Heat and Temperature

So Far
- Temperature vs Heat
- Measuring temperature
- Temperature scales and absolute zero
- Thermal expansion

To Come
- Heat capacity (absorption)
  - Specific heat
- phase transitions
- Heat and Work
- 1st law of thermodynamics
- heat transfer
  - conduction
  - convection
  - radiation
- Kinetic Theory of Gases

Today
Expansion with Temperature

Molecular explanation:

- As temperature of solid increases, molecules move faster around average locations further from each other (expansion).
- High enough temp (kinetic energy), bonds holding atoms in lattice break (liquid), but forces still hold atoms nearby (not fixed average location like solid) and more temp makes for greater separations (expansion).
- Raise temp even higher, bonds completely break --- gas --- more temp, more energy (expansion).

Solid  |  Liquid  |  Gas
Absolute Zero
Temperature

- **Fixed** amount of gas and **fixed** volume of bulb
- find empirically pressure and temperature proportional to each other
  \[ T = C_p \]

- extrapolate to lower temperatures
- all gases intercept at \( T_0 = -273.15^\circ C = \text{absolute zero} \)
Temperature Scales and Expansion

- Three temperature scales
- Mass of any system stays constant (until Relativistic Quant Mech)
- Empirically as temperature increases, most materials increase their volume
- Coefficient of linear expansion, $\alpha$ (small)

$$\Delta L = L \alpha \Delta T$$

$$\Delta A \equiv A \gamma \Delta T \quad ; \quad \gamma = 2\alpha$$

$$\Delta V \equiv V \beta \Delta T \quad ; \quad \beta = 3\alpha$$
Dangers - Diff Expansion

- Pour boiling water into beaker - inside expands faster than outside
- Train tracks laid too close together

**TABLE 10.1 Average Coefficients of Expansion for Some Materials Near Room Temperature**

<table>
<thead>
<tr>
<th>Material</th>
<th>Average Coefficient of Linear Expansion ((\text{[}^\circ\text{C}]^{-1}))</th>
<th>Material</th>
<th>Average Coefficient of Volume Expansion ((\text{[}^\circ\text{C}]^{-1}))</th>
</tr>
</thead>
<tbody>
<tr>
<td>Aluminum</td>
<td>24 \times 10^{-6}</td>
<td>Ethyl alcohol</td>
<td>1.12 \times 10^{-4}</td>
</tr>
<tr>
<td>Brass and bronze</td>
<td>19 \times 10^{-6}</td>
<td>Benzene</td>
<td>1.24 \times 10^{-4}</td>
</tr>
<tr>
<td>Copper</td>
<td>17 \times 10^{-6}</td>
<td>Acetone</td>
<td>1.5 \times 10^{-4}</td>
</tr>
<tr>
<td>Glass (ordinary)</td>
<td>9 \times 10^{-6}</td>
<td>Glycerin</td>
<td>4.85 \times 10^{-4}</td>
</tr>
<tr>
<td>Glass (Pyrex®)</td>
<td>3.2 \times 10^{-6}</td>
<td>Mercury</td>
<td>1.82 \times 10^{-4}</td>
</tr>
<tr>
<td>Lead</td>
<td>29 \times 10^{-6}</td>
<td>Turpentine</td>
<td>9.0 \times 10^{-4}</td>
</tr>
<tr>
<td>Steel</td>
<td>11 \times 10^{-6}</td>
<td>Gasoline</td>
<td>9.6 \times 10^{-4}</td>
</tr>
<tr>
<td>Invar (Ni-Fe alloy)</td>
<td>0.9 \times 10^{-6}</td>
<td>Air</td>
<td>3.67 \times 10^{-3}</td>
</tr>
<tr>
<td>Concrete</td>
<td>12 \times 10^{-6}</td>
<td>Helium</td>
<td>3.665 \times 10^{-3}</td>
</tr>
</tbody>
</table>
Expansion Issues

- Expansion of solids creates difficulties
- Note real-life solution on bridges, ...
Special Case of Water

- Note ice density \textit{much} smaller
  - $\rho_{\text{ice}} = 0.917 \text{ g/cm}^3$ \ldots \underline{phase change}
- Lakes and ponds in winter: ice floats, insulates water below; warmest water at bottom
  - Summer: water coolest at bottom

- Normal=
  - system of fixed mass increases its volume when temperature increases
- Density normally decreases as $T$ increases
- Water has peculiar property that density increases below $T = 4^\circ\text{C}$
- Water at the bottom of lakes is often at $T = 4^\circ\text{C}$ even in winter
Heat Capacity

- System at top in equilibrium with environment
- System at bottom at lower temperature. Heat energy, $Q$, will flow into system and temperature will rise
- For fixed $\Delta T$, amount of heat, $Q$ depends on system type and mass

\[ Q = C \Delta T \]
\[ Q = cm \Delta T \]

Note convention: $Q > 0$ heat added
$Q < 0$ heat removed

$C = $ heat capacity
$c = $ specific heat
Specific Heats

- Measure of how much heat energy it takes to raise unit mass one degree temperature
- Examples in table 19-3
- Note units and examples: water and copper or iron (more typical)
  - $c_w = 1 \text{ cal/g-°C} = 1 \text{ Btu/lb-°F}$
  - $c_w = 4190 \text{ J/kg-K}$
  - $c_{Cu} = 386 \text{ J/kg-K}$
  - $c_{Fe} = 460 \text{ J/kg-K}$
- Molar specific heat multiplies by the number of kg in a mole = $0.001 \times$ molecular wt
  - Note similar values - elements in table
  - 1 mole has $6.02 \times 10^{23}$ molecules (see later)

$$Q \equiv c_m \Delta T$$

Again, notice that water is different!!
Heat changing temperature
Sample prob 19-4

- Beaker: $C_b \ T_i$
- Water: $m_w \ T_i \ c_w$
- Copper: $m_c \ T_c \ c_c$
- Drop copper in water
- What is final temp, $T_f$?
- Intuitively, $T_i < T_f < T_c$

$$Q_w = c_w m_w (T_f - T_i)$$
$$Q_b = C_b (T_f - T_i)$$
$$Q_c = c_c m_c (T_f - T_c)$$
$$Q_w + Q_b + Q_c = 0$$

$$T_f = \frac{c_c m_c T_c + (C_b + c_w m_w) T_i}{c_c m_c + C_b + c_w m_w}$$
Changing phases ... heat required

- To go from **more ordered** state to less ordered state requires adding energy
- This added energy supplies the necessary potential energy of the bonding (or breaks the atomic bonds holding atoms together)
Heat of transformation

- Materials can exist in different “states” or phases. Typically
  - Solid → liquid at $T=\text{melting point}$
  - Liquid → gas at $T=\text{boiling point}$
  - Example water

- Go from **most ordered** (solid) to **least ordered** (gas) as heat added

- Certain amount of heat required to change the state
  - Heat of fusion, $L_F$...solid to liquid
  - Heat of vaporization, $L_V$...liq to gas

- See table 19-4 in text for examples
  - Most notable: WATER
    - $L_F = 333 \text{ kJ/kg}$
    - $L_V = 2256 \text{ kJ/kg}$
State or Phase Changes

Motions of Molecules

- If forces holding together a solid are more tightly bound, might expect
  - Melting temp higher
  - Expansion rate with temperature slower
- Correlation should exist!

Correlation exists for most metals
Heat changing the phase of water

- Water molecules vs ice molecules (avg locations)
- Behavior of 1 kg of water vs temp

- Fusion: 80 kcal, 333 kJ
- Vaporization: 539 kcal, 2256 kJ

$T \ (°C)$

Physics 1401 - L 20

Frank Sciulli
Problem 46 (not assigned)

An insulated Thermos contains 130 cm$^3$ of hot coffee, at a temperature of 80.0°C. You put in a 12.0 g ice cube at its melting point to cool the coffee.

By how many degrees has your coffee cooled once the ice has melted? Treat the coffee as though it were pure water and neglect energy transfers with the environment.
Back to Gases: Heat Doing Work ...

Work depends on how gas expands

\[
dW = F \cdot ds = pA \, ds
\]
\[
dW = p \, dV
\]

System of gas above: i→f
- insulated from outside world except reservoir
- reservoir supplies heat at rate determined by knob
- as piston on top rises, gas expands ( & vice versa)
- Can vary pressure and volume (with temperature and weights)

- Note that heat \( Q \) supplied to gas has a portion \( W \) doing work
- Work is area under \( p-V \) curve
- direction of process (sign \( W \)) determines whether work done by or on the system
First Law of Thermodynamics restricts the possibilities

First Law of Thermodynamics = Energy Conservation

Energy Conservation has three contributions

- $W = \text{work done (+) by system}$
- $Q = \text{thermal energy (+) added to system}$
- $E_{\text{int}} = \text{internal energy of system}$

$E_{\text{int}}$ is a consequence of the system state (e.g., could be a function of $P, V, T, \ldots$)

\[ dE_{\text{int}} = dQ - dW \]

Heat energy added to the system less the work done by the system equals the increase in the system's internal energy.
Example processes – see table 19-5

\[ dE_{\text{int}} = dQ - dW \]

- Adiabatic \( Q=0 \)
- Constant volume \( W=0 \)
- Cyclical \( \Delta E_{\text{int}}=0 \)
- Free expansion \( Q=W=0 \)
Sample Prob 19-5
work it through

- 1 kg boiling water at 1 atmosphere (1.01 \times 10^5 \text{ Pa})
- vol of water (10^{-3} \text{ m}^3)
  changes to steam (1.671 \text{ m}^3)
- What changes & how much?

\[ Q = L_f m \quad L_f = 2256 \text{ kJ/kg} \]
\[ W = \int_{V_i}^{V_f} p \, dV = p (V_f - V_i) \]
\[ \Delta E_{\text{int}} = Q - W \]

\[ Q = 2256 \text{ kJ} = 2.26 \text{ MJ} \]
\[ W = 1.69 \times 10^5 \text{ J} = 0.17 \text{ MJ} \]
\[ \Delta E_{\text{int}} = Q - W = 2.09 \text{ MJ} \]
Processes and Cycles

- In the above processes, work is done, and the gas state might also change
  - if $E_{\text{int}}$ is different at $i$ than at $f$
- Cycles are processes that return the gas to the starting point, have the gas end up in the same state as initially.
  - Note that, even though the gas is in the same state, net work can be done in the cycle
  - Some of the heat energy is converted to work
Problem 49 (not assigned)

A sample of gas expands from 1.0 m³ to 4.0 m³ while its pressure decreases from 40 Pa to 10 Pa.

How much work is done by the gas if its pressure changes with volume via each of the three paths shown in the p-V diagram in figure?

\[
W_A = 120 \text{ Joule} \\
W_B = 75 \text{ Joules} \\
W_C = 30 \text{ Joules}
\]
Problem 51 (not assigned)

Gas within a closed chamber undergoes the cycle shown in the p-V diagram of Fig. 19-36.

Calculate the net energy added to the system as heat during one complete cycle.

\[ Q = -30 \text{ Joules} \]

(system lost heat)
Processes – return later

\[ dE_{\text{int}} = dQ - dW = dQ - p\,dV \]

1. Adiabatic \( Q=0 \)
2. Constant volume \( W=0 \)
3. Cyclical \( \Delta E_{\text{int}}=0 \)
4. Free expansion \( Q=W=0 \)

- In chapters 20 and 21, we will discuss these more fully
- Also use processes to understand “ideal” and real gases
- And to understand “entropy”
Heat Transfer

3 principal mechanisms

- **Conduction**
  - Heat transfer through material
  - At microscopic level, thermal agitation of molecules causes adjacent molecules to also move more rapidly

- **Convection**
  - Occurs with fluids
  - Has macroscopic cause: hotter fluid has different (typically lower) density and moves up

- **Radiation**
  - NEW: completely different from those above
Heat and Temperature

So Far
- Temperature vs Heat
- Measuring temperature
- Temperature scales and absolute zero
- thermal expansion
- Heat capacity (absorption)
  - Specific heat
- phase transitions
- Heat and Work
- 1st law of thermodynamics
- heat transfer

To Come
- heat transfer
  - conduction
  - convection
  - Radiation
- Processes and Cycles
- Kinetic Theory of Gases
  - Where gas properties come from
  - ...
- Entropy ...