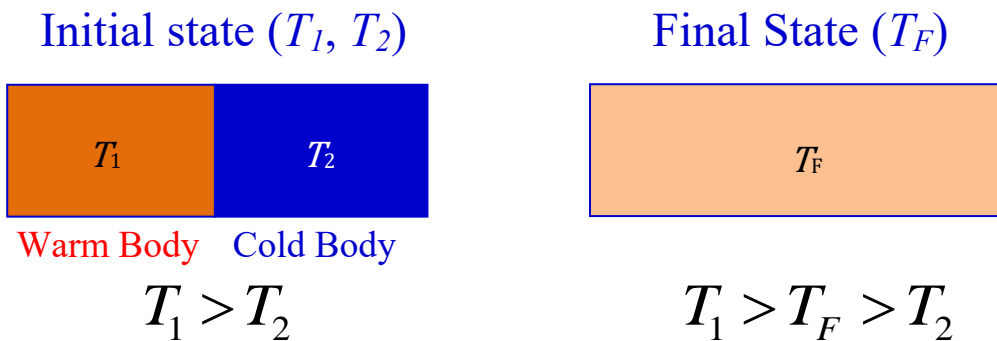


Lecture 6 Heat

(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.

From Eq. (3.6) we have

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

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 ΔT_1 and ΔT_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,

$$C = \frac{dQ}{dT} \tag{3.8}$$

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

$$c = \frac{dq}{dT} \tag{3.9}$$

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

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, $c_v=1463 \text{ J kg}^{-1} \text{ K}^{-1}$, $c_p=1952 \text{ J kg}^{-1} \text{ K}^{-1}$.

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

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 (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 $1\ \text{calorie} = 4.186\ J$.