# Energy and heat

In Phys 101 you discussed and practiced the conservation of mechanical energy,

i.e. the conversion between kinetic and potential mechanical energy.

Now we are going to expand this discussion and consider the possibility of converting mechanical energy into thermal energy.

For example, friction force converts mechanical (kinetic) energy of, say, moving block, into internal energy of the block and floor, as it slides along → their temperature increases.

In thermodynamics the temperature of an object characterizes its internal energy.

Interestingly, on a microscopic level it is still mechanical energy → kinetic (and maybe potential) energy of individually moving molecules.

~~Macroscopic~~

A block slides on a floor, decelerating due to a friction force → their temperature increases

Macroscopic

Kinetic energy of the block is lost and transferred into internal energy of the system

Microscopic

The organised coherent motion of all molecules is transformed into chaotic (extra) motion of each molecule. (system becomes more disordered)

Formal definition of the internal energy:

total energy of the system minus the Kinetic and potential energies of its motion as a whole.

or

The sum of the mechanical energy (kinetic + potential) of the particles that form the system

Ideal (non-interacting) gas → only kinetic energy (monoatomic)

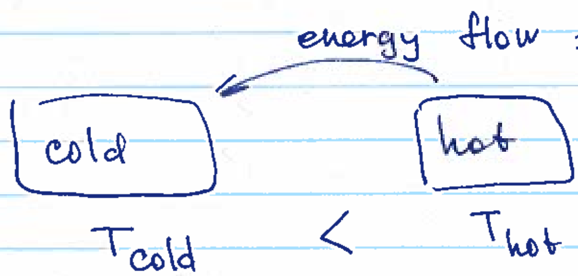
For each particle $\langle K \rangle = \frac{3}{2} k_B \cdot T$

$$\underline{E_{int}} = \underline{N \cdot \frac{3}{2} k_B \cdot T} = \underline{\frac{3}{2} n \cdot RT} \qquad (N \cdot k_B = n \cdot R)$$

Solids & liquids → particle interact strongly, thus we must include the potential energy of these interactions → difficult task! $\underline{\underline{\alpha T}}$

One more important term : <u>Heat</u>

Heat – the form of energy crossing the boundary of a thermodynamic system by virtue of a temperature difference across the boundary



Again, for complex systems (liquids, solids) we empirically find the relations b/w heat (amount of energy absorbed or lost by a system) and the change in its temperature

$$\underset{\text{heat}}{Q} = \overbrace{m}^{\text{mass of the object}} \cdot c \cdot \underbrace{\Delta T}_{\text{change in temperature}}$$

c – specific heat capacity

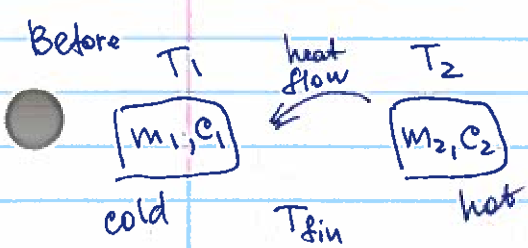
$Q>0 \rightarrow$ heat added to a system $\rightarrow \Delta T>0$
the temperature of the system increases

$Q<0 \rightarrow$ heat is lost $\rightarrow \Delta T<0 \rightarrow$ temperature decreased

$Q=0 \rightarrow$ no heat added or remove
If a system is thermoisolated, $Q=0$

What happens if two objects of different temperatute are brought together, in an isolated environment?
Their total energy stays constant, all the heat lost by the hotter object will be reabsorbed by the colder object $\longrightarrow$ until they are at the same temperature, and the heat flow stops.

General setup: Initially $m_1, T_1, c_1$ (given by mater
$m_2, T_2, c_2$




Final ( $T_{fin}$

$\Delta Q_{hot} = m_2 c_2 (T_{fin} - T_2) = - m_2 c_2 (T_2 - T_{fin})$

$\Delta Q_{cold} = m_1 c_1 (T_1 - T_{fin}) \qquad 0 = \Delta Q_{hot} + \Delta Q_{cold}$

Example 1: unknown heat capacity

Small test object of a known mass $m_{test}$.
dropped into a thermally isolated cup of water


Initial: $m_{test} = 0.3\,kg$ $\quad m_{water} = 1\,kg$

$T_{test} = 100^\circ C$ $\quad T_{water} = 25^\circ C$

$c_{water} = 4186 \frac{J}{kg \cdot ^\circ C}$

Final temperature is measured

$T_{fin} = 35^\circ C$

Heat balance $\quad Q_{water} = m_w \cdot c_w \,(T_{fin} - T_w) > 0$

$Q_{test} = m_{test}\, c_{test}\, (T_{fin} - T_{test}) < 0$

$Q_{water} + Q_{test} = 0 \qquad m_w c_w (T_{fin} - T_w) + m_{test}\, \underline{\underline{c_{test}}}\, (T_{fin} - T_{test}) =$

$$m_w c_w (T_{fin} - T_w) = m_{test}\, c_{test}\, (T_{test} - T_{fin})$$

$$c_{test} = \frac{m_w c_w (T_{fin} - T_w)}{m_{test} (T_{test} - T_{fin})} =$$

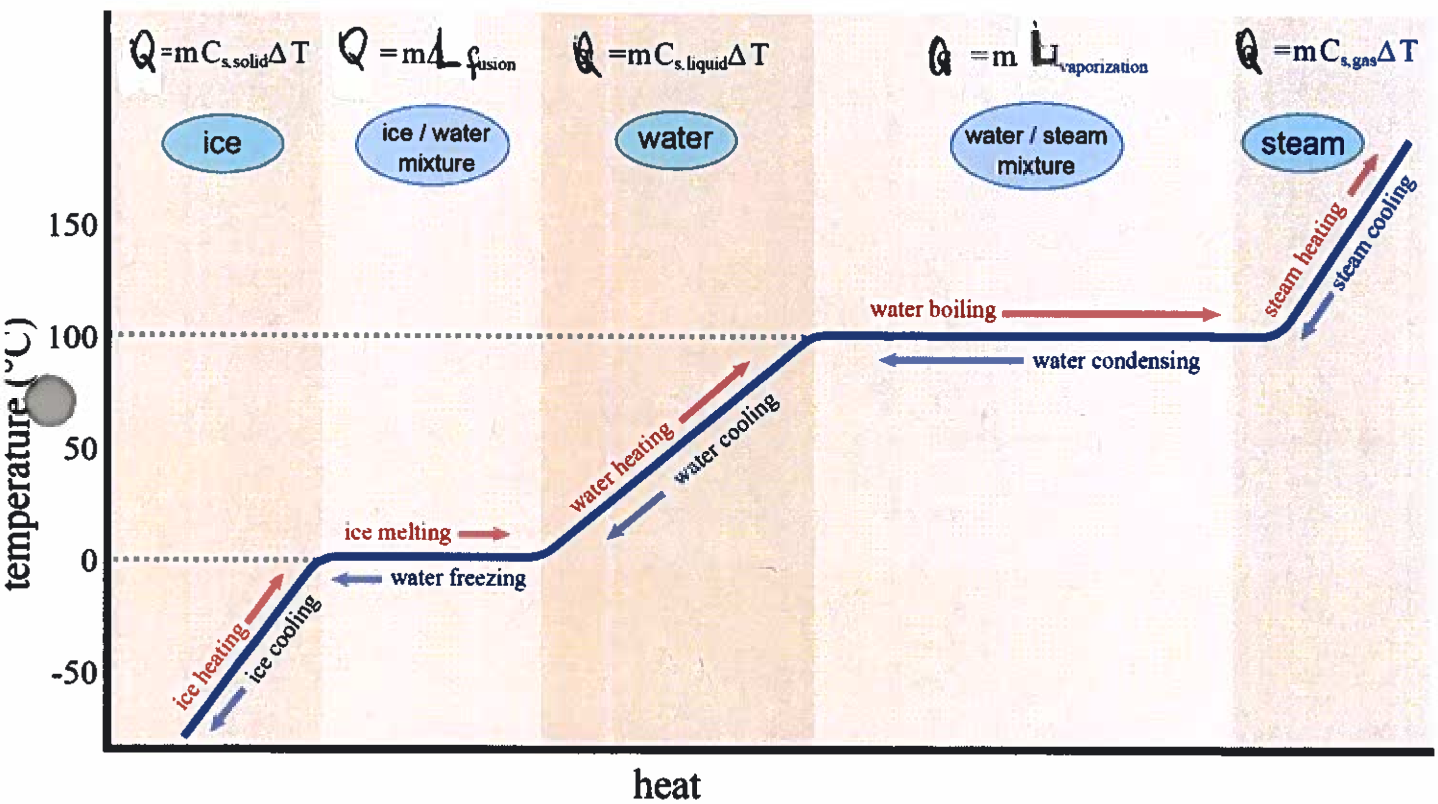
$$= \frac{1\,kg}{0.3\,kg} \cdot 4186 \frac{J}{kg \cdot ^\circ C} \cdot \frac{10^\circ C}{65^\circ C} = 2146\ J/kg\,^\circ C$$

But what if we drop a really big really hot test mass into the water? Above $100^\circ C$ water changes from liquid to gas

Microscopically molecules have enough energy to break the bonds (just like a really fast rocket can escape the gravity well of the Earth)
But this bond breaking requires extra energy, even if the temperature not changing

$Q = L \cdot m$

$L$ - latent heat




for $H_2O$

Specific heat

$C_{ice} = 2090\ J/kg \cdot °C$

$C_{water} = 4184\ J/kg \cdot °C$

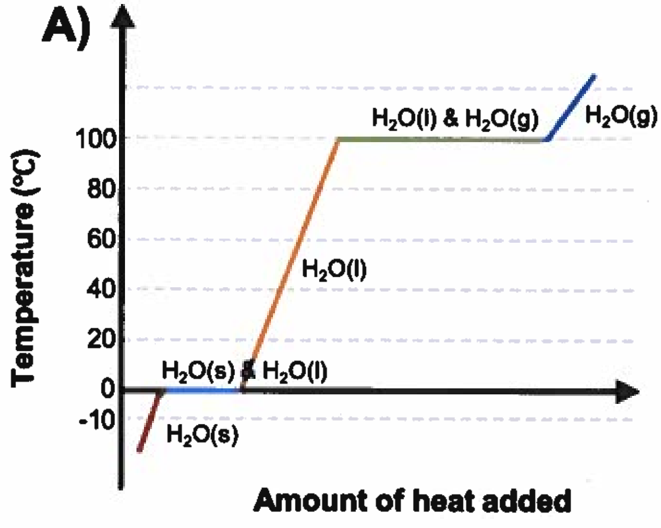
$C_{steam} = 2030\ J/kg \cdot °C$

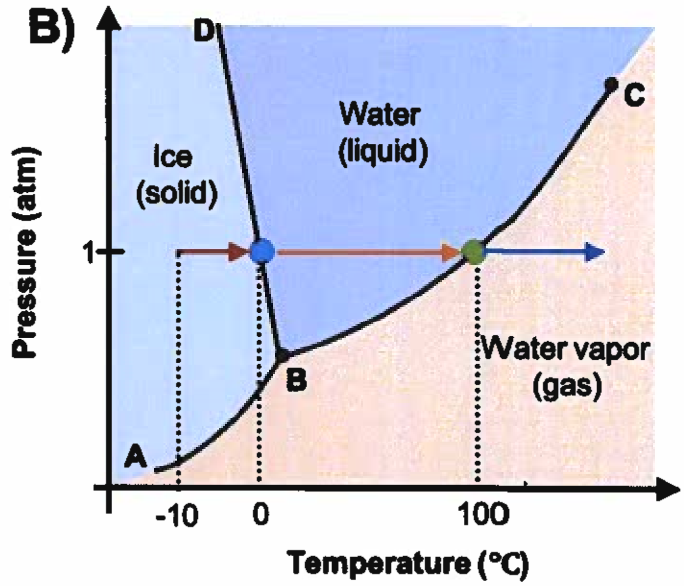
Latent heat

$L_{fusion} = 334\ J/g = 3.34 \cdot 10^5\ J/kg$

$L_{vaporization} = 2260\ J/g =$

$= 2.26 \cdot 10^6\ J/kg$


A)
Temperature (°C)
100
80
60
40
20
0
-10
H₂O(l) & H₂O(g)
H₂O(g)
H₂O(l)
H₂O(s) & H₂O(l)
H₂O(s)
Amount of heat added


B)
D
Water
(liquid)
Ice
(solid)
C
Pressure (atm)
1
Water vapor
(gas)
B
A
-10
0
100
Temperature (°C)