Lecture 2

Temperature & Heat

- ~~If~~ you take a can of Coke from fridge, & leave it on table, coke's temperature rises until it is in thermal equilibrium w/ room. (same temperature)
- For hot coffee, its temperature will fall if it is left out.

To generalize, we will call the coke or coffee the "system" @ temperature T_S & the surrounding area the "environment" @ temperature T_E.

In general, if $T_S \neq T_E$,

then T_S will change over time

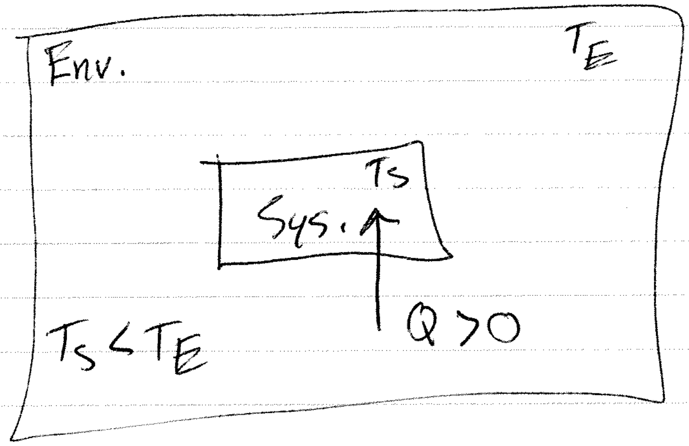
so that $T_S = T_E$ (T_E might change a little also, but we mostly neglect this)

- Change in temperature is due to a ~~change in~~ transfer of heat

(transfer of energy between system & environment)

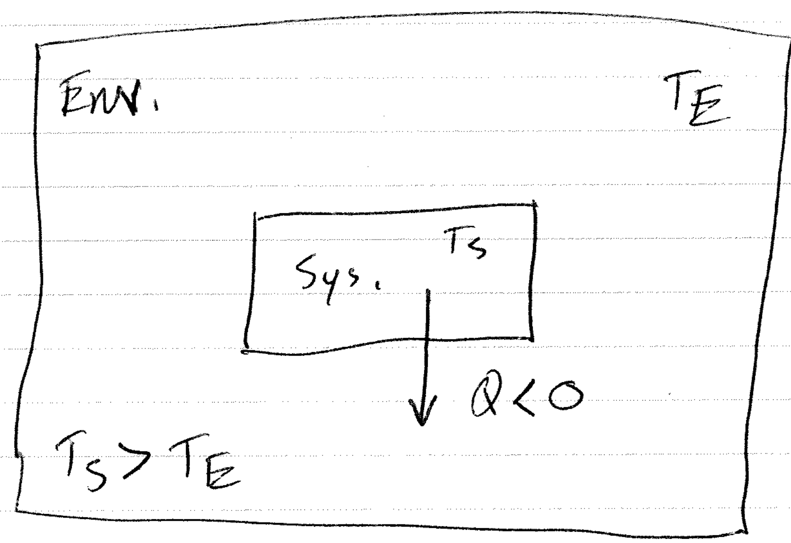
- Heat is denoted by Q (for quantity of heat)

Heat is positive when transferred from environment to system;



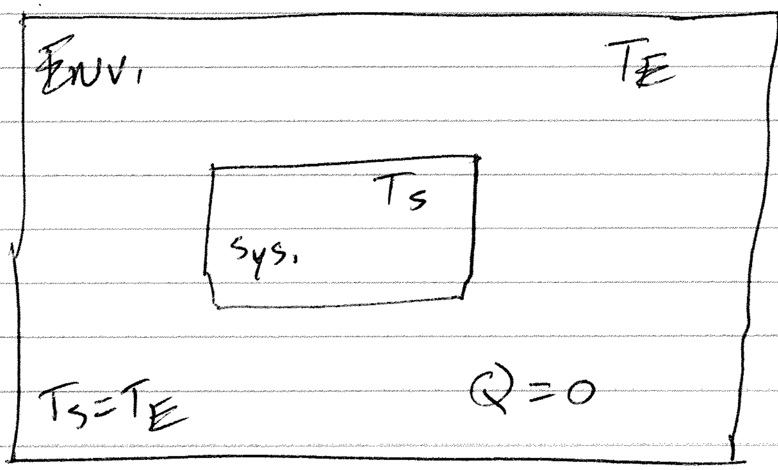
system gains thermal energy from environment when temperature of system is lower.

Heat is negative when transferred from system to environment



system loses thermal energy to environment when sys. temp. is higher.

No heat is transferred when $T_S = T_E$



Definition: Heat is energy transferred between system & environment

Since heat is energy, its SI unit is the joule (denoted by J)

Before scientists realized this, heat was measured in terms of its ability to raise the temperature of water.

- The calorie is the amount of heat energy needed to raise the temperature of 1 gram of water from 14.5 °C to 15.5 °C. 1 calorie = 4.1868 J

- food calorie is actually kilocalorie & measures the amount of energy in the food you consume. This energy can be used for running, lifting weights, thinking, etc.

Absorption of heat by solids & liquids

heat capacity C of an object is constant relating heat Q absorbed by an object & ~~its~~ the difference between its initial & final temperature:

$$Q = C\,\Delta T = C(T_f - T_i)$$

T_i = initial temperature

T_f = final temperature

heat capacity has units of

$\frac{\text{energy}}{\text{Kelvin}}$ or $\frac{\text{energy}}{\text{Celsius}}$

(recall that $T_K = T_C + 273.15$)

Specific Heat

- Two objects made of same material have heat capacities proportional to their masses.
- specific heat c is heat capacity per unit mass & depends on material

Molar specific heat

Sometimes the most convenient unit for indicating amount of substance is the mole (mol)

$$1 \text{ mol} = 6.02 \times 10^{23} \text{ elementary units}$$

Example: 1 mol of aluminum consists of 6.02×10^{23} atoms

1 mol of aluminum oxide consists of 6.02×10^{23} molecules

Molar specific heat C_{molar} gives

$$Q = C_{molar} \cdot n \cdot \Delta T$$

Q: amount of heat energy

C_{molar}: molar specific heat

n: # of moles

ΔT: temp. diff.

Then $C = cm$

& Previous equation becomes

$$Q = cm\,\Delta T = cm\,(T_f - T_i)$$

specific heat, mass, temp. diff.

Example: The Hot Jogger

specific heat of human body $= 3500 \frac{J}{kg \cdot C^\circ}$

$$Q = c\ m\ \Delta T$$

In a half hour, a 65 kg jogger can generate $8 \cdot 10^5$ J.

If the heat were not removed from the jogger's body, by how much will the jogger's temperature rise?

$$Q = cm\Delta T$$

$$\Rightarrow \frac{Q}{cm} = \Delta T$$

$$\frac{8 \cdot 10^5 \text{ J}}{3500 \text{ J}/(\text{kg}\cdot \text{C}^\circ) \cdot 65 \text{ kg}} = 3.5^\circ\text{C}$$

Vaporizing: this means that liquid state changes to vapor or gas state

Heat of transformation: (denoted by L)
Amount of energy per unit mass that must be transferred as heat for a sample to undergo phase change.

$$Q = Lm$$

heat transferred | heat of transformation | mass

When going from liquid to gas, this is called <u>heat of vaporization</u> L_V.

For water, it is

$$L_V = 539 \frac{cal}{g} = 2256 \frac{kJ}{kg}$$

When going from solid to liquid, this is called <u>heat of fusion</u> L_F

For water, it is

$$L_F = 79.5 \frac{cal}{g} = 333 \frac{kJ}{kg}$$

Example Problem

- Piece of copper ~~has~~ has mass 75g & is heated to $T = 312°C$.
- It is then dropped into a glass beaker ~~~~ containing $m_w = 220g$ of water.
- Heat capacity of beaker is $45 \frac{cal}{K}$.
- Initial temp. for water & beaker is 12°C.

$\boxed{\text{What is final temperature } T_f \text{ of whole isolated system @ thermal equilibrium?}}$

System is isolated &
so total energy cannot change
& only internal transfers occur:

$$Q_w + Q_b + Q_c = 0$$

heat transferred to water

heat transferred to beaker

heat transferred from copper

$$Q_w = c_w m_w (T_f - T_i)$$

$$Q_b = C_b (T_f - T_i)$$

$$Q_c = c_c m_c (T_f - T)$$

$$\Rightarrow c_w m_w (T_f - T_i) + C_b (T_f - T_i) + c_c m_c (T_f - T) = 0$$

$$\Rightarrow T_f\left(c_w m_w + C_b + c_c m_c\right)$$

$$= c_w m_w T_i + C_b T_i + c_c m_c T$$

$$\Rightarrow T_f = \frac{c_w m_w T_i + C_b T_i + c_c m_c T}{c_w m_w + C_b + c_c m_c}$$

looking up values for c_w & c_c,

we can solve this