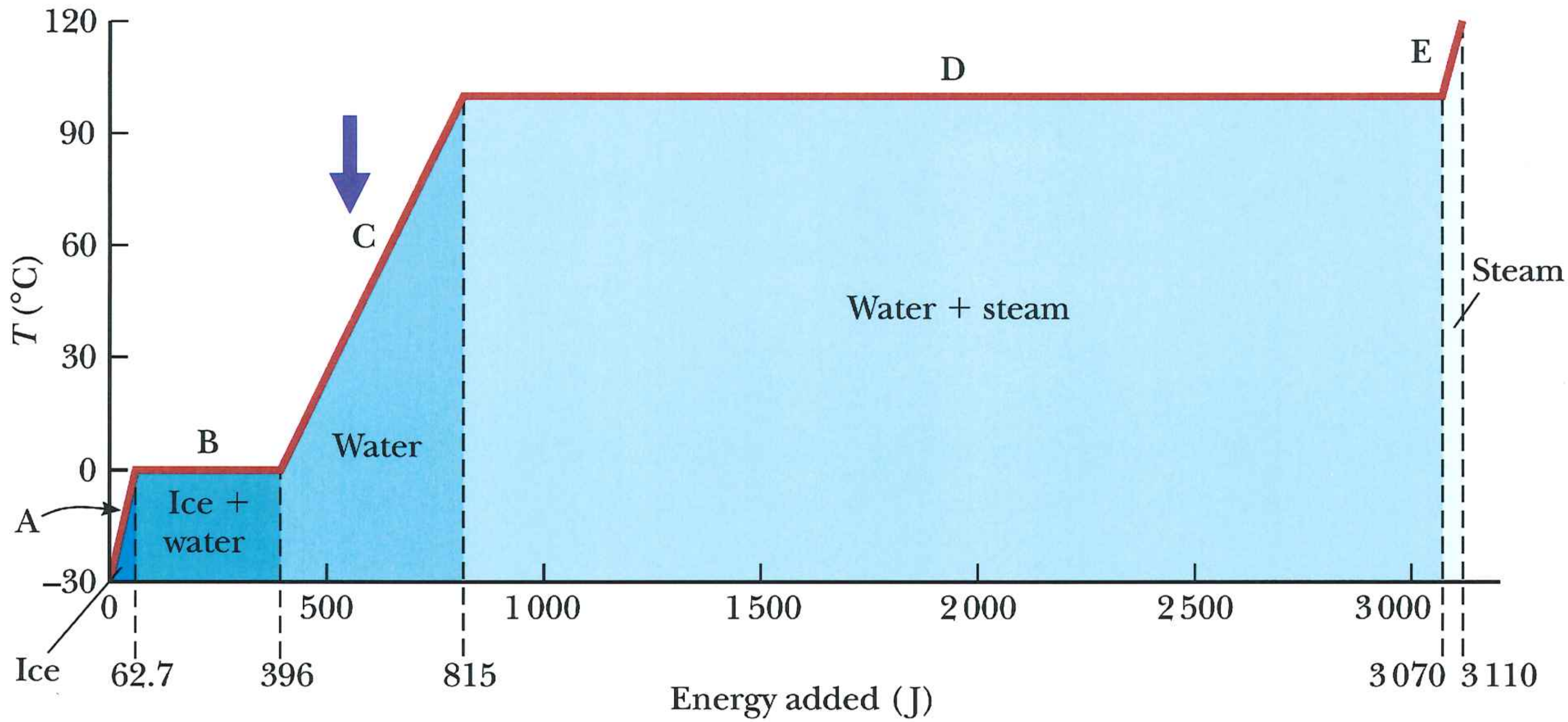
$$Q = mc_{\text{ice}}\Delta T = \left(1.00\times10^{-3}\ \text{kg}\right)\left(2090\ \text{J/kg}\cdot{}^{\circ}\text{C}\right)\left(30.0^{\circ}\text{C}\right) = 62.7\ \text{J}$$

Heat and Internal Energy



$$Q = mL_f = \left(1.00\times10^{-3}\ \text{kg}\right)\left(3.33\times10^{5}\ \text{J/kg}\right) = 333\ \text{J}$$

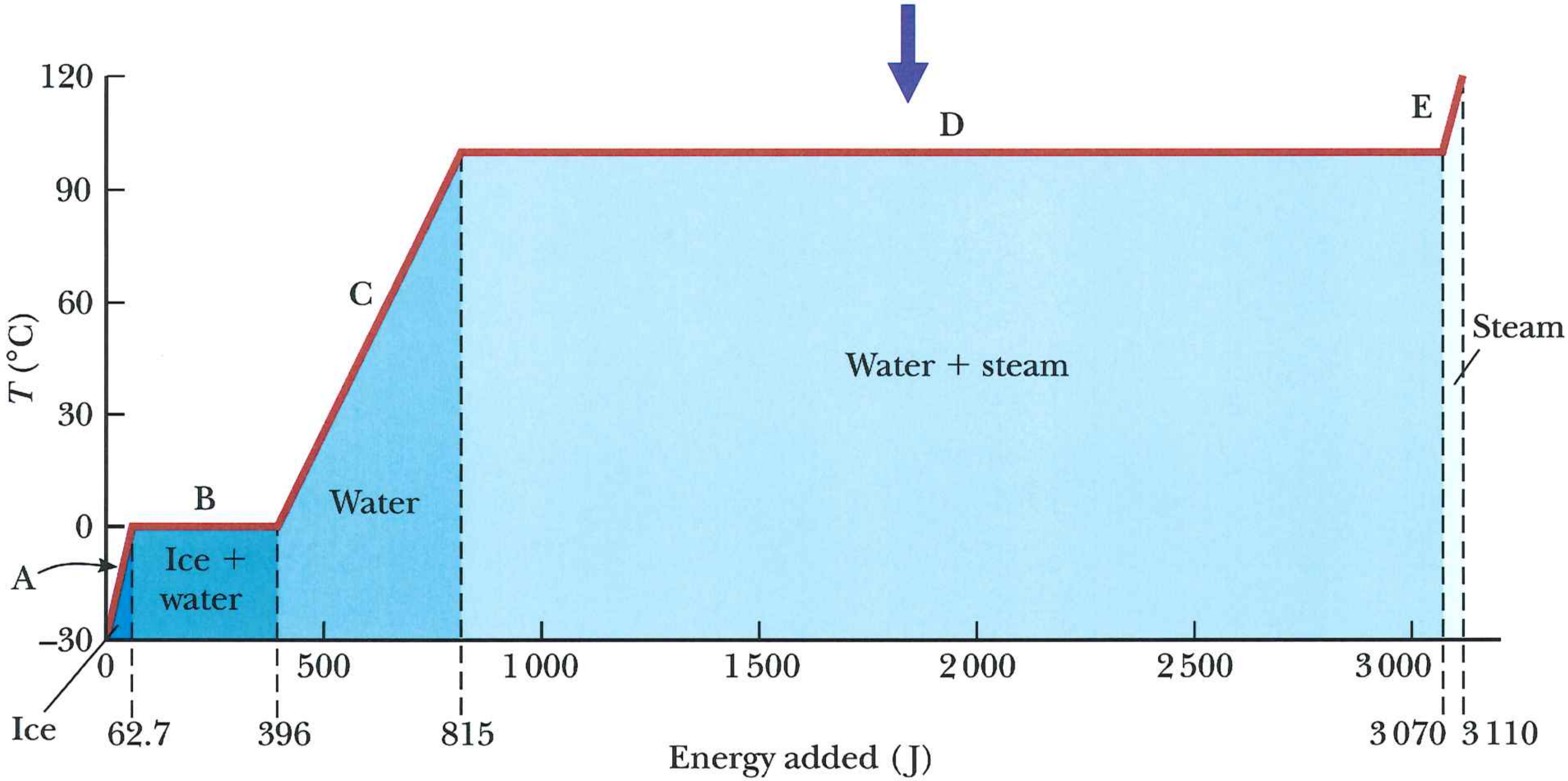
Heat and Internal Energy



$$Q = mc_{water}\Delta T = \left(1.00\times10^{-3}\ \text{kg}\right)\left(4.19\times10^{3}\ \text{J/kg}\cdot{}^{\circ}\text{C}\right)\left(1.00\times10^{2}\ {}^{\circ}\text{C}\right)$$

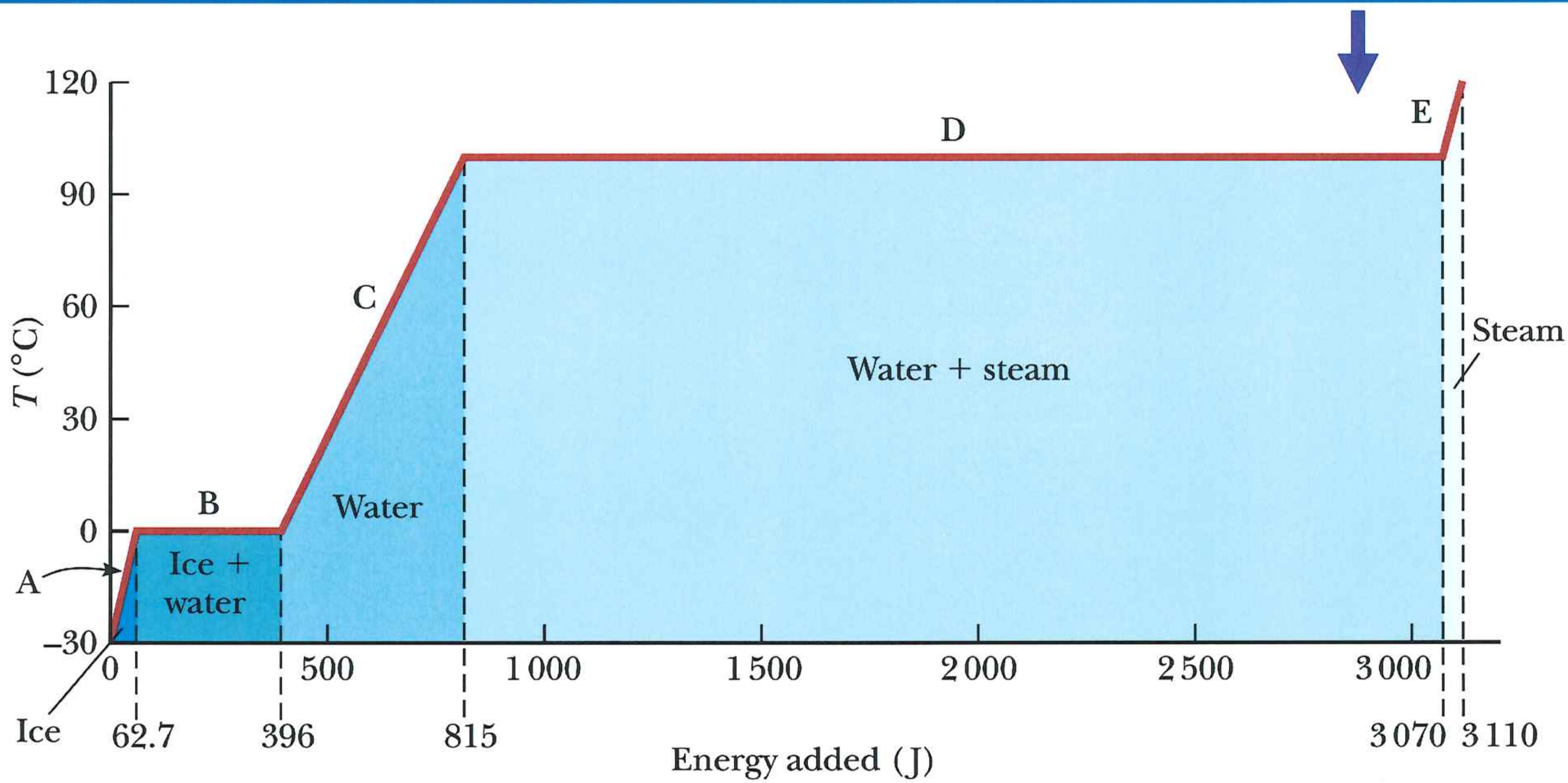
$$= 4.19\times10^{2}\ \text{J}$$

Heat and Internal Energy



$$Q = mL_v = \left(1.00\times10^{-3}\ \text{kg}\right)\left(2.26\times10^{6}\ \text{J/kg}\right) = 2.26\times10^{3}\ \text{J}$$

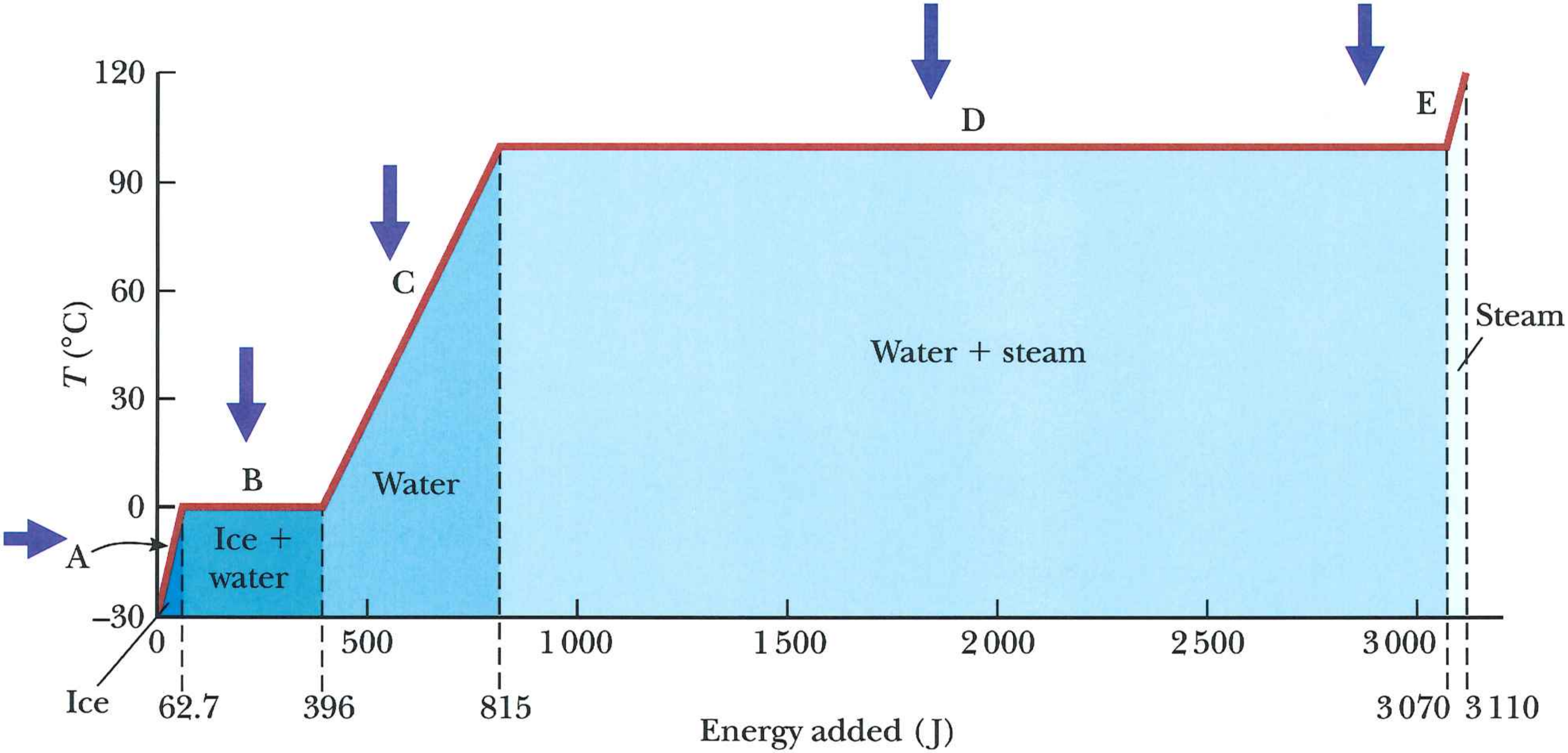
Heat and Internal Energy



$$Q = mc_{\text{steam}}\Delta T = \left(1.00\times10^{-3}\ \text{kg}\right)\left(2.01\times10^{3}\ \text{J/kg}\cdot{}^{\circ}\text{C}\right)\left(20.0^{\circ}\text{C}\right)$$

$$= 40.2\ \text{J}$$

Heat and Internal Energy



$$Q_{\text{total}} = 62.7\text{ J} + 333\text{ J} + 4.19\times10^{2}\text{ J} + 2.26\times10^{3}\text{ J} + 40.2\text{ J}$$

$$= 3.11\times10^{3}\text{ J}$$

Chapter 11: Heat – Phase Changes

Example

Compute the heat of fusion of a substance from the data: 31.15 kJ will change 0.500 kg of the solid at 21 C to liquid at 327 °C, the melting point. The specific heat of the solid is 0.129 kJ/kg K.

$$Q = m c \Delta T + m L_f$$

$$L_f = \frac{Q - m c \Delta T}{m} = 22.8\,\mathrm{kJ/kg}$$