Thermal Properties

Temperature
What is temperature? It is a measure of the amount of "atomic jiggling". When something is hot (has a high temperature), its atoms are jiggling a lot. When it is cold (has a low temperature), its atoms are jiggling little.

As temperature falls, atoms jiggle less and less. At "absolute zero" T = 0 K, all atoms stop, no motion.

Temperature T = measure of energy per atom

Various temperature scales:

\[ T_F = \frac{9}{5} T_C + 32 \]
\[ T_K = T_C + 273.15 \]

°F = Fahrenheit
°C = Celsius
K = Kelvin

1 °C = 1 K,  
1 °F = (5/9) °C

room temperature = 72°F = 22°C = 295 K ≈ 300 K

absolute zero = 0 K = −273°C = −459°F

[In the ideal gas law, \( p V = N k T = n R T \) (N = #molecules, n = #moles), must always use T in Kelvin.]

Thermal energy \( U \) = total energy of all atoms (random motion)

Heat \( Q \) = amount of thermal energy transferred to a body.

\[ [Q] = \text{energy, SI unit of heat} = \text{joule} \]

popular unit of energy = 1 calorie (cal) = 4.184 J Notice calorie spelled with a small "c".
1 cal = energy to raise T of 1 gram of water by 1°C

1 kcal = 1000 cal = 1 Cal = 4184 J = "food Calorie" Notice Calorie spelled with a big "C"

[ Some primitive cultures use the BTU (British Thermal Unit) = energy to raise a pound of water by 1°F. 1 BTU = 1060 J ]

**Definition:** heat capacity of an object = heat added per temperature rise (J/K)

**Definition:** specific heat (or specific heat capacity) of a material = c = amount of heat added per unit mass per degree Celsius rise in temperature. If we have a mass m of some material, and we add an amount of heat $\Delta Q$ and that produces a temperature rise of $\Delta T$, then specific heat is defined as..

$$c = \frac{\Delta Q}{m \Delta T} \quad [c] = J/(kg \cdot ^\circ C) \quad \text{(SI units)}$$

Usually write the equation as $\Delta Q = m \cdot c \cdot \Delta T$

This equation says that if I have a mass m of some material with specific heat c, and I want to raise its temperature by $\Delta T$, then I have to add an amount of heat $\Delta Q = m \cdot c \cdot \Delta T$.

$$c_{\text{water}} = \frac{1 \text{ cal}}{g \cdot ^\circ C} = \frac{1 \text{ kcal}}{kg \cdot ^\circ C} = \frac{4186 \text{ J}}{kg \cdot ^\circ C}$$

**Example:** Heat your mug of coffee (which is mostly water) from room temperature to near boiling: m = 200 g, T = 20°C → 90°C.

$$\Delta Q = m \cdot c \cdot \Delta T = (200g) \left(1 \frac{\text{cal}}{g \cdot ^\circ C}\right)(70^\circ C) = 14000 \text{ cal} = 14.0 \text{ kcal} \times \left(\frac{4186 \text{ J}}{\text{kcal}}\right) = 58600 \text{ J}$$

Different materials have different specific heats:

<table>
<thead>
<tr>
<th>material</th>
<th>c (cal/g·°C)</th>
</tr>
</thead>
<tbody>
<tr>
<td>water</td>
<td>1.00</td>
</tr>
<tr>
<td>ice</td>
<td>0.53</td>
</tr>
<tr>
<td>aluminum</td>
<td>0.22</td>
</tr>
<tr>
<td>dry air</td>
<td>0.24</td>
</tr>
<tr>
<td>iron</td>
<td>0.11</td>
</tr>
</tbody>
</table>

(notice that liquid water has a high specific heat compared to other materials)
Example: Suppose we have 2 objects, labeled A and B (water and steel, say), with object A hotter than object B. They initially have temperatures $T_A$ and $T_B$.

\[
\begin{array}{c|c}
\text{(hot)} & \text{(cold)} \\
A & B \\
m_A, T_A, c_A & m_B, T_B, c_B \\
\end{array}
\]

$T_A > T_B$

Bring A and B together, allowing them to exchange heat with each other, but not with the outside world $\Rightarrow$ A will cool, B will heat and both will reach same final temperature $T_f$.

Object A will lose heat: $\Delta Q_A < 0$
Object B will gain heat: $\Delta Q_B > 0$

\[
\Delta Q_A = -\Delta Q_B \\
=m_A c_A \Delta T_A = -m_B c_B \Delta T_B \\
m_A c_A (T_f - T_A) = -m_B c_B (T_f - T_B) \\
T_f (m_A c_A + m_B c_B) = m_A c_A T_A + m_B c_B T_B
\]

…solve for $T_f$ (does not matter if T is in Celsius or Kelvin, but must be consistent).

**Phase changes.**

phase = solid, liquid, or gas (S, L, or G)

S $\leftrightarrow$ L (freezing/melting) or L $\leftrightarrow$ G (boiling/condensing) or S $\leftrightarrow$ G (sublimation)

Solid water (ice) can have any temperature in the range $-273^\circ C < T \leq 0^\circ C$

Liquid water can have any temperature in the range $0^\circ C \leq T < 100^\circ C$

Can have a mixture of ice and water both at $T = 0^\circ C$

If heat is added to the mixture at $T = 0^\circ C$, some ice melts, but T stays at $0^\circ C$ until all the ice has melted.

*Latent heat or heat of transformation* = heat required to cause phase change

Latent heat of solid/liquid trans. $L_{SL}$ = heat needed to melt 1 g of ice at $0^\circ C$.

$L_{SL}$ (water) = 79.7 cal/g

Requires 80 cal to melt a single gram of ice, but only 1 cal to raise temp of the liquid by $1^\circ C$. 

**Example:** How much heat required to change 100 g of ice at $T = -10^\circ C$ into liquid water at $T = +10^\circ C$?

1. Heat ice to $T = 0^\circ C$  
   \[ \Delta Q_1 = m c_{\text{ice}} \Delta T \]
2. Melt ice at $T = 0^\circ C$  
   \[ \Delta Q_2 = m L_{\text{SL}} \]
3. Heat water to $T_{\text{final}}$  
   \[ \Delta Q_3 = m c_{\text{water}} \Delta T \]

\[ \Delta Q_{\text{total}} = 100(0.5)(10) + 100(80) + 100(1)(10) \]
\[ = 500 \text{ cal} + 8000 \text{ cal} + 1000 \text{ cal} = 9500 \text{ cal} \]

(heat ice)  
(melt ice)  
(heat water)

Note that most of the energy went into melting the ice because of the large latent heat of water/ice transformation. This is good for people in Boulder. If $L_{\text{SL}}$ was not large, we would have big floods every spring, because all the snow would suddenly melt as soon as the temperature rose above melting.

$L_{\text{LG}} = \text{heat of vaporization} = \text{heat needed to transform 1 g of liquid water into vapor at 100^\circ C}$.  
$L_{\text{LG}} (\text{water}) = 539 \text{ cal}$. This is a very large amount of heat ⇒ very expensive to distill water.

**Example:** Tiger, tiger, burning bright... In the 1956 science fiction movie, *Forbidden Planet*, Captain Adams (played by Leslie Nielsen) vaporizes a tiger with one shot from his "blaster pistol". This tiger is only about 6 meters away from the captain. About how much energy is required to vaporize a tiger? Is it a good idea to release this much energy this close to you?

A tiger is mostly water and has a mass of about $m = 250 \text{ kg}$ (three time the mass of a man). In order to make the tiger boil away, you have to first raise the temperature of the tiger (water) from $T = 30^\circ C$ (healthy tiger temp.) to $100^\circ C$ (boiling). Then you have to evaporate the water at $T = 100^\circ C$. For each gram of tiger, the first step requires 70 cal ($= m c \Delta T$), and the second step requires 539 cal ($= m L_{\text{LG}}$), so let's say, roughly, at least 600 cal is needed per gram.

\[ Q = 250 \text{ kg} \times \frac{1000 \text{ g}}{\text{kg}} \times \frac{600 \text{ cal}}{\text{g}} \times \frac{4 \text{ J}}{\text{cal}} = 6 \times 10^8 \text{ J} \]  

(just a rough calculation so 1 cal ≈ 4 J.)

How much energy is this $6 \times 10^8 \text{ J}$? This energy is about 200 kW·hr. [One kilowatt-hour is 1000W (ten 100W light bulbs on) for 1 hour.] The power company charges about $20 for this much energy (at 10 cents per kW·hr). This energy is also the energy content of about 5 gallons of gasoline or about 300 sticks of dynamite. Releasing this much energy all at once would kill everyone nearby and make a huge, choking cloud of tiger smoke.

**Heat Transfer**

There are three (and only three) ways to transfer heat.

1) Conduction : heat transfer by direct touch
2) Convection : heat transfer by bulk movement of hot matter
3) Radiation : heat transfer by light (electromagnetic radiation)