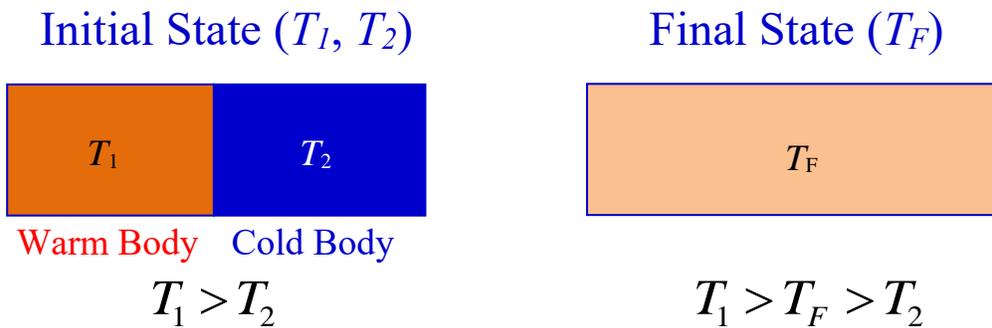


## Lecture 6 Heat

(Ch. 3 First Law of Thermodynamics  
(3.2a Heat))

Consider the following system:



The energy which exchanged from the warm body to the cold body is called **heat**.

The amount of **heat loss by the warm body** is proportional to the temperature difference between the initial and the final states of that body:

$$-\Delta Q = C_1(T_F - T_1) \quad (3.5)$$

Similarly, the amount of **heat gained by the cold body** is proportional to  $T_F - T_2$ ,

$$\Delta Q = C_2(T_F - T_2) \quad (3.6)$$

where  $C_1$  and  $C_2$  are called **heat capacities** of the two these bodies.

The amount of **heat gained by the cold body** must be the same as the amount of heat lost by the warm body. Adding (3.5) and (3.6) gives

$$C_1(T_F - T_1) + C_2(T_F - T_2) = 0 \quad (3.7)$$

From Eq. (3.5), we have

$$C_1 = \frac{-\Delta Q}{(T_F - T_1)} = \frac{-\Delta Q}{-(T_1 - T_F)} = \frac{\Delta Q}{\Delta T_1}$$

where  $\Delta T_1 = T_1 - T_F$ . Thus,  $C_1$  is the heat loss by body 1 (warm body) lost per degree K.

Similarly,  $C_2$  can be derived from Eq. (3.6),

$$C_2 = \frac{\Delta Q}{(T_F - T_2)} = \frac{\Delta Q}{\Delta T_2}$$

where  $\Delta T_2 = T_F - T_2$ . Thus,  $C_2$  is the heat gained by body 2 (cold body) gained per degree K.

Clearly, as  $\Delta T_1$  and  $\Delta T_2$  approach 0,  $T_1$ ,  $T_2$  and  $T_F$  approach to the same value. In other words, we have

$$C_1 = \lim_{\Delta T_1 \rightarrow 0} \frac{\Delta Q}{\Delta T_1} = \lim_{\Delta T_2 \rightarrow 0} \frac{\Delta Q}{\Delta T_2} = C_2 = \left. \frac{dQ}{dT} \right|_{T=T_F}$$

Thus, we can define the **heat capacity** at temperature  $T$  as,

$$C = \frac{dQ}{dT} \quad (3.8)$$

Divide (3.8) by mass  $m$ , we get the expression for **specific heat capacity**,

$$c = \frac{dq}{dT} \quad (3.9)$$

where  $q=Q/m$ ,  $c=C/m$ . The **unit for the specific heat is  $J K^{-1} kg^{-1}$** .

If the transformation process from  $T_1$  to  $T_F$  is **isobaric**, the specific heat is called the **specific heat at constant pressure**:

$$c_p = \left( \frac{dq}{dT} \right) \Big|_p.$$

Similarly, if the transformation process is **isosteric**, the specific heat is called the **specific heat at constant volume**:

$$c_v = \left( \frac{dq}{dT} \right) \Big|_v.$$

For example,

For water vapor,  $c_v = 1463 J kg^{-1} K^{-1}$ ,  $c_p = 1952 J kg^{-1} K^{-1}$ .

For dry air,  $c_v = 717 J kg^{-1} K^{-1}$ ,  $c_p = 1004 J kg^{-1} K^{-1}$ .

Note that **in general  $c_p > c_v$** .

This may be explained by considering the piston-cylinder system. The amount of heat needed to heat the air in the cylinder with a constant pressure ( $c_p$ ) is larger than that with a constant pressure ( $c_v$ ) because part of the heat is used to do work of expansion on the environment.

Note also that  $c_p$  and  $c_v$  are temperature dependent. Strictly speaking, an exact value of  $c_p$  or  $c_v$  only applies to the condition of a certain temperature. However, for gas (e.g., air and water vapor), the difference for different temperatures is small. Thus, normally we just use the constant values as shown above.

Note that in SI or MKS units, heat has a unit of *Joule* ( $N\ m$  or  $kg\ m^2\ s^{-2}$ ) same as work and energy. Sometimes heat is measured by *calorie*, which is defined as the quantity of heat needed to raise the temperature of 1 gram of pure water from 14.5 °C to 15.5 °C. The relation between  $J$  and calorie is 1 calorie = 4.186  $J$ .