



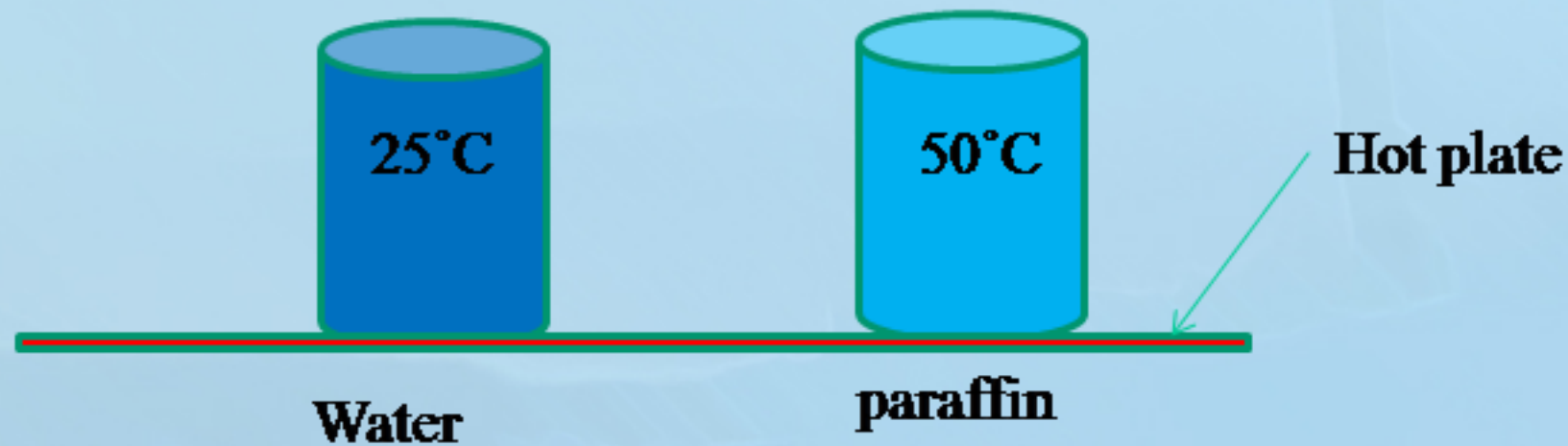
✓ Enthalpy

✓ Heat Capacity

The First Law of Thermodynamics

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Consider 2 beaker's filled with paraffin and water



If both are heated on the same heat source, over the same amount of time, the temperature rise of the paraffin will be twice as high.

The paraffin heats up twice as much given the same amount of energy as the water does. We say that the specific heat capacity or the amount of energy it takes to raise the temperature of the water is higher than the specific heat capacity of the paraffin.

Specific Heats

Specific heat is the energy required to raise the temperature of the unit mass of a substance by one degree.

The specific heat capacity (c) of a substance is the amount of energy needed to raise the temperature of 1kg of the substance by 1K (or 1 °C)

The specific heat capacity $c \equiv \frac{C}{m}$

Specific heat capacity is usually quoted as the energy(j) absorbed to raise the Temperature of a mass of 1 Kg by 1 K or 1 Degree Celsius. It's units are $\text{J Kg}^{-1} \text{K}^{-1}$ or C^{-1} and it's symbol is (c).

E.g

If a substance has a specific heat capacity of $1000\text{J Kg}^{-1}\text{C}^{-1}$

Then applying 1000J of heat energy will cause 1 kg of that substance to increase its temperature by 1 degree.

Applying 2000J increases the temperature by 2C .

It could also be said that applying 1000J will cause the temperature of 0.5 kg of that substance to rise 2C .

WHERE'S THE MATH

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

Q = change in thermal energy

m = mass of substance

ΔT = change in temperature ($T_f - T_i$)

c = specific heat of substance

Practice:

1. How much heat is given out when an iron ball of mass 2 Kg and specific heat capacity $440 \text{ Jkg}^{-1}\text{C}^{-1}$ cools from 300 to 200C?

88.000 J

2. A piece of aluminium of mass 0.5Kg is heated to 100C and then placed in 0.4kg of water at 10C. If the resulting temperature of the mixture is 30C, what is the specific heat capacity of aluminium if that of water is $4,200\text{J/kgC}^{\circ}$

Self learning :

- Problems of enthalpy (1)
- The relation between C_p and C_v

Next time :

The Second Law of Thermodynamics

Enthalpy

Enthalpy is the **heat content** of a system, or the **amount of energy within a substance**, both kinetic and potential.

Every substance possesses both stored energy, due to the nature of the chemical (and nuclear) bonds holding the substance together, and kinetic energy which arises from the constant motion of the particles. This total amount of energy is **enthalpy**.

The enthalpy symbol with a degree symbol, H° , indicate that enthalpy was measured under standard conditions.

However, it is not possible to actually measure the heat content of a substance that is just "sitting there" – so we can only measure how much enthalpy *changes* during a process (ΔH).

What would happen if we created a set of conditions under which no work is done by the system on its surroundings, or vice versa, during a chemical reaction?

Under these conditions, the heat given off or absorbed by the reaction would be equal to the change in the internal energy of the system.

$$\Delta U = Q \text{ (if only if } W = 0)$$

The easiest way to achieve these conditions is to run the reaction at constant volume, where no work of expansion is possible. At constant volume, the heat given off or absorbed by the reaction is equal to the change in the internal energy that occurs during the reaction.

$$\Delta U = Q_v \text{ (at constant volume)}$$

Most processes or reactions, however, are run in open flasks and beakers. When this is done, the volume of the system is not constant because gas can either enter or leave the container during the reaction. The system is at constant pressure, however, because the total pressure inside the container is always equal to atmospheric pressure.

If a gas is driven out of the flask during the reaction, the system does work on its surroundings. If the reaction pulls a gas into the flask, the surroundings do work on the system. We can still measure the amount of heat given off or absorbed during the reaction, but it is no longer equal to the change in the internal energy of the system, because some of the heat has been converted into work.

$$Q = \Delta U + W \quad \rightarrow \quad \Delta U = Q - W$$

Enthalpy (H), is the sum of the internal energy of the system plus the product of the pressure of the gas in the system times the volume of the system.

$$H = U + PV$$

The change in the enthalpy of the system during a process is equal to the change in its internal energy plus the change in the product of the pressure times the volume of the system.

$$\Delta H = \Delta U + \Delta PV$$

Contoh:

Suatu reaksi berlangsung pada volume tetap disertai penyerapan kalor sebanyak 200 kJ. Tentukan nilai ΔU , ΔH , Q dan W reaksi itu.

Jawab :

Sistem menyerap kalor sebanyak 200 kJ , berarti :

$$Q = + 200 \text{ kJ}$$

Reaksi berlangsung pada volume tetap , maka

$$W = 0 \text{ kJ.}$$

$$\Delta U = Q + W = + 200 \text{ kJ} + 0 \text{ kJ} = 200 \text{ kJ}$$

$$\Delta H = \Delta U + P\Delta V = + 200 \text{ kJ} + 0 = 200 \text{ kJ}$$

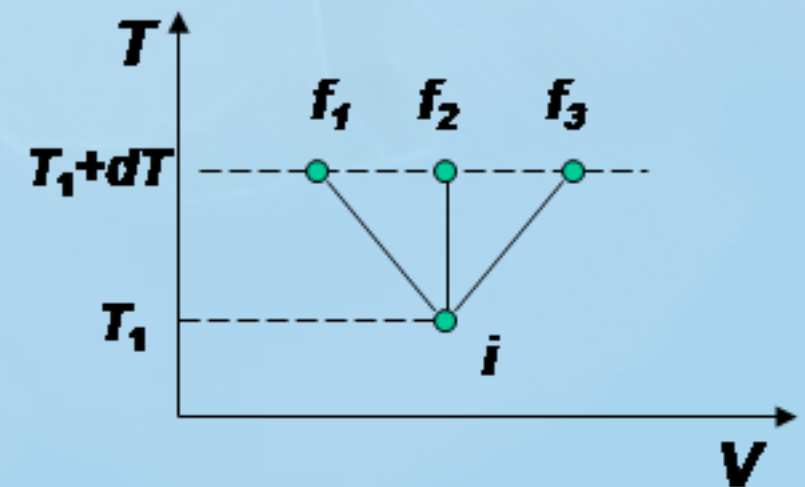
$$\Rightarrow \Delta H = \Delta U = Q$$

Heat Capacity

The heat capacity of a system - the amount of energy transfer due to heating required to produce a unit temperature rise in that system

$$C \equiv \frac{\delta Q}{\Delta T}$$

C is NOT a state function (since **Q** is not a state function) – it depends on the path between two states of a system



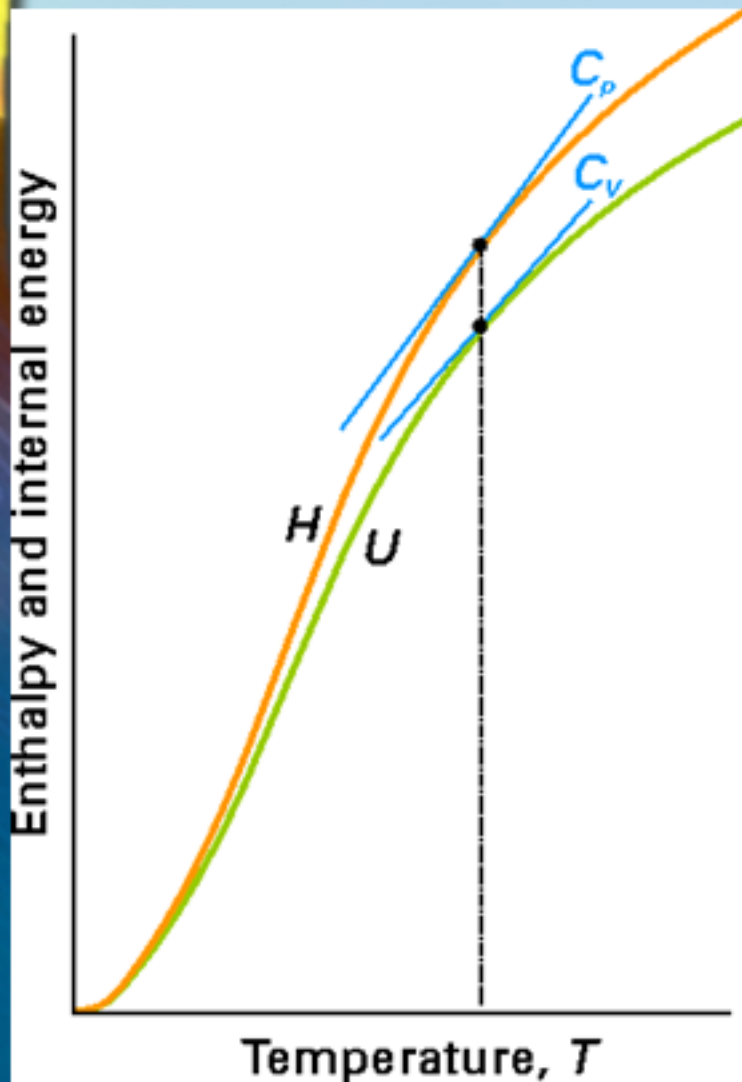
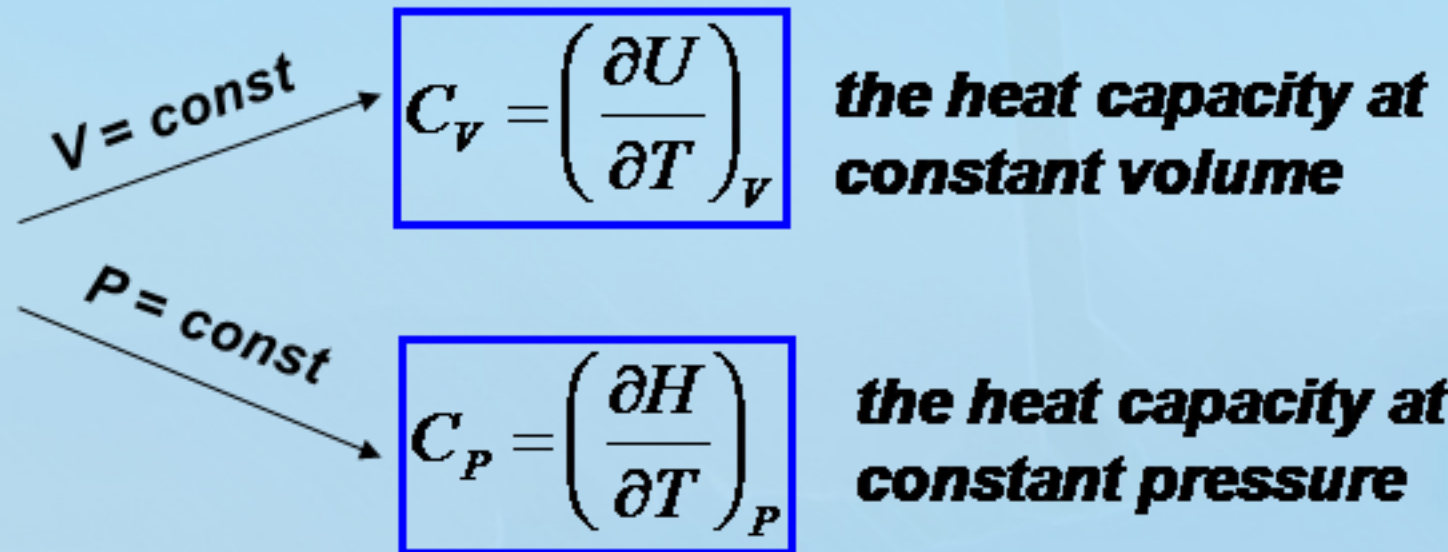
isothermic $C = \infty$

adiabatic $C = 0$

The specific heat capacity $c \equiv \frac{C}{m}$

C_V and C_P

$$C = \frac{\delta Q}{dT} = \frac{dU + PdV}{dT}$$



To find C_P and C_V , we need $f(P, V, T) = 0$ and $U = U(V, T)$

For an ideal gas $U = \frac{f}{2} Nk_B T$ $H = \left(\frac{f}{2} + 1 \right) Nk_B T$

$$C_V = \frac{f}{2} Nk_B = \frac{f}{2} nR$$

$$C_P = \left(\frac{f}{2} + 1 \right) nR$$

of moles

For one mole of a monatomic ideal gas:

$$C_V = \frac{3}{2} R \quad C_P = \frac{5}{2} R$$