

Units and Values

Avogadro’s number: $N_A = 6.023 \times 10^{23}$
Boltzmann’s cosntant: $k = 1.381 \times 10^{-23}$ J/K
Boltzmann’s cosntant: $k = 8.617 \times 10^{-5}$ eV/K

$R = 8.31\text{J/Kmol}$, $nR = Nk$.
Calorie: 1cal = 4.2J (amount of heat needed to raise the temperature of a gram of water by 1°C.)
Absolute zero: 0K = −273.15°C.
Plank’s constant: 6.626×10^{-34} J/s, 4.136×10^{-15} eV/s
One atmosphere pressure: 1atm = 101.325kPa.
One atomic mass unit: $1.6605402 \times 10^{-27}$ kg

Definitions

Isotherm: constant temperature
Isobar: constant pressure
Isothermal compression: so slow that temperature does not rise
Adiabatic: so fast heat does not escape ($Q = 0$)
Quasi static: Process with constant entropy ($\Delta S = 0$), slow process so that system can always equilibrate to the changing conditions.
Isentropic: Adiabatic ($Q = 0$) and quasistatic ($\Delta S = Q/T = 0$)
isochoric: Constant volume $\Delta V = 0$. **fundamental assumption of statistical mechanics**: In an isolated system in thermal equilibrium, all accessible microstates are equally probable
thermodynamic limit: Let number of particles go to infinity $P = T \left(\frac{\partial S}{\partial V} \right)_{U,N}$, $T^{-1} = \left(\frac{\partial S}{\partial U} \right)_{N,V}$, $\mu = -T \left(\frac{\partial S}{\partial N} \right)_{U,V}$

The laws of thermodynamics

- 1. Defines temperature
- 2. Defines energy (U), and energy transformation (W, Q).
- 3. Defines entropy S
- 4. Fixes the value of entropy at $T = 0\text{K}$.

Important formulas

$$\Omega_{total}(N, n) = \Omega_A(N, n)\Omega_B(N, n)$$
$$\Delta S = 0 \rightarrow \text{reversible}$$
$$\Delta S \geq 0 \rightarrow \text{irreversible}$$

Important facts

- Macrostates with maximum multiplicity are the most likely and they correspond to the average values
- ΔS is path independent, choose the simplest path between the states to compute

Mol

For a given substance, find it’s atmoic mass, in units of u (hydrogen 1u, oxygen 16u, lead 207u). Then the mole is the mass of the substance in gram (for example 1g), and divide with the atomic mass (for example 16u). The substance is then 1/16mole.

Ideal Gas

Low density gas $PV = NkT$, ignore particle interactions. $\langle v^2 \rangle = 3kT/m$. Compression $W = NkT \ln V_i/V_f$. For indistinhuishable particles (we get $1/N!$)

$$\Omega(U, V, N) = f(N)V^N U^{3N/2} \approx \frac{V^N}{N!(3N/2)!} (2\pi mU/h^2)^{3N/2}$$
$$S = Nk \left[\ln \left(\frac{V}{N} \left(\frac{4\pi mU}{3Nh^2} \right)^{3/2} \right) + \frac{5}{2} \right] \text{ (Sackur-Tetrode)}$$
$$\Delta S = Nk \ln V_f/V_i \quad (U, N \text{ fixed})$$

Equipartition theorem

For each quadratic degree of freedom (kinetic v^2 , spring k^2 , rotational ω^2 , etc.) contributes to $1/2kT$ to the average energy at equilibrium at temperature T . For N particles with f degrees of freedom $U = NfkT/2$. Safest to apply in changes of energy. Coutning f : one for each dimension, 2 from rotation (as one axis is symmetric), vibrateing counts twice (frozen out at room temperature) in a solid 3 vibrational directions $f+=6$. Equipartition theorem only works for kinetic energy in liquids. Use $f = 3$ for a monatmoic gas and $f = 5$ for a diatomic gas, 6 for polyatomic.
Derived by Boltzmannfactors and converting to integral

$$Z = \frac{1}{\Delta q} \sum_q e^{-\beta c q^2} \Delta q$$

and calculating energy from partition function

Heat and Work

Heat = any spontaneous flow of energy from one object to another due to a difference in temperature; konduction, konvection, radiation.
Work = any other transfer of energy inn and out of system. Both refer to energy in transit.

$$\Delta U = Q + W$$

Assuming quasistatic compression $W = -P\delta V$. Use pressure from POV of gas, negative work means gas is doing work, positive work means the outside is doing work on the gas.

$$W = - \int_{V_i}^{V_f} P(V) dV \quad \text{quasistatic}$$

For a quasistatic process $Q = TdS$.

Heat Capacity

Heat capacity: amount of heat needed to raise temperature per degree temperature increase $C = Q/\Delta T$.
Specific heat capacity: Heat capacity per unit mass $c = C/m$.

$$C = (\Delta U - W)/\Delta T, C_V = \left(\frac{\partial U}{\partial T} \right)_V, C_P = \left(\frac{\partial U}{\partial T} \right)_P + P \left(\frac{\partial V}{\partial T} \right)$$

At the phase transition you can put heat into system without increasing temperature, $C = \infty$. Latent heat $L = Q/m$ is the heat needed to accomplish phase transition per mass, assume constant pressure 1atm, and no other work.

$$C_V = T \left(\frac{\partial S}{\partial T} \right)_V \quad C_P = T \left(\frac{\partial S}{\partial P} \right)_V$$

$$C_V \rightarrow 0 \quad \text{as} \quad T \rightarrow 0$$
$$C_V \rightarrow \text{constant} \quad \text{as} \quad T \rightarrow \infty$$

$$dS = \frac{dU}{T} = \frac{Q}{T} = \frac{C_V dT}{T} \rightarrow \Delta S = \int_0^{T_f} \frac{C_V}{T} dT$$

Adiabatic compression

So fast no heat flows out, still quasistatic, $\Delta U = W$.

$$V^\gamma P = \text{constant}, \gamma = (f + 2)/f$$

Derived from $\Delta U = W$, from equipartition $dU = fNk/2dT = -PdV$, plugging in ideal gas law: $f/2dT/T = -dV/V$.

Two state system

Multiplicity of the macrostate with N particles where N_\uparrow point up

$$\Omega(N, n) = \frac{N!}{N_\uparrow!(N - N_\uparrow)!} = \frac{N!}{N_\uparrow!N_\downarrow!}$$

Einstein solid

$N/3$ atoms as harmonic oscillators which can oscillate in 3 dimensions with $q = (U_n - N\hbar\omega/2)/\hbar\omega$ units of energy

$$\Omega(N, q) = \binom{q + N - 1}{q} = \frac{(q + N - 1)!}{q!(N - 1)!} \approx \frac{(q + N)!}{q!N!} N_\downarrow!$$

If a composite system: $q = q_A + q_B$ and $N = N_a + N_b$ and $\Omega = \Omega_A \Omega_B$.

$$\text{low } T \quad \Omega \approx \left(\frac{Ne}{N} \right)^q, \quad q < < N$$
$$\text{high } T \quad \Omega \approx \left(\frac{qe}{N} \right)^N, \quad N < < q$$
$$S = kN \left[\ln \left(\frac{q}{N} + 1 \right) \right], \quad q = U/\hbar\omega$$

thermodynamic potentials

- Internal Energy: U Total internal thermal energy of the system. The change in the stored energy equal the sum of «energies in transit» $\Delta U = Q + W$ (first law of thermodynamic). The infinitesimal change in internal energy: $dU = \delta Q + \delta W$. For infinitesimal reversible process: $dU = TdS - PdV$. Can also be $dU = TdS - PdV + \mu dN$.

- Enthalpy: $H = U + PV$, $\Delta F \leq W$ (at constant T). Energy needed to create the system out of nothing, and make room for it. $dH = TdS + VdP + \mu dN$

- Helmholtz free energy: $F = U - TS$, (constant pressure). Total energy needed to create the system, minus the heat you get from enviornment. Thermodynamic identity: $dF = -SdT - PdV + \mu dN$. Minimized in thermal bath, with only energy exchange. $\Delta F \leq W$. $\mu = \left(\frac{\partial F}{\partial N} \right)_{T,V}$, $S = - \left(\frac{\partial F}{\partial T} \right)_{V,N}$, $P = - \left(\frac{\partial F}{\partial V} \right)_{T,N}$.

- Gibbs free energy: $G = U - TS + PV = \mu N$, $\Delta G \leq W_{other}$. Create system in constant temperature and pressure. Minimized when change in volume with constant pressure. $dG = -SdT + VdP + \mu dN$. $\mu = \frac{\partial G}{\partial N}$, $V = \frac{\partial G}{\partial P}$, $-S = \frac{\partial G}{\partial T}$

- Chemical potential, the energy increase by adding a particle in to the system with constant temperature and pressure. Diffusive equilibrium \rightarrow equal chemical potentials. $\mu = \left(\frac{\partial G}{\partial N} \right)_{T,P}$, $\mu = -T \left(\frac{\partial S}{\partial N} \right)_{U,V}$, $\mu = \left(\frac{\partial U}{\partial N} \right)_{S,V}$ Particles tend to flow from the system with higher μ to the system with lower μ .
For monatomic ideal gas $\mu = -kT \ln \left(\frac{V Z_{int}}{N V Q} \right)$ must be equal

At constant U and T , S increases.
At constant V and T , F decreases.
At constan T and P , G decreases.

Reversible process

At every step of the path the system is at equilibrium. For infinitesimal reversible process: $dU = TdS - PdV$.
Clausius equality: $dS = \delta Q_{rev}/T$, constant volume, no work. Irreversible heat transfer is smaller than the reversible heat exchange at a given T

$$\frac{\delta Q_{irrev}}{T} < \frac{\delta Q_{rev}}{T}$$

Clausius inequality: $dS \geq \delta Q/T$. For an isolated system entropy tends to increase as the system spontaneously finds its equilibrium state $dS \geq 0$.

$$C_V = \left(\frac{\delta Q}{dT} \right)_V = \left(\frac{\partial U}{\partial T} \right) \rightarrow dU = C_V dT$$

Isolated system at equilibrium

Multiplicity of a macrostate $\Omega(U, V, N)$ counts all equally likely accessible microstates, if the particles are indistinguishable the total number of microstates is reduced by $N!$. Boltzmann’s formula $S = k \ln \Omega = -k \sum_s P(s) \ln P(s)$, for equilibrium state at fixed U , entropy is maximized at equilibrium, $dS = 0$.

$$dS = \frac{1}{T} dU + \frac{P}{T} dV - \frac{\mu}{T} dN$$

For a composite system

$$\frac{\partial S_A}{\partial U_A} = \frac{\partial S_B}{\partial U_B} \quad (\text{thermal equilibrium})$$
$$\frac{\partial S_A}{\partial V_A} = \frac{\partial S_B}{\partial V_B} \quad (\text{mechanical equilibrium})$$
$$\frac{\partial S_A}{\partial N_A} = \frac{\partial S_B}{\partial N_B} \quad (\text{diffusive equilibrium})$$

Extensive and Intensive Quantities

An extensive quantity doubles when you double a system. The quantities which are unchanged when you double the system are intensive.

- **Extensive**: $V, N, S, U, H, F, G, \text{mass}$
- **Intensive**: $T, P, \mu, \text{density}$

extensive \times intensive = extensive
 $\frac{\text{extensive}}{\text{extensive}} = \text{intensive}$
extensive \times extensive = neither, you did something wrong
extensive + extensive = extensive
extensive + intensive = not allowed

Van der Waals Equation

$$\left(P + \frac{aN^2}{V^2} \right) (V - Nb) = NkT$$

Nb makes it impossible to compress to zero volume. The a term accounts for short range attractive forces. Low density gas $PV = NkT$, ignore particle interactions. $\langle v^2 \rangle = 3kT/m$. Compression $W = NkT \ln V_i/V_f$.

Maxwell Construction and phase transition

Phase transiton is a discontinuous change in the substance properties when the surroundings changes infinitesimally. Makes it possible to find the pressure at the phase transition from the PV -diagram, create straight line which makes even areas in the PV -line. Gives us the point of phase transition between liquid and gas. Gas at high volume, liquid at low volume. The intersected regions are unstable. The phase transition happens at constant PRESSURE! We find the pressure on the P - V diagram by the equal areas construction of a given isotherm. Phase transition from liquid to vapor at a constant Gibbs free energy $G(P, T, N)$, since the entropy effects the temperature dependence of the Gibbs, and the pressure effects the volume dependence of the gibbs, this connects the gibbs free energies $dG_{liquid} = dG_{gas}$.
Clausius Clapeyron relation

$$-S_l dT + V_l dP = -S_g dT + V_g dP$$

$$\frac{dP}{dT} = \frac{S_g - S_l}{V_g - V_l} = \frac{\Delta S}{\Delta V} = \frac{L}{T \Delta V}$$

Entropy jumps going from a liquid to a gas. Volume expansion going from a liquid to a gas. The Clausius Clapeyron relation tells us how much the phase transition pressure changes with changing temperature. Derivation: $dG = SdT + VdP + \mu dN$. Along the phase boundary where we at any point have chemical equilibrium between substances, the changes dG for solid and gas must be the same for changes dT and dP . $S_g dT + V_g dP = S_s dT + V_s dP$.

Sterlings approximation

$$N! \approx N^N e^{-N} \sqrt{2\pi N}$$
$$N! \approx N^N e^{-N}$$
$$\ln(N!) \approx N \ln N - N$$
$$\ln \binom{N}{n} \approx N \ln N - n \ln n - (N - n) \ln (N - n)$$

Maxwell relations

Functions are usually well behaved, so order of differential does not matter, can then use

$$\frac{\partial}{\partial V} \left(\frac{\partial U}{\partial S} \right) = \frac{\partial}{\partial S} \left(\frac{\partial U}{\partial V} \right)$$

By using thermodynamic identities for dU one can derive the maxwell relations, here they are

$$\frac{\partial^2 U}{\partial S \partial V} \rightarrow \left(\frac{\partial T}{\partial V} \right)_S = - \left(\frac{\partial P}{\partial S} \right)_V$$
$$\frac{\partial^2 H}{\partial S \partial P} \rightarrow \left(\frac{\partial T}{\partial P} \right)_S = + \left(\frac{\partial V}{\partial S} \right)_P$$
$$- \frac{\partial^2 F}{\partial T \partial V} \rightarrow \left(\frac{\partial S}{\partial V} \right)_T = + \left(\frac{\partial P}{\partial T} \right)_V$$
$$\frac{\partial^2 G}{\partial T \partial P} \rightarrow \left(\frac{\partial S}{\partial P} \right)_T = - \left(\frac{\partial V}{\partial T} \right)_P$$

