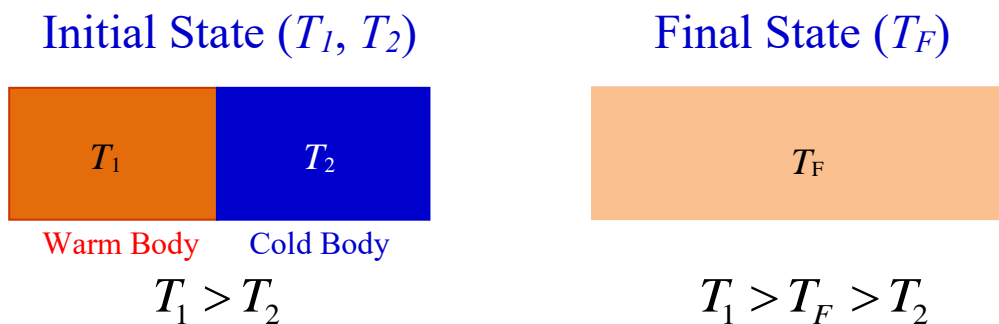


## Lecture 6 Heat

(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:

(a) 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.

**(b)  $c_p$  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.

**(c) Heat has a unit of *Joule* ( $N m$  or  $kg m^2 s^{-2}$ )**

It is the 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  $J$ .