

Lecture 34

Heat and phase changes.

Heat

We saw that energy is transferred due to a temperature difference. There is no work involved here. This transferred energy is called **heat**.

Units:

SI: J (Joules)

cal (calorie)

1 cal = heat required to raise the temperature of 1 g of water from 14.5°C to 15.5°C

Specific heat

How much heat is needed to change by ΔT the temperature of a mass m of material X?

$$Q = mc_X \Delta T \quad c_X = \text{specific heat}$$

c_X does have some temperature dependence, but very small (ie, negligible for this course)

Definition of calorie!

water: $c = 1 \text{ cal}/(\text{g } ^\circ\text{C}) = 4186 \text{ J}/(\text{kg K})$ ← Water has a very high specific heat.

iron: $c = 470 \text{ J}/(\text{kg K})$

It's "hard" to increase the temperature of water.

Example: Kettle

Your electric kettle is labeled 2000 W. Estimate how long it will take to boil enough water to fill your 0.5 liter thermos if water comes out of the tap at 15°C.

Heat needed to warm up the water: $Q_{\text{water}} = m_{\text{water}} c_{\text{water}} \Delta T$

$$Q_{\text{water}} = \left(0.5 \text{ liters} \frac{1 \text{ kg water}}{1 \text{ liter water}} \right) \left(4186 \frac{\text{J}}{\text{kg } ^\circ\text{C}} \right) (100 - 15) ^\circ\text{C} = 1.78 \times 10^5 \text{ J}$$

Assuming that all the heat produced by the kettle is used to warm up the water,

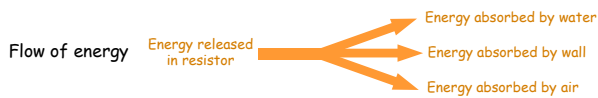
$$t = \frac{Q}{P} = \frac{1.78 \times 10^5 \text{ J}}{2000 \text{ W}} = 89 \text{ s} = 1.5 \text{ min}$$

But of course in reality it will be a little longer. What are we neglecting?

Heat produced by the internal resistance of the kettle also warms up:

- air (negligible if kettle has a lid and is well insulated)
- kettle (at least internal wall)

$$Q_{\text{water}} + Q_{\text{wall}} + Q_{\text{air}} = P_{\text{produced}} t$$



What does specific heat depend on?

Temperature = average kinetic energy of particles

Degrees of freedom (= ways to move, ie, to increase kinetic energy):

Example:

A molecule of helium is made of one atom. It can basically just "bounce around" in 3 directions (3 degrees of freedom)

A molecule of hydrogen is made of two atoms. It can bounce around (3 dof) and it can also rotate (+ 2 dof, total 5 degrees of freedom)

Molar mass

Heavier molecules store require more additional energy to increase their average speed.

Molar heat capacity

How much heat is needed to change by ΔT the temperature of n moles of a certain substance?

$$m = nM \quad \begin{array}{l} n = \text{number of moles} \\ M = \text{mass of one mole (molar mass)} \end{array}$$

$$Q = mc\Delta T = nMc\Delta T \quad C = Mc$$

$$Q = nC\Delta T \quad C = \text{molar heat capacity}$$

Phases or states of matter

Three basic states of matter:

- Solid
- Liquid
- Gas

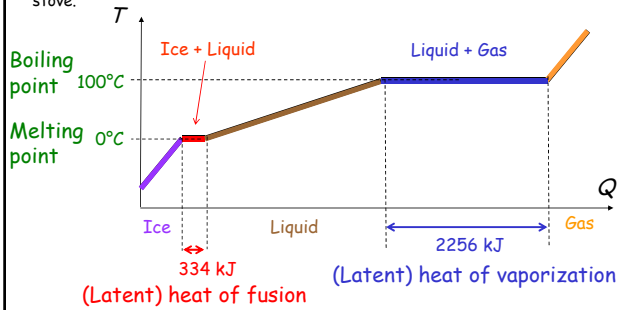
A **phase change** involves a **critical change in the microscopic structure of matter**.

Example:

Ice to water: Lattice disappears, molecules are free to move around.

Analysis: Phase changes for water

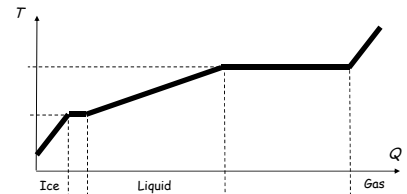
1 kg of ice is placed on a pan on the stove. Plot temperature and heat supplied by stove.



ACT: Specific heats

Which water phase has the largest specific heat?

- A. Ice
- B. Liquid water**
- C. Steam



$$Q = mc\Delta T \Leftrightarrow \text{slope } \frac{dT}{dQ} = \frac{1}{mc} \quad \text{Small slope, large } c$$

Indeed: $c_{\text{ice}} = 2100 \frac{\text{J}}{\text{kg K}}$ $c_{\text{water}} = 4186 \frac{\text{J}}{\text{kg K}}$

Latent heat

This is the heat for 1 kg...

Water: $L_{\text{fusion}} = 334 \text{ kJ/kg}$ $L_{\text{vaporization}} = 2256 \text{ kJ/kg}$

This energy is not used to increase the kinetic energy of the particles (does not increase the temperature) but to **change the structure of matter**.

During a phase change, two or more phases coexist in dynamic equilibrium.

- Examples:
- Vapor and liquid water exactly at 100°C
 - Ice and liquid water exactly at 0°C
 - Vapor/Liquid water/Ice at the triple point (273.16 K and 610 Pa)

In-class example: Ice melting

How much heat is needed to turn 10 g of ice at -5°C into liquid water at 20°C?

A. 105 J $Q_{\text{total}} = Q_{\text{warm ice to } 0^\circ\text{C}} + Q_{\text{melt ice}} + Q_{\text{warm water to } 20^\circ\text{C}}$
 B. 420 J $Q_{\text{warm ice to } 0^\circ\text{C}} = (0.010 \text{ kg}) \left(2100 \frac{\text{J}}{\text{kg}^\circ\text{C}} \right) (0 - (-5^\circ\text{C})) = 105 \text{ J}$
 C. 837 J $Q_{\text{melt ice}} = (0.010 \text{ kg}) \left(333 \frac{\text{kJ}}{\text{kg}} \right) = 3330 \text{ J}$
D. 3330 J $Q_{\text{warm water to } 20^\circ\text{C}} = (0.010 \text{ kg}) \left(4186 \frac{\text{J}}{\text{kg}^\circ\text{C}} \right) (20^\circ\text{C} - 0) = 837 \text{ J}$
E. 4272 J $Q_{\text{total}} = 4272 \text{ J}$
 (mostly from melting)

Follow-up example: Iced coffee

10 g of ice at -5°C are added to 30 ml of hot coffee inside a thermos that is then tightly closed. After the system reaches equilibrium, the temperature of the mix is 20°C . What was the initial temperature of the coffee?

Because the system is thermally isolated (closed thermos), the hot coffee is the only source of energy, so it must provide the necessary 4272 J of heat.

$$Q_{\text{cool coffee}} = m_{\text{coffee}} c_{\text{water}} (20^{\circ}\text{C} - T_{\text{initial}}) = -4272 \text{ J}$$

For the coffee, this is a decrease in energy

$$T_{\text{initial}} = 20^{\circ}\text{C} - \frac{Q_{\text{cool coffee}}}{m_{\text{coffee}} c_{\text{water}}}$$

$$= 20^{\circ}\text{C} - \frac{-4272 \text{ J}}{(30 \times 10^{-3} \text{ kg}) \left(4186 \frac{\text{J}}{\text{kg}^{\circ}\text{C}} \right)} = 54^{\circ}\text{C}$$

10 g of ice at -5°C are added to 30 ml of hot coffee inside a thermos that is then tightly closed. After the system reaches equilibrium, the temperature of the mix is 20°C . What was the initial temperature of the coffee?

Energy balance:

$$Q_{\text{warm ice to } 0^{\circ}\text{C}} + Q_{\text{melt ice}} + Q_{\text{warm water to } 20^{\circ}\text{C}} + Q_{\text{cool coffee}} = 0$$

+ + + -

Energy is absorbed Energy is released

Flow of energy: from hot object to cold object

