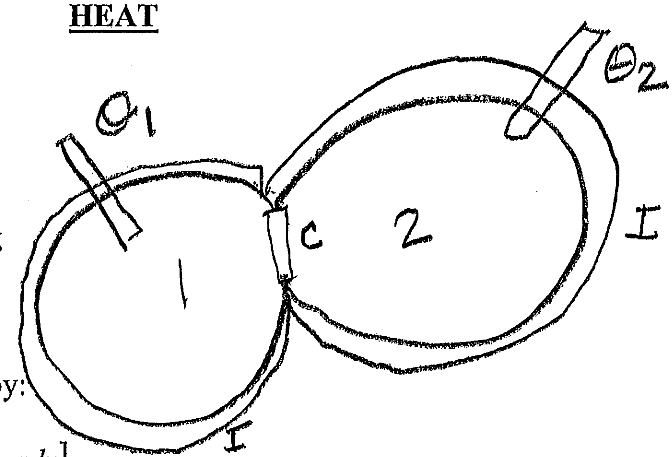
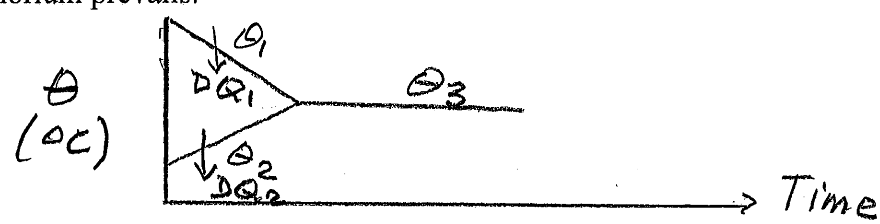
HEAT

Now we know that in order to understand the experiment in which two Thermodynamic Systems are allowed to "talk" to one another through a conducting wall, we must use Temperature as the relevant variable. Once we do that, the two observations described earlier will be described by:



Case I: $\theta_1 = \theta_2$ [*equilibrium already*]

Case II: $\theta_1 = \theta_2$ are different but both change until they become equal and then equilibrium prevails.



In case II we will find that invariably the warmer system (high θ) cools and the colder system warms until the final temperature (θ_3) is reached.

The next questions are why do the two temperatures change and what is exchanged between the systems to cause the changes. This brings out the definition of Heat: If two systems at different temperatures are separated by a conducting wall, energy will "flow" from one to the other. This Exchange of energy is called HEAT.

Definition: HEAT IS A FORM OF ENERGY WHICH IS EXCHANGED BETWEEN SYSTEMS WHEN THEY ARE AT DIFFERENT TEMPERATURES AND NO AGENCY IS BEING USED TO PREVENT THE EXCHANGE.

[Immediate Consequence: It is meaningless to talk about the quantity of Heat within a system.]

We use DQ to indicate that we are talking about Exchange only. Later, we will learn that this exchange depends on how the process is carried out and that is why we use a capital "D".

Of course, energy must be conserved so in the above experiment heat lost by the warmer system must be exactly equal to that gained by the cooler system. That is

$$DQ_1 + DQ_2 = 0$$

No heat is exchanged with the surroundings because the walls are insulators.

UNIT OF HEAT

We use the properties of water to define the unit of heat. In order to change the temperature of one gram of water from 14.5°C to 15.5°C, it will take 1 calorie of heat [Heat M $L^2 T^{-2}$ cal scalar].
A kilocalorie requires 1 kg of water.
Next, for any solid or liquid, it turns out that the quantity of heat-required to change the temperature depends on the mass, hence one can write

$$DQ = m\,C(\theta_f - \theta_i)$$

where C is the specific heat (quantity of heat required to change temperature of a kilogram of material by one degree).
Determination of C gets us into the Science of Calorimetry. A calorimeter is a device whose walls are totally insulating. Our two systems can then be a quantity m_w of water at temperature θ_w and an amount m of material whose specific heat C we want to measure. We heat it to a temperature θ, drop it into water, close the calorimeter and wait for equilibrium. Then

$$mC(\theta_f - \theta_i) + m_w C_w(\theta_f - \theta_w) = 0$$

and calculate C. For example, Lead has C = 0.0305 cal/g °C.
However, transference of heat can have another effect. The temperature does not change but the solid turns into a liquid (or vice verse) or a liquid turns into vapor. That brings into play Latent Heat

$$D\,Q = m\,L$$

L: quantity of heat required to change the state [Sol$\rightarrow$ Liq, Liq$\rightarrow$vapor] without changing the temperature.
Examples: It takes 80 calories/g to change ice into water at 0°C and nearly 540 calories to change 1g of water into 1g of steam at 100°C.

Mechanical Equivalent

When you rub your hands together, they get warmer but NO HEAT IS INVOLVED. What you have discovered is that a certain amount of mechanical (frictional) work will produce the same effects as heat. Indeed, we will do an experiment in which you warm water by rubbing a copper can with a rough loaded string and learn that 4.186 Joules of mechanical work will MIMIC the effects of 1 calory of heat, BUT IT IS NOT HEAT!