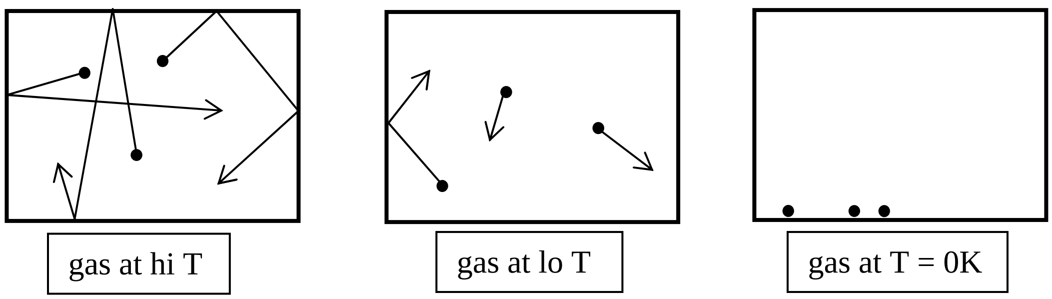
Thermal Properties

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:

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

^oF = Fahrenheit
^oC = Celsius
K = Kelvin

1 ^oC = 1 K , 1 ^oF = (5/9) C^o

room temperature = 72^{o}F = 22^{o}C = 295 K ≈ 300 K

absolute zero = 0 K = –273^{o}C = – 459^{o}F

[In the ideal gas law, p V = N k T = n R T (N = #molecules, n = #moles), <u>must</u> 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] = energy, SI unit of heat = 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^{o}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^{o}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 ΔQ and that produces a temperature rise of ΔT, then specific heat is defined as..

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

Usually write the equation as $\boxed{\Delta Q = m\,c\,\Delta T}$

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 ΔT, then I have to add an amount of heat $\Delta Q = m\,c\,\Delta T$.

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

Example: Heat your mug of coffee (which is mostly water) from room temperature to near boiling: m = 200 g, T = 20^{o}C $\rightarrow$ 90^{o}C.

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

Different materials have different specific heats:

material	c (cal/g·C)
water	1.00
ice	0.53
aluminum	0.22
dry air	0.24
iron	0.11

(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 .

(hot) (cold)

m_A , T_A , c_A m_B , T_B , c_B

$T_A > T_B$

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: $\Delta Q_B > 0$

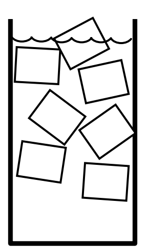
$$\begin{aligned} \Delta Q_A &= -\Delta Q_B \\ m_A\, c_A\, \Delta T_A &= -m_B\, c_B\, \Delta T_B \\ m_A\, c_A (T_f - T_A) &= -m_B\, c_B (T_f - T_B) \\ T_f (m_A c_A + m_B c_B) &= m_A c_A T_A + m_B c_B T_B \end{aligned}$$

…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 $-273^oC < T \le 0^oC$
Liquid water can have any temperature in the range $0^oC \le T < 100^oC$

Can have a <u>mixture</u> of ice and water <u>both</u> at $T = 0^oC$


ice + water

If heat is added to the mixture at $T = 0^oC$, some ice melts, but T stays at 0^oC 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^oC.

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^oC.

Example: How much heat required to change 100 g of ice at T = –10^{o}C into liquid water at T = +10^{o}C?

1. Heat ice to T = 0^{o}C $\Delta Q_1 = m\, c_{ice}\, \Delta T$
2. Melt ice at T = 0^{o}C $\Delta Q_2 = m\, L_f$
3. Heat water to T_{final} $\Delta Q_3 = m\, c_{water}\, \Delta T$

ΔQ_{total} = 100(0.5)(10) + 100(80) + 100(1)(10)

= 500 cal (heat ice) + 8000 cal (melt ice) + 1000 cal (heat water) = 9500 cal

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_f 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_{LG} = *heat of vaporization* = heat needed to transform 1 g of liquid water into vapor at 100^{o}C.
L_{LG} (water) = 539 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 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^{o}C (healthy tiger temp.) to 100^{o}C (boiling). Then you have to evaporate the water at T = 100^{o}C. For each gram of tiger, the first step requires 70 cal (= m c ΔT), and the second step requires 539 cal (= m L_{LG}), so let's say, roughly, at least 600 cal is needed per gram.

$$Q \;=\; 250\text{ kg}\times\frac{1000\text{ g}}{\text{kg}}\times\frac{600\text{ cal}}{\text{g}}\times\frac{4\text{ J}}{\text{cal}} \;=\; 6\times10^{8}\text{ J}$$ (just a rough calculation so 1 cal ≈ 4 J.)

How much energy is this 6×10^8 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)