Physics 101: Lecture 26
Kinetic Theory and Heat

What concepts did you find most difficult, or what would you like to be sure we discuss in lecture?

- list of exam topics
  - Exam 3: next W-F, covers Lectures 16 (AM only!) **-21**
  - (angular momentum through oscillations) Sign up!
  - Review session *today* 7pm+ **1310 DCL**
- Finally -- after weeks of trekking through foreign territory with no oasis in sight I have arrived on land I recognize! O fresh water! O wonderful rest! O PV=nRT!
- what's an isotropic solid?
- EXPANSION
- The last question with the holes in the aluminum plates.

Kinetic Theory

The relationship between energy and temperature (for monatomic ideal gas)

$$\Delta p_x = 2mv_x$$

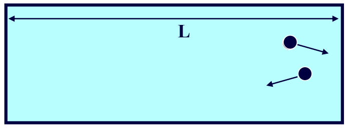
$$\Delta t = 2\frac{L}{v_x}$$

$$F_{avg} = \frac{\Delta p_x}{\Delta t} = \frac{mv_x^2}{L}$$

For N molecules, multiply by N

$$P = \frac{F}{A} \qquad P = \frac{Nmv_x^2}{V}$$

Note $K_{tr} = \frac{1}{2} m v^2 = 3/2\, m v_x^2$



$$\boxed{P = \frac{2}{3}\frac{N}{V}\langle K_{tr}\rangle}$$

⟨⟩ means *average.*

Using PV = NkT

$$\boxed{\langle K_{tr}\rangle = \frac{1}{2}m\langle v^2\rangle = \frac{3}{2}kT}$$

kT/2 energy per degree of freedom = equipartition theorem

Example

- What is the rms (root mean squared) speed of a nitrogen (N_2) molecule in this lecture hall?

$$\langle KE\rangle = \frac{3}{2}k_B T$$

$$\frac{1}{2}m\langle v^2\rangle = \frac{3}{2}k_B T$$

$$\langle v^2\rangle = \frac{3k_B T}{m}$$

v = 511 m/s

= 1143 mph!

(Speed of sound is 767 mph)

$$\langle v^2\rangle = \frac{3\,(1.38\times10^{-23}\,\mathrm{J/K})(273+20)\mathrm{K}}{(28\,\mathrm{u})\times(1.66\times10^{-27}\mathrm{kg/u})}$$

Internal Energy

- Energy of all molecules, including
  - Random motion of individual molecules
    - » $\langle K_{tr} \rangle = (3/2)\ k\ T$ for ideal gas
    - » Vibrational energy of molecules and atoms
  - Chemical energy in bonds and interactions
- DOES NOT INCLUDE
  - Macroscopic motion of object
  - Potential energy due to interactions w/ other objects
  - $E_{tot} = K + U + U_{internal}$

Heat

- Definition: Heat is the flow of energy between two objects due to difference in temperature
  - Changes internal energy
  - Note: similar to WORK
  - Object does not "have" heat (it has energy)
- Units: Joules or *calories*
  - calorie: Amount of heat needed to raise 1g of water 1ºC
  - 1 Calorie = 1000 calories = 4186 Joules

Specific Heat (for solids and liquids)

- Heat adds energy to object/system
- IF system does NO work then:
  - Heat increases internal energy. $Q = \Delta U$
  - Heat increases temperature!
- $Q = c\ m\ \Delta T$
  - Specific heat c, units = J/kg°C
  - Heat required to increase Temp depends on amount of material (m) and type of material (c)

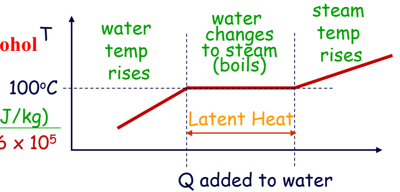
Demo: water in paper cup takes up heat from flame

Latent Heat L

- As you add heat to water, the temperature increases to the boiling point, then it remains constant despite the additional heat!

Demo: evaporative cooling with alcohol

Substance	L_f (J/kg)	L_v (J/kg)
water	33.5×10^4	22.6×10^5



- Latent Heat L [J/kg] is heat which must be added (or removed) for material to *change phase* (liquid-gas).
- Latent Heat of Fusion (L_f) and Latent Heat of Vaporization (L_v)

Example

- How much ice (at 0 C) do you need to add to 0.5 liters of a water at 25 C, to cool it down to 10 C?

(L_f = 80 cal/g, c = 1 cal/g C)

Key ideas: conserve energy by accounting for all the heat flow

1) Q leaving water goes into heating ice.

2) Final temps are same

$Q_{water} = -Q_{ice}$

$$\underbrace{m_{water}c(T_f - T_0)}_{\text{Cool Water}} = \underbrace{-m_{ice}L_f}_{\text{Melt Ice}} - \underbrace{m_{ice}c(T_f - 0)}_{\text{Warm water that was ice}}$$

$$m_{water}c(T_f - T_0) = -m_{ice}(L_f + cT_f)$$

Advice: compute the heat leaving one object (keep it positive) and set it to heat entering the other (keep it positive)

$$m_{ice} = m_{water}c(T_0 - T_f)/(L_f + cT_f)$$

$$= 83.3\ g$$

Demo: Hot brass into water, What's the final Temperature?

Example

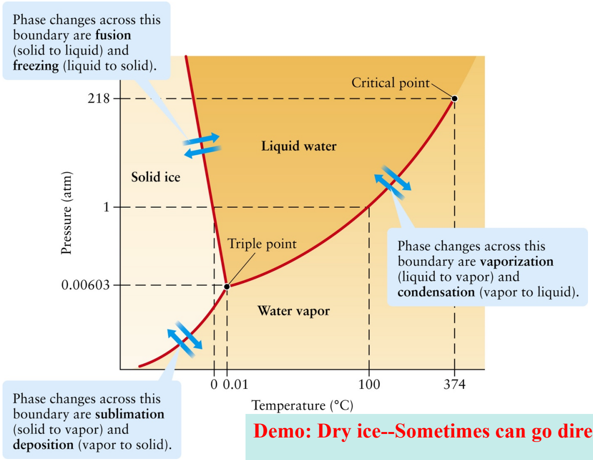
Summers in Phoenix Arizona are very hot (125 F is not uncommon), and very dry. If you hop into an outdoor swimming pool on a summer day in Phoenix, you will probably find that the water is too warm to be very refreshing. However, when you get out of the pool and let the sun dry you off, you find that you are quite cold for a few minutes (yes...you will have goose-bumps on a day when the air temperature is over 120 degrees).

How can you explain this?

When you leave the pool, the water on your skin starts to evaporate. This evaporation takes energy from the surface of your skin, which is why you feel cool when you get out of the pool. This is similar to sweating.

-- explanation from a previous Phys 101 student

Phase Diagrams

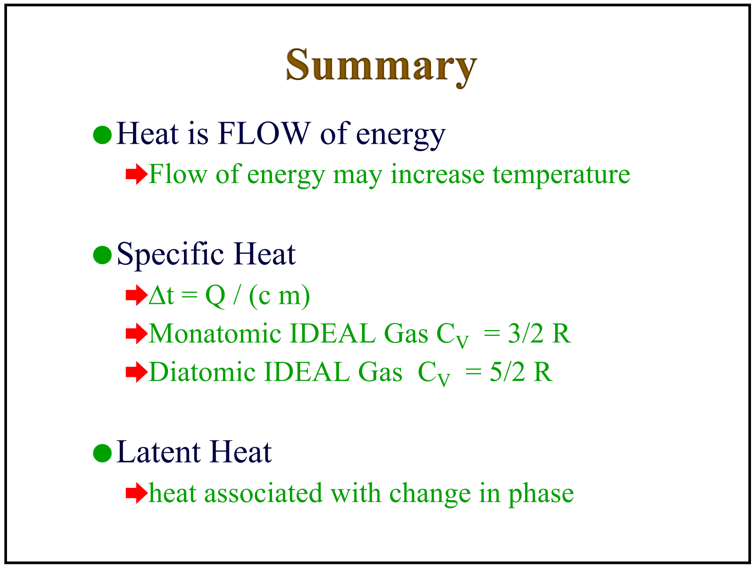


Phase changes across this boundary are **fusion** (solid to liquid) and **freezing** (liquid to solid).

Phase changes across this boundary are **vaporization** (liquid to vapor) and **condensation** (vapor to liquid).

Phase changes across this boundary are **sublimation** (solid to vapor) and **deposition** (vapor to solid).

Demo: Dry ice--Sometimes can go directly from solid to gas, called sublimation



Summary

- Heat is FLOW of energy
  - Flow of energy may increase temperature
- Specific Heat
  - $\Delta t = Q / (c\ m)$
  - Monatomic IDEAL Gas $C_V = 3/2\ R$
  - Diatomic IDEAL Gas $C_V = 5/2\ R$
- Latent Heat
  - heat associated with change in phase