Lec #8: Thermal Energy (Chaps. 3 & 4)

LAST: Mechanical Energy
- Laws of Motion; Forces
- Work, Kinetic Energy, Potential Energy, Power

TODAY: Thermal Energy. I.
- Internal Kinetic and Potential Energy
- Temperature and Heat
- Laws of Thermodynamics
- Specific Heat & Phase Transitions

THURSDAY: Thermal Energy. II.
- Heat Transfer (conduction, convection, radiation)
- Heat Engines & Efficiency

Recap

- Work = Force x Distance x (cos θ)
- Power=d/dt (Work) (instantaneous)
- KE = (1/2) m v^2
  - change in speed -> change in KE
  - note: can change velocity w/out change in KE
- PE = Force x Distance (e.g. mgh for gravity)
- If forces are "conservative":
  - Mechanical KE + Mechanical PE = constant
  - Work = change in Mechanical Energy
- If *not* conservative, where does the energy go?

How Do We Measure Total Energy?

- Total Energy = External (M.E.) + Internal
- Internal *Kinetic*: Thermal
- Internal *Potential*: Chemical; Nuclear
  - molecular bonds
  - atomic bonds
  - nuclear bonds
- We can't measure Total Energy, but we know that it's huge and takes many forms
- We can, however, measure *changes*....

First Law of Thermodynamics

- ΔE = ΔExternal + ΔInternal = Work + Heat
- In practice, "**heat**" usually refers only to a **change** in internal (thermal) energy, not a basic property of a substance.
  - objects don't contain *heat*, but they do contain *energy*
- "Thermal energy" usually refers only to internal *kinetic* energy, though this is only a small fraction of the total internal energy
- To measure thermal energy, we use "temperature"

Temperature

- Temperature **not** measure of **total** internal energy!
- Temperature **is** a measure of **average** kinetic energy of the molecules
- Internal K.E. -> 0 at "absolute zero", increases with temperature (but must use absolute scale)
- When 2 objects are brought into contact
  - if $T_1>T_2$, "thermal energy" transfer (heat) from T_1 to T_2
  - If $T_1 = T_2$, no energy transfer
- "Heat" is the transfer of thermal energy from higher Temp ---> lower Temp

Table 4.1 TEMPERATURES OF SOME COMMON PHENOMENA*

	°C	°F	K
Water, ice point	0	32	273
Water, boiling point	100	212	373
Absolute zero	−273	−460	0
Liquid nitrogen boiling point	−196	−319	77
Liquid helium boiling point	−269	−454	4
Zinc, melting point	420	787	693
Gold, melting point	1063	1945	1336
Solid CO_2 (Dry Ice) sublimation**	−78	−109	195

*At atmospheric pressure
**Process of going from a solid directly to a gas phase

You should know how to convert between F & C.
Absolute scales: Kelvin (°C) and Rankine (°F).

Temperature

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- When 2 objects are brought into contact
  - if $T_1 > T_2$, "thermal energy" transfer (heat) from T_1 to T_2
  - if $T_1 = T_2$, no (net) energy transfer
- "Heat" is the transfer of thermal energy
  from higher Temp ---> lower Temp
- thought experiment: why do some objects feel colder?

Temperature (continued)

- Temperature parameterizes *average* energy
  - And it's something we can measure!
- Total internal energy = average * # of particles
  - energy content depends on <u>mass</u> and <u>temperature</u>
  - change in energy ~ change in temperature
- We can "parameterize" other internal energies (kinetic or potential) with temperatures
  - T_k kinetic temperature of molecules
  - electron temperature, ion temperature, atomic excitation temperature, radiation temperature.....

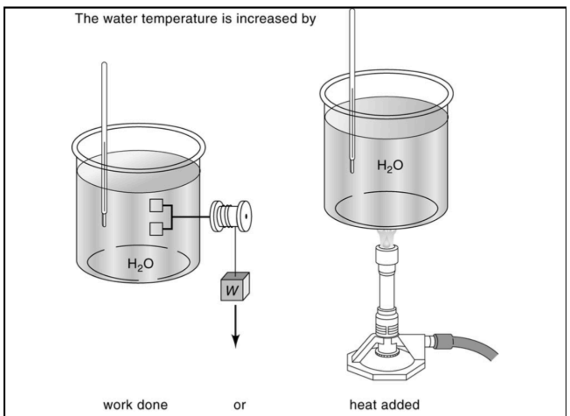
Second Law of Thermodynamics

3 equivalent statements

- an isolated system will steadily change until it reaches "THERMAL EQUILIBRIUM"
  - internal energy shared equally among all forms!
  - everything at same temperature, all temps. same!
- degree of "disorder" always increases until thermal equilibrium
  - state of maximum disorder
  - disorder ("entropy") never decreases on its own
- impossible to transfer heat completely into work:
  $\Delta E = W + Q$ W->Q up to 100% Q--> W < 100%
  - many processes in nature are irreversible
  - even reversible processes aren't 100% efficient

Thermal Energy Units

- 1 calorie = energy required to change temperature of 1 g of water by 1^{o} Celsius
- 1 food Calorie = 1 kilocalorie (1000 calories)
- 1 BTU = energy required to change 1 pound of water by 1^{o} Fahrenheit
- "Mechanical Equivalent of Heat":
  - historical reasons for different units
  - 1 calorie = 4.184 Joules
  - 1 BTU = 1055 J = 252 cal = 778 ft-lb



Specific Heat

- Heat "flows" from higher temp. to lower temp.
  - can either do work (mechanical energy)
  - or change internal energy (change in temperature)
- Change in Temperature for given amount of heat depends on substance
- $Q = c_s \, m \, \Delta T$
  - Q = heat = energy difference
  - c_s is "specific heat"
  - depends on material (why?)
  - also depends on the nature of the process (e.g. constant volume, constant pressure, ...)
  - it even depends on the temperature and pressure

Material	Specific Heat (J/kg/°C)	Specific Heat (Btu/lb/°F)
Water ⟵	4186	1.00
Aluminum	900	0.22
Iron	448	0.12
Copper	387	0.093
Concrete	960	0.23
Glass	840	0.20
White pine	2800	0.67
Ice ⟵	2090	0.50
Air	1004	0.24
Rock	840	0.20

Table 4.2 SPECIFIC HEATS OF COMMON SUBSTANCES

Specific Heat: amount of energy needed to change the temperature of a fixed mass of a substance by a fixed amount

Note that water has a very high specific heat!
Why? How can we take advantage of this?

Phase Transitions

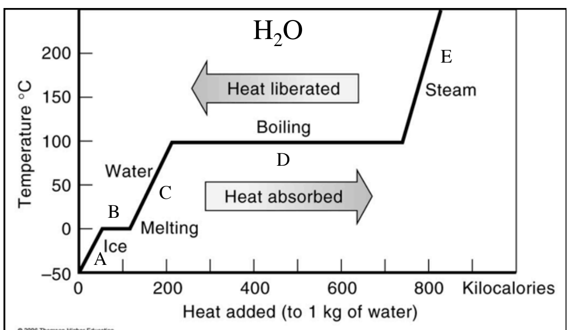
T increasing ------> <----- P increasing

SOLID LIQUID GAS PLASMA

melt-> *vaporize->* *ionize->*

<-fuse *<-condense* *<-recombine*

- generally, $Q=mc\Delta T$
  - c_s = "specific heat" ; m=mass
  - can be different for same material in diferent phases (e.g. H_2O solid 0.5, liquid 1.0, gas 0.48)
- *during* phase transition, $Q=mL$
  - L = "latent heat"
  - no change in temperature! (where does energy go?)



A) c = 2090 J/(kg K) C) c = 4190 J/(kg K) E) c=2010 J/(kg K)

B) L_f = 333000 J/kg D) L_v = 2260000 J/kg