FINAL

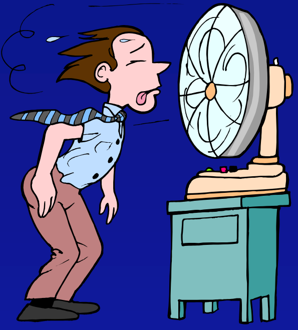
Physics 101: Lecture 25
Heat

Today's lecture will cover Textbook Chapter 14.1-14.5

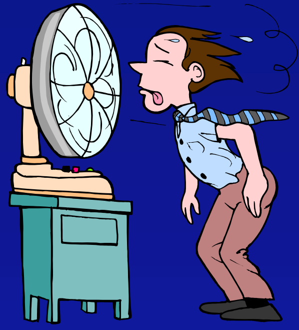
Internal Energy

- 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



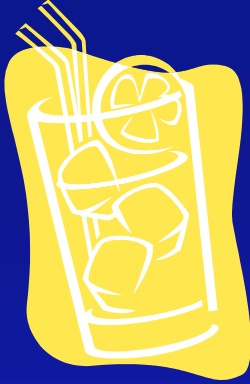
- Definition: Flow of energy between two objects due to difference in temperature
  - ➔ Changes internal energy
  - ➔ Note: similar to WORK
  - ➔ Object does not "have" heat (it has energy)

- Units: Joules or calories
  - ➔ Amount of heat needed to raise 1g of water 1ºC
  - ➔ 1 Calorie = 1000 calories = 4186 Joules

Specific Heat

- 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)
- $\Delta T = Q/cm$

Act



- After a grueling work out, you drink a liter (1kg) of cold water (0 C). How many Calories does it take for your body to raise the water up to body temperature of 37 C?
  (Specific Heat of water is 1 calorie/gram C)

1) 37 2) 370 3) 3,700 4) 37,000

Prelecture 1 & 2

Suppose you have two insulated buckets containing the same amount of water at room temperature. You also happen to have two blocks of metal of the same mass, both at the same temperature, warmer than the water in the buckets. One block is made of aluminum and one is made of copper. You put the aluminum block into one bucket of water, and the copper block into the other. After waiting a while you measure the temperature of the water in both buckets. Which is warmer?

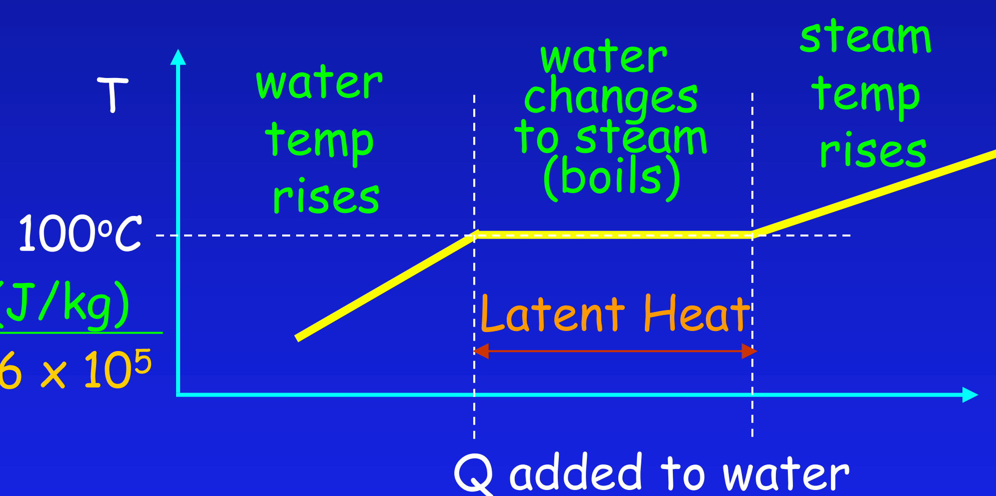
1. The water in the bucket containing the aluminum block
2. The water in the bucket containing the copper block
3. The water in both buckets will be at the same temperature

Substance	c (heat capacity)
	J/(kg-C)
aluminum	900
copper	387

Latent Heat L

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

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)

Ice Act

- Which will do a better job cooling your soda, a "cooler" filled with water at 0C, or a cooler filled with ice at 0 C.

A) Water B) About Same C) Ice

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

Cooling Act

- During a tough work out, your body sweats (and evaporates) 1 liter of water to keep cool (37 C). How much water would you need to drink (at 2C) to achieve the same thermal cooling? (recall c = 4.2 J/g C for water, L_v=2.2x10^3 J/g)

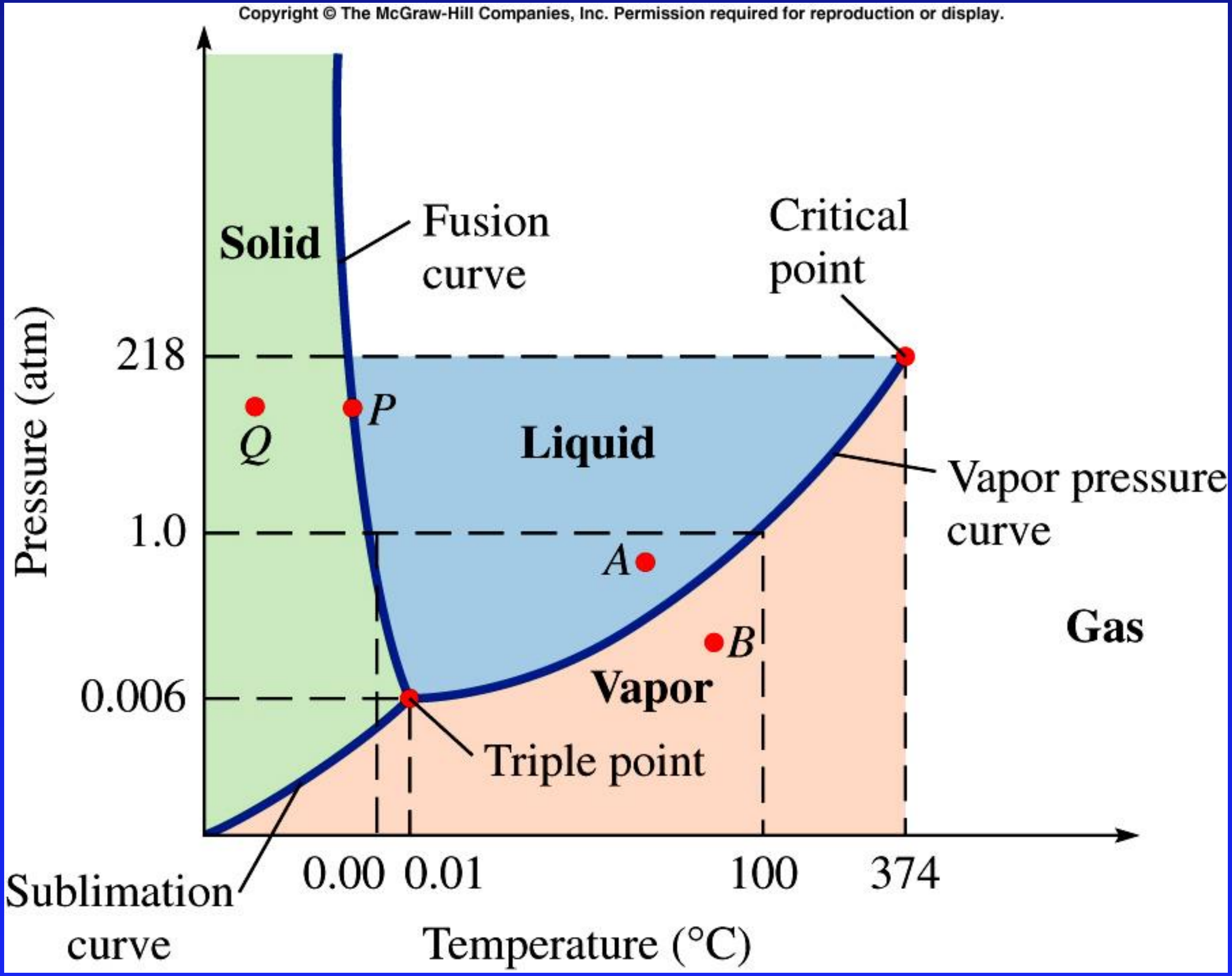
A) 0.15 liters B) 1.0 liters C) 15 liters D) 150 liters

Prelecture 3

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?

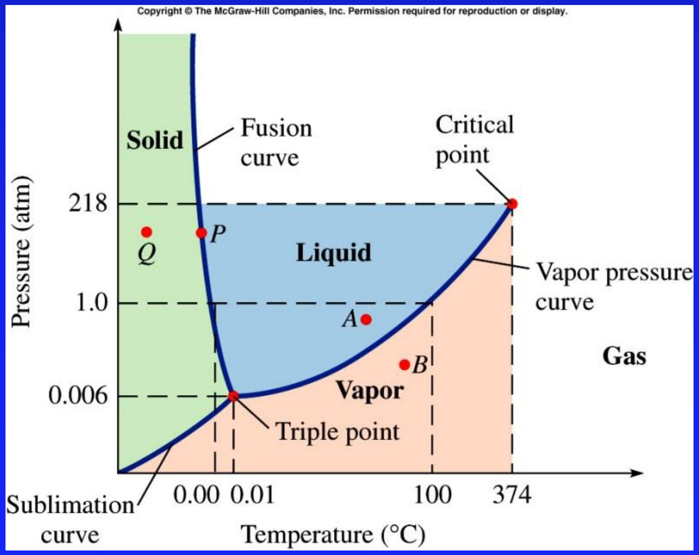
Phase Diagrams



Cooling ACT

- What happens to the pressure in the beaker when placed in ice-water

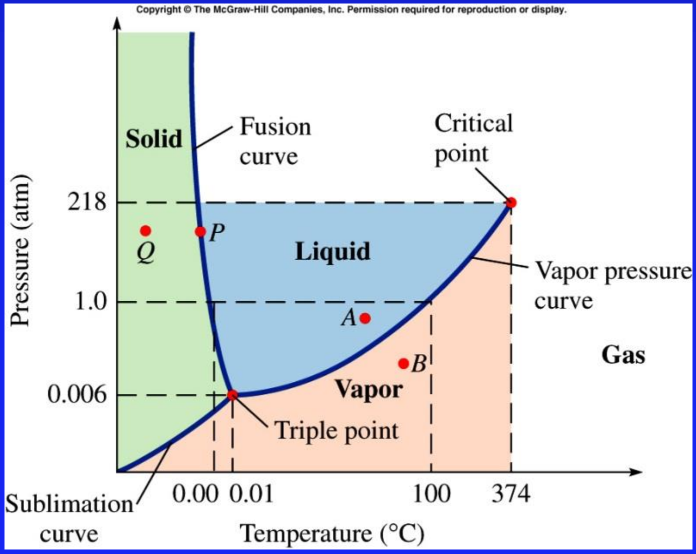
1) Increases 2) Decreases 3) Same



Cooling ACT continued…

- What happens to the boiling point when beaker is placed in ice-water

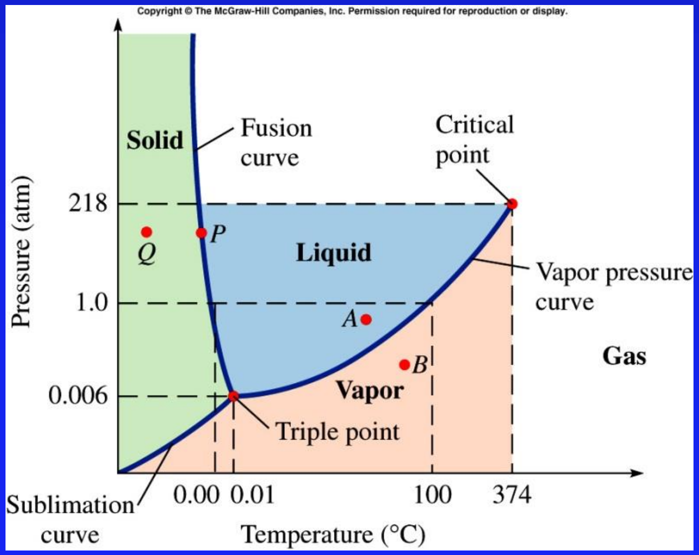
1) Increases   2) Decreases   3) Same



Cooling Act Continued…

- What will happen to the water in the container when I pour ice water over the container

1) cool down 2) Boil 3) Both 4) Neither



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

Key ideas

1) Q leaving water goes into heating ice.

2) Final temps are same

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

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

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

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

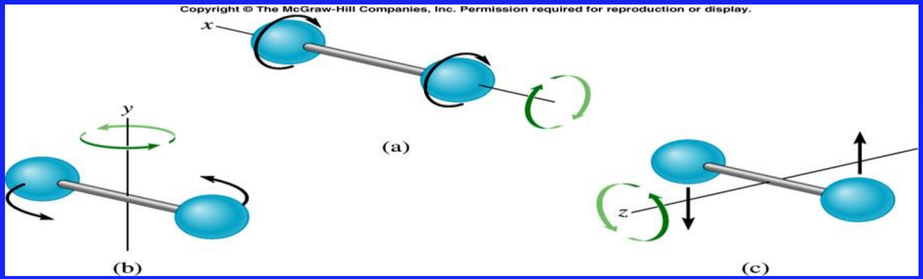
$$= 83.3\ g$$

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

Specific Heat for Ideal Gas

- Monatomic Gas (single atom)
  - ➔All energy is translational Kinetic
  - ➔At constant Volume work = 0
  - ➔$Q = \Delta K_{tr} = 3/2\ nR\Delta T$
  - ➔$C_V = 3/2\ R = 12.5$ J/(K mole)
- Diatomic Gas (two atoms)
  - ➔Can also rotate
  - ➔$C_V = 5/2\ R = 20.8$ J/(K mole)



	Gas	$C_V \left(\frac{J/K}{mol}\right)$
Monatomic	He	12.5
	Ne	12.7
	Ar	12.5
Diatomic	H_2	20.4
	N_2	20.8
	O_2	21.0
Polyatomic	CO_2	28.2
	N_2O	28.4