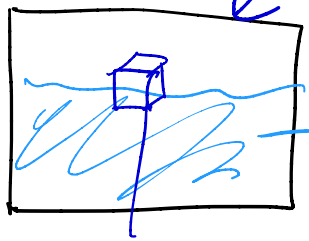
Entropy

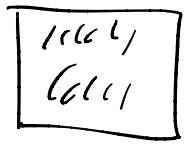


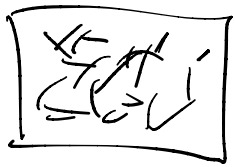
1st Law of Thermodynamics

$\Delta E = q_{water} + q_{ice}$

$= 0$

The 1st Law of Thermodynamics cannot explain why ice not only melts but also forms, since in an isolated system both phase changes would automatically have an overall energy change of zero.


solid


liquid


gas

A gas has more entropy than a liquid, which in turn has more entropy than a solid. This is because a gas has freedom of movement and therefore more ways energy can be distributed.

$MgSO_4(s) \xrightarrow{H_2O} MgSO_4(aq)$

When magnesium sulfate dissolves, the solution becomes cold, meaning that heat is absorbed from water as the solution forms ⟶ endothermic process

This process can occur because the entropy change outweighs the enthalpy change.

Since a solution allows for far greater distribution of matter and energy, forming a solution involves a large, positive entropy change, which has an energetic effect that makes up for the fact that heat must be absorbed for the process to occur.

Matter and energy spreading out is a release of energy itself,

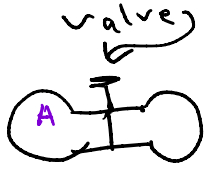
$S = k \ln W$

- S: entropy
- k: Boltzmann Constant, $k = R/N_A$
- W: the number of possible configurations of energy (microstates)

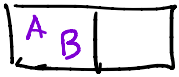
Individual ways in which the same amount of matter or energy can be distributed in a systmem.

W

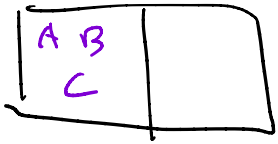
1



1 — 2 molecules



1 — 3 molecules

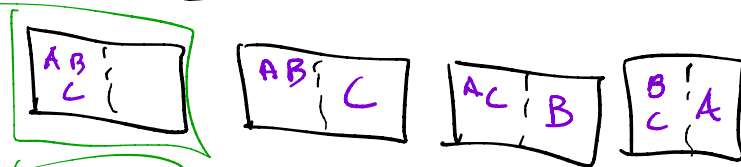


W (# of microstates)

$2 = 2^1$

$4 = 2^2$

$8 = 2^3$



In general, if there are n molecules in this gas example, there are 2^n microstates created by doubling volume.

Entropy is a state function, meaning that the number of ways matter and energy can be distributed depends only on the state the system is in, not how the state formed.

Boltzmann constant

$$S_{init} = k \ln W = k \ln 1 = 0$$

$$S_{expand} = k \ln W = k \ln 2^n$$

For 1 mol ($n = 6.02 \times 10^{23}$), $S = 5.76\ J/K$

$$\Delta S_{sys} = \frac{q_{rev}}{T}$$