Lecture 43

(Ch13:1)

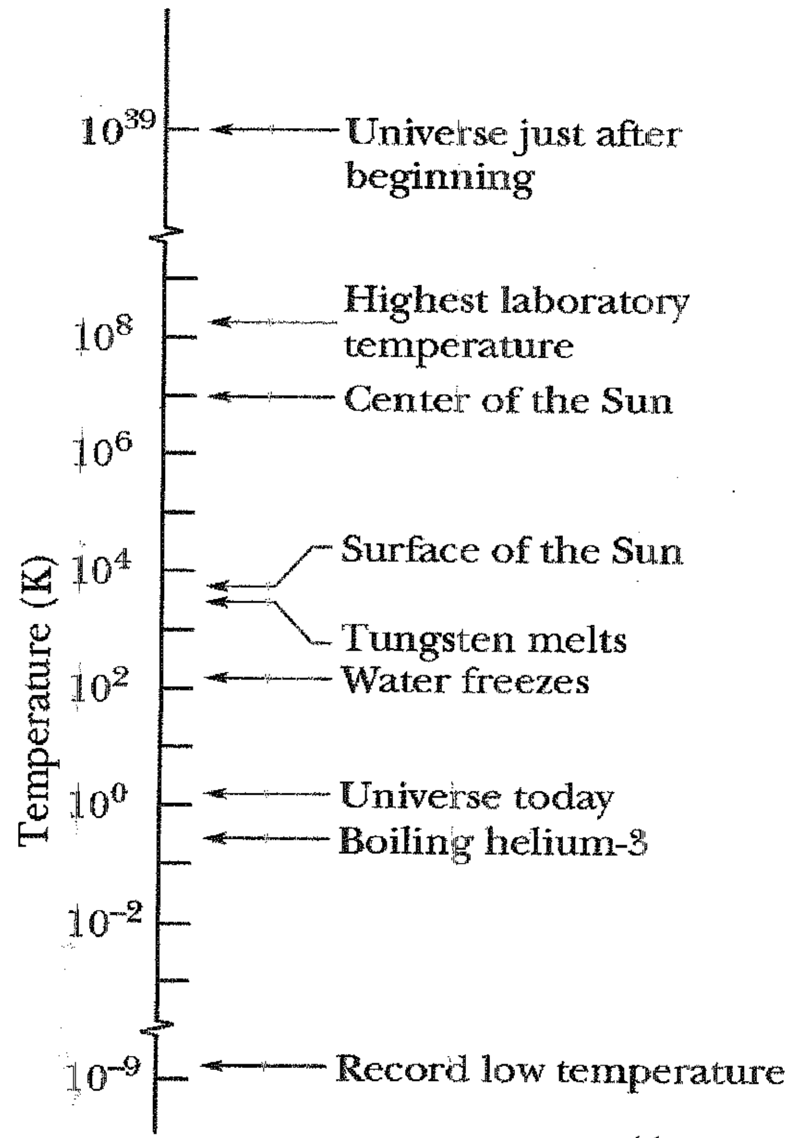
Thermodynamics & Temperature

- **Mechanics**: mechanical energy of systems governed by Newton's laws

- **Thermodynamics:** internal energy of bodies – *thermal energy*

- **Temperature**: central concept - sometimes a measure of the internal energy

  - Kelvin temperature - absolute temperature (see Fig 18-1)



Celsius and Fahrenheit Scales

- Celsius scale: used worldwide

  $T_C = T - 273.15\,K$ → $T[K] = T_C + 273.15$

  - 1° C = 1 K
  - 0° C => ice-water phase transition, normal pressure
  - 100° C => water-vapor phase transition
  - 37° C => normal human body temperature
  - 20° C => normal room temperature (68° F)

WHERE DO WE FEEL HEAT IN DAILY LIFE

1. SUN
2. BOILING WATER
3. STOVE
4. LAMP
5. MAKING A FIRE
6. JUST COOKED FOOD
7. TEA / COFFEE

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

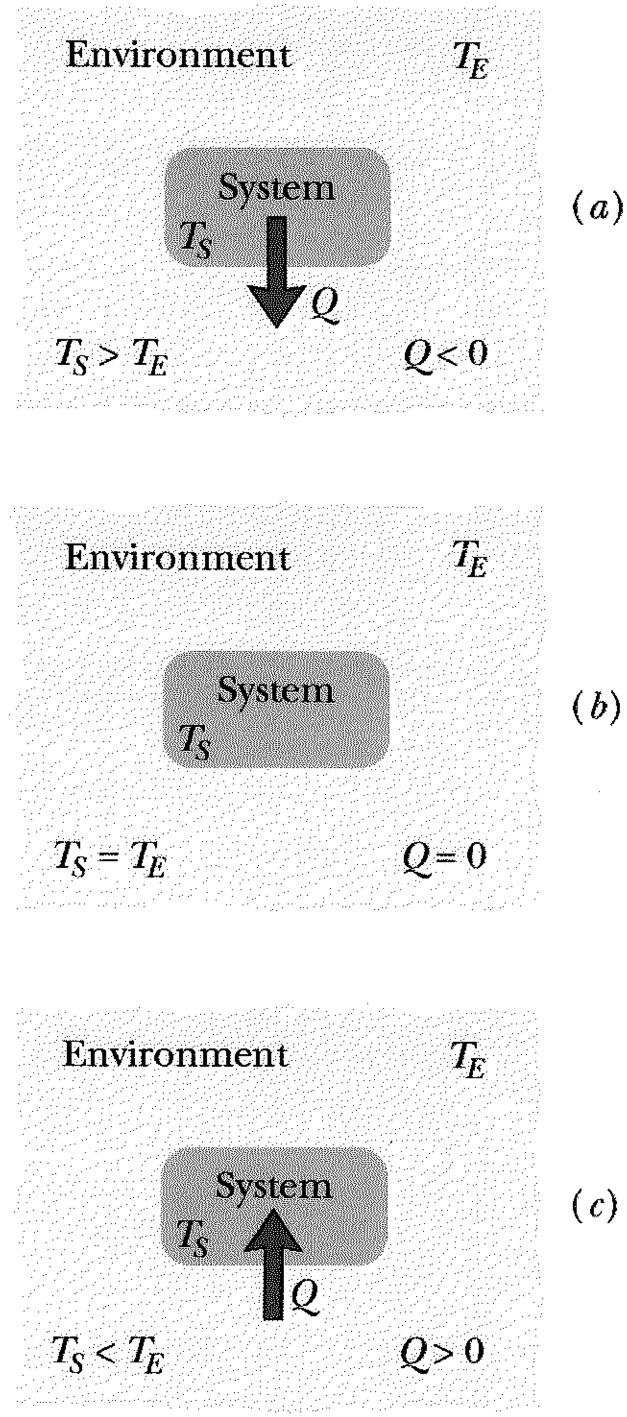


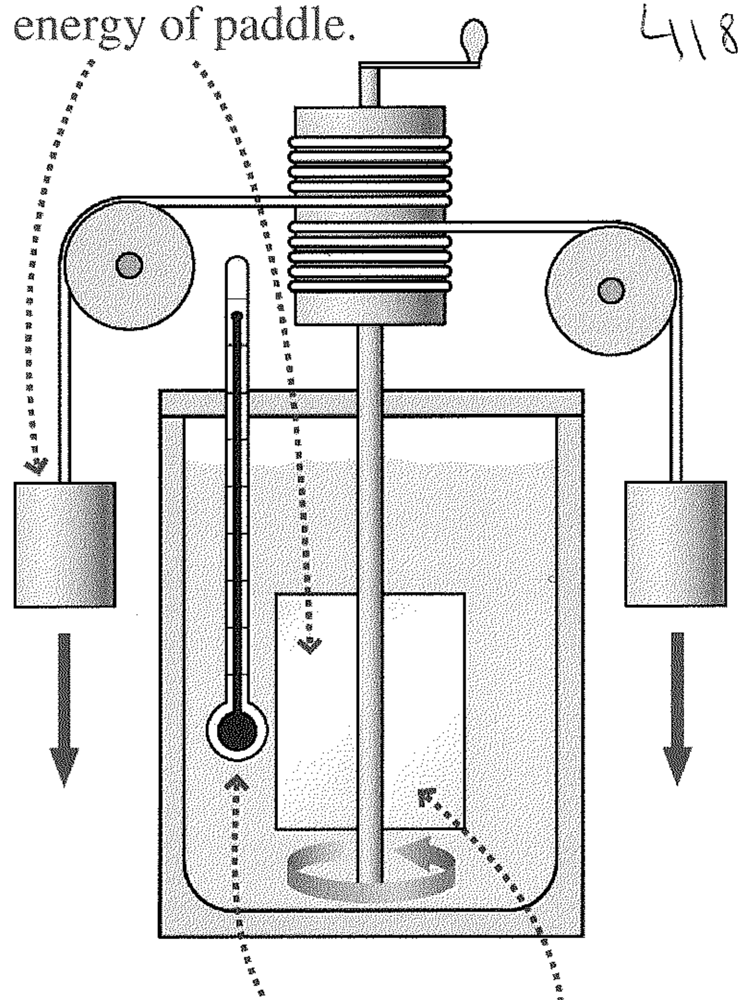
Figure 13.1

Mechanical equivalent of heat

4186 J raise 1kg H_2O by 1°C

James Joule

Potential energy of falling weights becomes kinetic energy of paddle.



The paddle's kinetic energy in turn becomes internal energy of the water, indicated by rising temperature.

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- **Units of heat (Q):**
  - *calorie* (cal): heat 1 gram of water from 14.5^o C to 15.5^o C
  - *British thermal unit* (Btu): heat 1 lb of water from 63^o F to 64^o F
  - *Joule* (J): SI unit ; 1 cal = 4.186 J

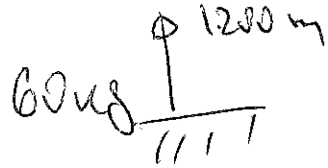
1 cal = 3.969 x 10^{-3} Btu = 4.186 J

1 Food Calorie = 1,000 cal = 4186 J

200 Cal → J ; 1 Food C = 4186 J

22. SOLVE The conversion between food calories to Joules gives:

$$200\ \text{Cal}\left[\frac{4186\ \text{J}}{1\ \text{Cal}}\right] = 837\ \text{kJ}$$



; E=? in Cal

23. SOLVE The change in gravitational potential energy is equal to:

$$U = mgh = (60\text{ kg})(9.80\text{ m/s}^2)(1200\text{ m}) = 706\text{ kJ}\left[\frac{1\text{ Cal}}{4186\text{ J}}\right] = 169\text{ Cal}$$

A 2000-kg car is going 55 mph. Find the thermal energy generated in the car's brakes when stopping the car.

Note 1 mph = (1609.3 m)/(60 min.60 sec) = 0.447 m/s

Therefore 55 mph = (0.447 m/s).(55) = 24.59 m/s

Moving car possesses Kinetic energy $K = (\frac{1}{2}).m.v^2 =$

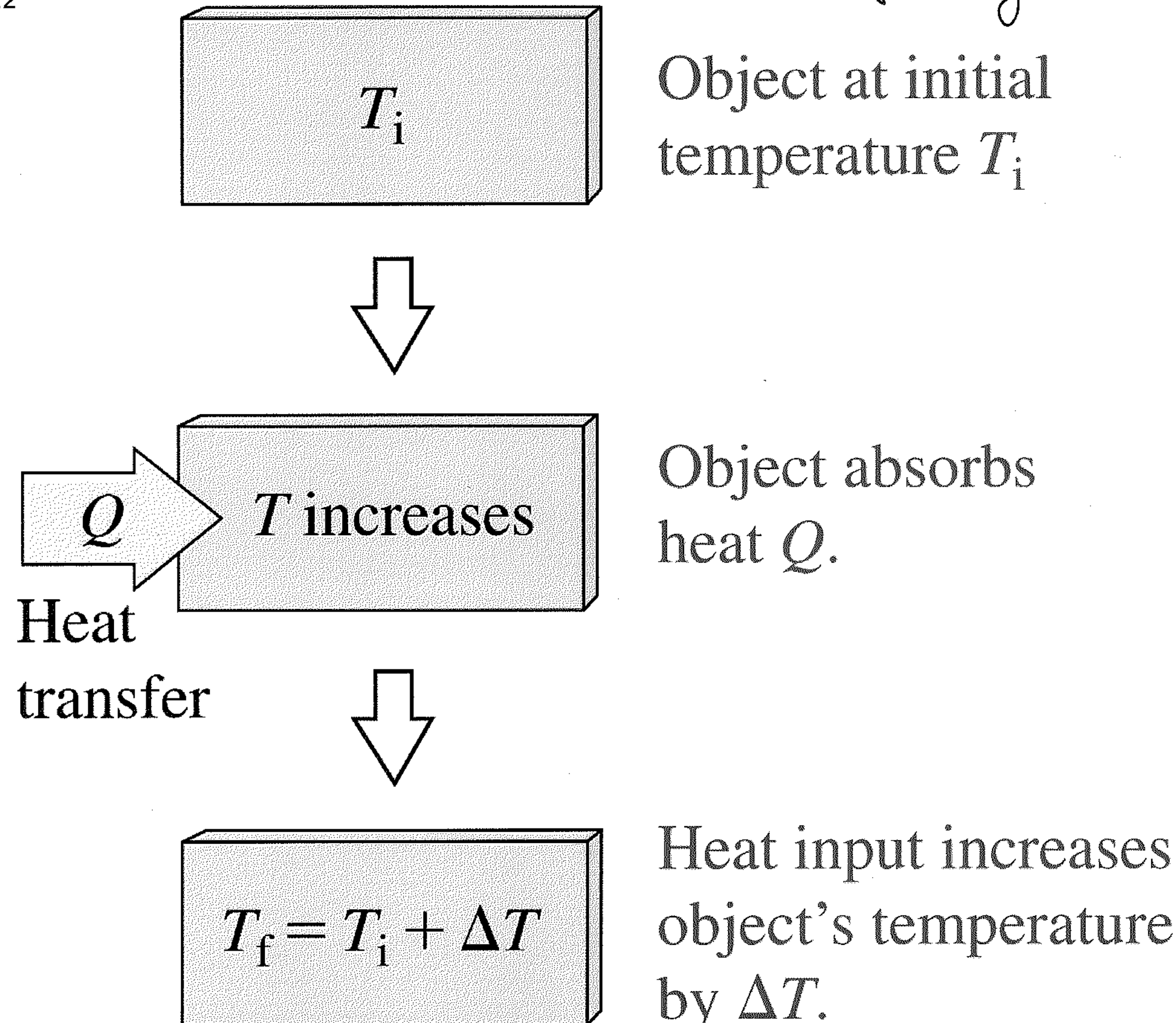
$$= (\tfrac{1}{2}).(2000\ kg).(24.59\ m/s)^2 = 6.045 \times 10^5\ J$$

When the car stops its kinetic energy transforms into thermal energy

So the thermal energy generated in the car's brakes is = 6.045×10^5 J

Heat capacity

Figure 13.2



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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

A piece of metal absorbs 3.6 kJ of heat, increasing its temperature by 33 °C. What is its heat capacity ?

By definition $Q = C\Delta T$ so heat capacity $C=Q/\Delta T$

In our case $Q = 3.6$ kJ and $\Delta T = 33$ °C

Therefore $C = (3600\ J)/(33\ °C) = 109.1\ J/°C$

T↑ 25°C ← Metal ← 2.48 kJ | C = ? Q = ? so T↑ 200 °C

43. ORGANIZE AND PLAN The definition of the heat capacity is from Equation 13.1: $Q = C\Delta T$. The value for C found in part (a) should remain valid for the temperature increase in part (b).

Known: $Q = 2.48\ \text{kJ}$, $\Delta T = 25^\circ\text{C}$ for part (a); $\Delta T = 200^\circ\text{C}$ for part (b).

SOLVE (a) From the definition of heat capacity:

$$C = \frac{Q}{\Delta T} = \frac{2.48\ \text{kJ}}{25^\circ\text{C}} = 99.2\ \text{J}/^\circ\text{C}$$

Q = CΔT

(b) The heat capacity is a constant of the material, so we can use it to find the heat absorbed for other temperature increases:

$$Q = C\Delta T = (99.2\ \text{J}/^\circ\text{C})(200^\circ\text{C}) = 19.8\ \text{kJ}$$

- **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)

$c(H_2O) \rightarrow c_w = 1 cal / g \cdot K = 1 Btu / lb \cdot F = 4190 J / Kg \cdot K$

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

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189 L $$ = ? $\Delta T = 60^\circ - 10^\circ C$; \$0.12 kWh

46. **ORGANIZE AND PLAN** We will need Equation 2.2: $Q = mc\Delta T$, along with the density of water: 1.00 kg/L. Note too that a kilowatt-hour (kWh) is a unit of energy, so we'll need to convert J into kWh.

Known: $V = 189\text{ L}$, $R = \$0.12/\text{kWh}$, $\Delta T = 60^\circ\text{C} - 10^\circ\text{C} = 50^\circ\text{C}$.

SOLVE The 189 L is equivalent to 189 kg, so the heat required to raise its temperature by 50 degrees is:

$$Q = mc\Delta T = (189\text{ kg})(4186\text{ J/kg}\cdot^\circ\text{C})(50^\circ\text{C}) = 3.96\times10^7\text{ J}$$

C_{H_2O}

To convert this to kWh, recall that $1\text{ W} = 1\text{ J/s}$, or equivalently $1\text{ J} = 1\text{ W}\cdot\text{s}$:

$$3.96\times10^7\text{ J}\left[\frac{1\text{ W}\cdot\text{s}}{1\text{ J}}\right]\left[\frac{1\text{ kW}}{1000\text{ W}}\right]\left[\frac{1\text{ h}}{60\cdot60\text{ s}}\right] = 11\text{ kWh}$$ (* \$0.12)

The cost of this much energy is \$1.32.

$P[W] = \frac{W[J]}{\Delta t[s]}$; $1W = 1\frac{J}{s} \rightarrow 1J = 1W * 1s$

2.3 mL Hg  | Q = ? so T ↑ 100 °C
at 0 °C

55. ORGANIZE AND PLAN This is a straightforward use of Equation 13.2: $Q = mc\Delta T$, where the specific heat of mercury is from Table 13.1: $c = 140\ \text{J/kg}\cdot{}^\circ\text{C}$. → c_{Hg}
The one thing we will need is the density of liquid mercury from Table 10.1: $\rho = 13{,}600\ \text{kg/m}^3$. $1\,L = 10^{-3}\,m^3$, $1\,mL = 10^{-3}\,L \rightarrow 1\,mL = 10^{-6}\,m^3$
Known: $V = 2.30$ mL, $\Delta T = 100^\circ\text{C}$.

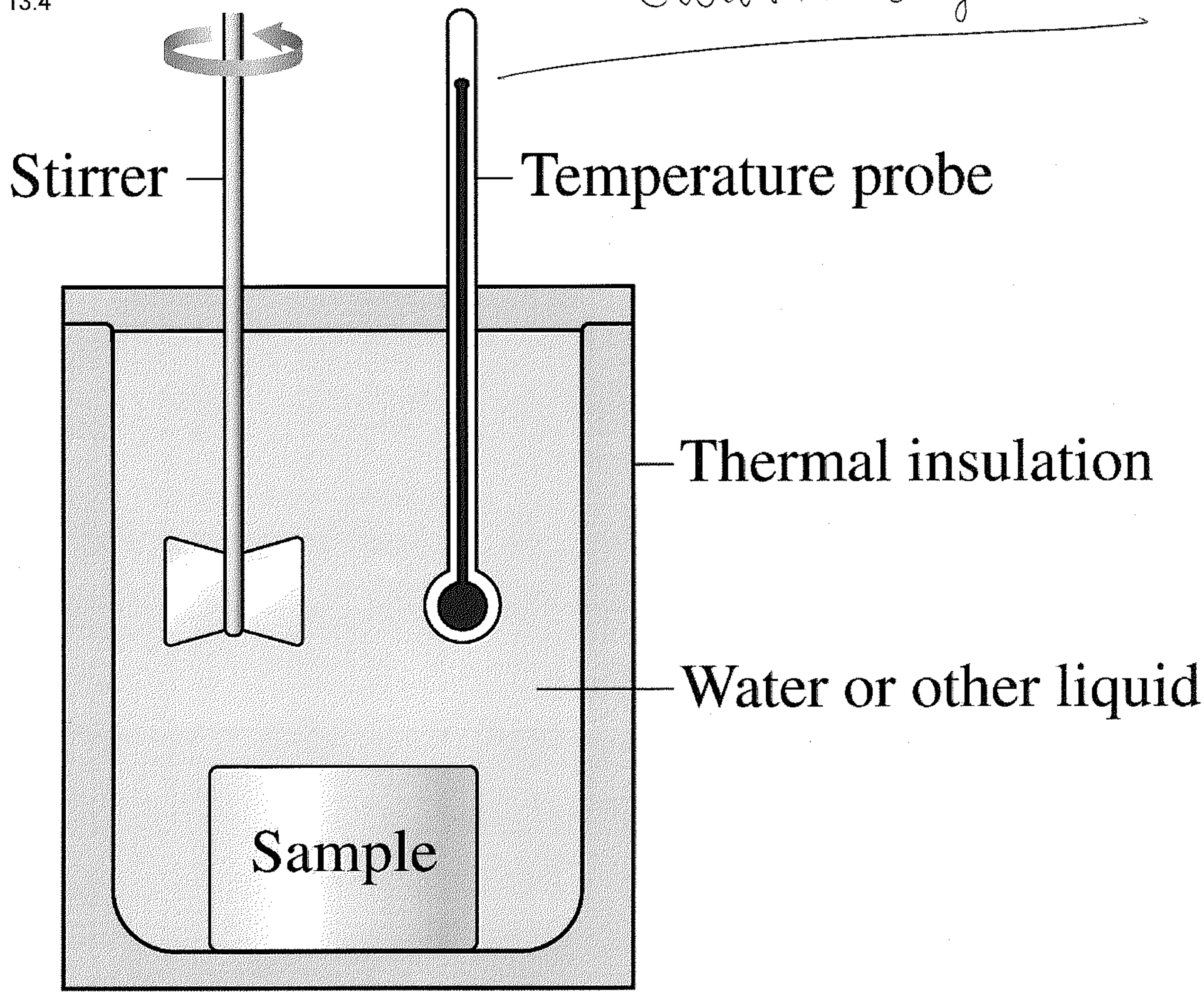
SOLVE Plugging the mass of mercury $(m = \rho V)$ into Equation 13.2:

$$Q = \rho V c \Delta T = (0.0136\ \text{kg/mL})(2.30\ \text{mL})(140\ \text{J/kg}\cdot{}^\circ\text{C})(100^\circ\text{C}) = 438\ \text{J}$$

REFLECT This heat causes the mercury to expand slightly, which results in the liquid rising inside the thermometer. Because this rise is uniform, we can use it to measure the temperature. In this way, the thermometer works simply by absorbing heat (or losing heat) to the environment.

Figure 13.4

Calorimetry



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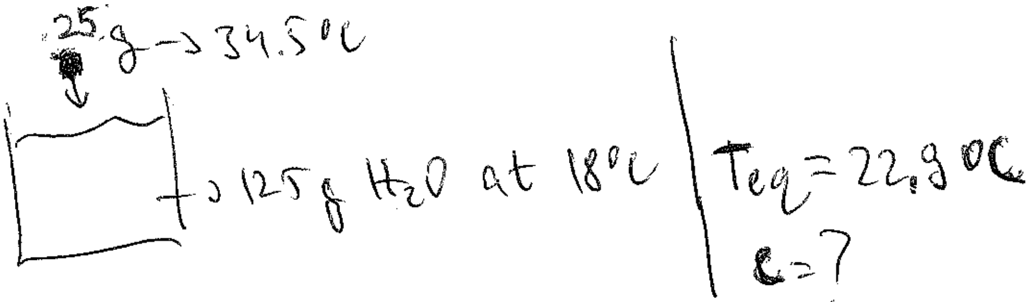

53. ORGANIZE AND PLAN We know that the sum of the heat lost by the material and gained by the water is zero: $Q_M + Q_W = 0$, so we'll use that to solve for the unknown specific heat: c_M. The temperature changes for the material and the water are:

$$\Delta T_M = 22.9°\text{C} - 34.5°\text{C} = -11.6°\text{C}$$
$$\Delta T_W = 22.9°\text{C} - 18°\text{C} = 4.9°\text{C}$$

Known: $m_M = 25.0$ g, $m_W = 125$ g.

SOLVE Using the heat exchange and Equation 13.2: → c_{H_2O}

$$c_M = -\frac{m_W c_W \Delta T_W}{m_M \Delta T_M} = -\frac{(0.125 \text{ kg})(4186 \text{ J/kg}\cdot°\text{C})(4.9°\text{C})}{(0.0250 \text{ kg})(-11.6°\text{C})} = 8841 \text{ J/kg}\cdot°\text{C}$$

$$m_M \Delta T_M c_M + m_W c_W \Delta T_W = 0$$

$$m_M \Delta T_M c_M = -m_W c_W \Delta T_W$$

Joule 817 lb ↓ 1 ft | H_2O | → 1 lb by 1 °F | is this true?

48. ORGANIZE AND PLAN We will use hindsight to see how accurate Joule's measurements were. We'll compare the gravitational potential of 817 pounds at a height of one foot, and compare that to the heat required to raise the temperature of 1 pound of water by 1°F.

Known: $m_{\text{weight}} = 817 \text{ lb}, \ h = 1 \text{ ft}, \ m_{\text{water}} = 1 \text{ lb}, \ \Delta T = 1°\text{F}$

SOLVE (a) The gravitational energy of the weight in Joule's experiment was:

$$U = m_{\text{weight}} gh = (817\text{lb})\left[\frac{0.457 \text{ kg}}{1 \text{ lb}}\right](9.80 \text{ m/s}^2)(1 \text{ ft})\left[\frac{0.305 \text{ m}}{1 \text{ ft}}\right] = 1116 \text{ J}$$

This energy went into turning a paddle wheel and thereby heated the pound of water. Using the specific heat of water, the heat that would be needed to raise the water's temperature by 1°F:

$$Q = m_{\text{water}} c\Delta T = (1 \text{ lb})\left[\frac{0.457 \text{ kg}}{1 \text{ lb}}\right](4186 \text{ J/kg°C})(1°\text{F})\left[\frac{1°\text{C}}{1.8°\text{F}}\right] = 1063 \text{ J}$$

$T°C = \frac{5}{9}[T°F - 32]$

$\frac{9}{5} 1°C = 1°F - 32$

The discrepancy between these values is:

$$\frac{U-Q}{U} = \frac{1116 \text{ J} - 1063 \text{ J}}{1116 \text{ J}} = 5\%$$

(b) What mass should have Joule used to generate the expected heat input, assuming that the weight again falls 1 foot?

$$m_{\text{weight}} = \frac{Q}{gh} = \frac{1063 \text{ J}}{(9.80 \text{ m/s}^2)(1 \text{ ft})}\left[\frac{1 \text{ ft}}{0.305 \text{ m}}\right] = 356 \text{ kg}\left[\frac{1 \text{ lb}}{0.457 \text{ kg}}\right] = 779 \text{ lb}$$

REFLECT There are several possible reasons why Joule found a larger mass than necessary in his experiment. Some of the potential energy of the weight may have been lost, perhaps to friction in the mechanical device, or to the heating of the surrounding. If so, only about 95% of the potential energy wound up heating the water.