

CHEM 26200 (Thermodynamics) Notes

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Part I

Statistical Mechanics

Week 1

The Boltzmann Factor and Partition Functions

1.1 Overview of Major Results

- 1/10:
- In this course, we will review thermochemistry from intro chem, but go deeper with statistical mechanics.
 - TA: Haozhi.
 - Did his undergrad at Oxford.
 - Has already taught this class in the PME.

- **Boltzmann constant:** The following constant. *Denoted by k_B . Given by*

$$k_B = 1.381 \times 10^{-23} \text{ J/K}$$

- Equal to the quotient of the ideal gas constant and Avogadro's constant.
- **Ideal gas law:** The following relationship between the pressure P , volume V , number of moles n , and temperature T of an ideal gas, and the ideal gas constant R .

$$PV = nRT$$

- Multiplying by the quotient of Avogadro's constant with itself yields

$$PV = nN_A \frac{R}{N_A} T$$

$$PV = Nk_B T$$

where N is the number of molecules in the system.

- The unit for PV is Joules.
 - Thus, the above form states that PV is equal to the number of particles times a tiny unit of energy.
- Relating PV to the kinetic energy of gas molecules/atoms^[1].
 - Pressure originates microscopically from the collisions of particles with the walls of their container.
 - As such, we first seek to derive an expression for the number of collisions per second per area.

¹This derivation differs from that on Labalme (2021a, pp. 3–4) and Labalme (2021d, pp. 18–19), in that its approach is from a flux perspective.

- Consider the number $N(v_x)$ particles with speed v_x in the x -direction.
 - The quotient $N(v_x)/V$ is the density in the container of particles with speed v_x .
 - Thus, the flux “through”/to/at the wall is this density, times the area of the wall, times the x -velocity of the particles.
- Assume an elastic collision of each particle with the wall. Thus, when each particle of mass m collides with the wall, it transfers $2mv_x$ of momentum.
 - Therefore, since $F = dp/dt$, the overall force exerted on the wall by the gas particles moving with speed v_x is $2mv_x \cdot N(v_x)/V \cdot v_x \cdot \text{Area}$ times per second.
 - But, of course, we must sum over all possible v_x , so the total force

$$F = \int_{v_x > 0} 2mv_x \cdot \frac{N(v_x)}{V} \cdot v_x \cdot \text{Area} dv_x$$

- It follows that

$$\begin{aligned} P &= \frac{F}{\text{Area}} \\ &= \int_{v_x > 0} 2mv_x^2 \cdot \frac{N(v_x)}{V} dv_x \end{aligned}$$

The factor of $1/2$ in the following line comes from the fact that we are only integrating over half of the possible v_x^2 s (i.e., the positive ones).

$$\begin{aligned} &= 2m \cdot \frac{N}{V} \cdot \frac{1}{2} \langle v_x^2 \rangle \\ &= \frac{N}{V} m \langle v_x^2 \rangle \\ PV &= Nm \cdot \langle v_x^2 \rangle \end{aligned}$$

Assuming that the gas is not moving in any one direction means that $\langle v_x^2 \rangle = \langle v_y^2 \rangle = \langle v_z^2 \rangle = \frac{1}{3} \langle v^2 \rangle$. Therefore,

$$\begin{aligned} &= Nm \cdot \frac{1}{3} \langle v^2 \rangle \\ &= \frac{2}{3} N \cdot \frac{1}{2} m \langle v^2 \rangle \\ &= \frac{2}{3} N \cdot \langle E_{KE} \rangle \\ \langle E_{KE} \rangle &= \frac{3}{2} \frac{PV}{N} \\ \langle E_{KE} \rangle &= \frac{3}{2} k_B T \end{aligned}$$

- Note that this applies to all sorts of regimes — we used no properties of the particles (e.g., atom vs. molecule) to derive this relationship.
- Getting the distribution of the gas energies or speed is the next logical step.
 - First, though, we consider alternate occurrences of $k_B T$.
- The activation energy of Arrhenius (1889): “To collide is to react” is inaccurate; it must collide with sufficient energy. The molecule must be “activated.”

$$k = Ae^{-E_a/RT} = Ae^{-E_a/k_B T}$$

- The first E_a is the molar energy of activation; the second is the molecular energy of activation.

- Yields the probability distribution of a molecule reacting.
- Nernst equation:

$$E_{\text{cell}} = E_{\text{cell}}^0 - \frac{RT}{nF} \ln Q$$

- $\ln Q$ is the ratio inside vs. outside the membrane.
 - $F = N_A e$ where e is the charge of an electron.
 - Thus,
- $$\Delta E = \frac{RT}{nF} = \frac{k_B T}{ne}$$
- If the potential across the membrane is approximately $k_B T$, then $\ln Q \approx 1$, so $Q \approx e$.
 - Thus, at body temperature ($T = 310 \text{ K}$), $k_B T/e = 26 \text{ mV}$.
 - The speed of sound: Certainly sound cannot travel faster than the molecules. Therefore, we can derive the following approximation for the speed of sound.

$$\begin{aligned} \frac{1}{2} m \langle v^2 \rangle &= \frac{3}{2} k_B T \\ \sqrt{\langle v^2 \rangle} &= \sqrt{\frac{3k_B T}{m}} \\ v_{\text{rms}} &= \sqrt{\frac{3k_B T}{m}} \end{aligned}$$

- This estimate is within 20 – 30 % — take m to be the average mass of air.
- de Broglie wavelength: A molecule has a kinetic energy approximately equal to $k_B T$. Additionally, the quantum mechanical kinetic energy of a molecule aligns with this, as $\hbar^2 k^2 / 2m \approx k_B T$. Furthermore, the particle-wave duality relates the momentum to wavelength by $p = \hbar k = h/\lambda$. Therefore,

$$\lambda \approx \sqrt{\frac{h^2}{2mk_B T}}$$

- Thus, a gas at STP has a very small de Broglie wavelength and behaves classically.
- Only at very low temperatures with very light gasses do quantum considerations come into play.
- A H_2 molecule at 300 K has de Broglie wavelength $\lambda = 1.78 \text{ \AA}$.
- Note that the quantum mechanical kinetic energy of a free particle is derived as follows.

$$\begin{aligned} \hat{H}\psi &= E\psi \\ -\frac{\hbar^2}{2m} \frac{\partial^2}{\partial x^2} (e^{ikx}) &= E e^{ikx} \\ \frac{\hbar^2 k^2}{2m} e^{ikx} &= E e^{ikx} \\ E &= \frac{\hbar^2 k^2}{2m} \end{aligned}$$

- **Boltzmann factor:** Gives the relative probability p_2/p_1 of two states E_1, E_2 , provided their respective energies E_1, E_2 . Given by

$$\frac{p_2}{p_1} = e^{-(E_2 - E_1)/k_B T}$$

- Consider states E_1, E_2, E_3, \dots , denoted by their energies.
- Consistency check: Given

$$\frac{p_2}{p_1} = e^{\frac{-(E_2 - E_1)}{k_B T}} \qquad \frac{p_3}{p_2} = e^{\frac{-(E_3 - E_2)}{k_B T}}$$

we do indeed have

$$\frac{p_3}{p_1} = \frac{p_3}{p_2} \cdot \frac{p_2}{p_1} = e^{\frac{-(E_3 - E_2)}{k_B T} + \frac{-(E_2 - E_1)}{k_B T}} = e^{\frac{-(E_3 - E_1)}{k_B T}}$$

- We'll take this as God-given for now. Boltzmann derived it with a very good knowledge of the thermodynamics of freshman chemistry.
- We're starting with the above exciting result, and then going back and building up to it over the next three weeks.
- We write the Boltzmann factor for degenerate states as follows.
 - Consider four states at E_2 and one state at E_1 .
 - The Boltzmann factor still tells us that $p_2/p_1 = e^{-(E_2-E_1)/k_B T}$, but we have to make the following adjustment. Indeed, the total probability of being in one of the four states at energy E_2 is $p(E_2) = 4p_2$, while the total probability of being in the one state at energy E_1 is still just $p(E_1) = p_1$.
 - In each state E_2 ,

$$\frac{p(E_2)}{p(E_1)} = \frac{N_2}{N_1} e^{-(E_2-E_1)/k_B T}$$

- The weekly quiz.
 - The first quiz will be next week.
 - A Canvas quiz – we'll have 24 hours to take it, but only 1 hour to take it.

1.2 Boltzmann Factor Examples / Partition Function

- 1/12:
- We will apply the Boltzmann factor to electronic, magnetic, translational, rotational, and vibrational molecular states.
 - Example: Sodium lamp – two lines at 589.6 nm and 589.0 nm with intensity ratio 1:2.

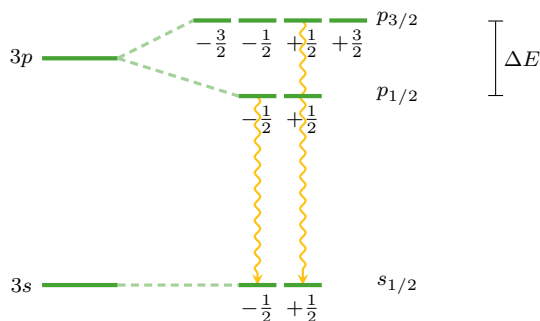


Figure 1.1: Sodium lamp energy levels.

- Street lamps use this (very efficient).
- Also used in astronomy.
- In the sodium atom, there are two energy levels (3s and 3p).
- The states have a spin-orbit coupling effect.
 - 3s (with $S = 1/2$) splits into two degenerate states $s_{\pm 1/2}$ based on spin.
 - 3p (with $L = 1$ and $S = 1/2$) splits into two nondegenerate states ($l = \pm 1$ [called $p_{3/2}$] and $l = 0$ [called $p_{1/2}$]), which further subdivide into four (resp. two) degenerate states ($-3/2, -1/2, 1/2, 3/2$ and $-1/2, 1/2$).

- Let ΔE be the difference in energy between the $p_{3/2}$ and $p_{1/2}$. Then

$$\frac{\Delta E}{k_B} = \frac{1}{k_B} \left(\frac{hc}{\lambda_1} - \frac{hc}{\lambda_2} \right) = 25 \text{ K}$$

where $\lambda_1 = 589.6 \text{ nm}$ and $\lambda_2 = 589.0 \text{ nm}$.

- Thus, $e^{-\Delta E/k_B T} \approx 1$ for $T = 300 \text{ K}$ (the temperature in the sodium vapor lamp).
- Therefore,

$$\begin{aligned} \frac{p(E_2)}{p(E_1)} &= \frac{4}{2} \cdot 1 \\ p(E_2) &= 2p(E_1) \end{aligned}$$

- Example: MRI.

- The magnetic field polarizes the spins of the hydrogen protons in our body with $\Delta E = \mu_B B$.
- If we also take $B = 6 \text{ T}$ and $T = 310 \text{ K}$ (body temperature), then

$$\frac{\mu_B B}{k_B T} = 2 \times 10^{-5}$$

- Thus, very few protons actually flip, but with modern technology we can still measure this.

- **Proton magnetic moment:** The magnetic moment of a proton. Denoted by μ_B . Given by

$$\mu_B = 1.4 \times 10^{-26} \text{ J/T}$$

- Example: Rotational.

- The rotational energy E_J of a molecule depends on the angular momentum quantum number J and the moment of inertia of the molecule $I = \mu R^2$ via the following relation.

$$E_J = \frac{\hbar^2}{2I} J(J+1)$$

- Microwave spectroscopy can be used to find molecules out in the universe.
- At 300 K,

$$\frac{p(J=1)}{p(J=0)} = \frac{3}{1} e^{\frac{-(E_1-E_0)}{k_B T}} = 2.95$$

■ As before $J = 1$ corresponds to states $j = -1, 0, 1$.

- There is a range of angular momenta due to the temperature that for $T = 300 \text{ K}$ peaks around $J = 5$.
- See Figure 2.1.

- Example: Vibrational.

- Here, $\Delta E = E_n - E_{n-1} = h\nu$ for every energy level since $E_n = h\nu(n + 1/2)$.
- It follows that

$$\frac{h\nu}{k_B} = 2800 \text{ K}$$

for CO, meaning that at 300 K, CO will be largely in its ground state.

- The partition function tells us everything we would want to know about a system.

$$Q = \sum_i e^{-E_i/k_B T}$$

- All we need to know is the energy of every state in the system.
- This is impossible for an infinite system, but the Schrödinger equation gives us the energy of a system, so its a great place to start.
- Calculating the average energy of a system from the partition function.
 - Take the energy E_1 of the ground state to be our zero of energy.
 - Under this assumption, we may rewrite the partition function as follows.

$$Q = \sum_i e^{-(E_i - E_1)/k_B T} = \frac{p_1}{p_1} + \frac{p_2}{p_1} + \frac{p_3}{p_1} + \dots = 1 + e^{\frac{-(E_2 - E_1)}{k_B T}} + e^{\frac{-(E_3 - E_1)}{k_B T}} + \dots$$

- From MathChapter B, the expected energy is equal to

$$\langle E \rangle = E_1 p_1 + E_2 p_2 + E_3 p_3 + \dots$$

- Since $E_1 = 0$, the above can be rewritten as

$$\langle E \rangle = p_1 \left[E_2 \frac{p_2}{p_1} + E_3 \frac{p_3}{p_1} + \dots \right]$$

- With this result established, we now take a moment to note that

$$\frac{\partial}{\partial T} \left(e^{-E_i/k_B T} \right) = \frac{E_i}{k_B T^2} e^{-E_i/k_B T} = \frac{1}{k_B T^2} \left(E_i \frac{p_i}{p_1} \right)$$

and

$$\begin{aligned} p_1 &= 1 - (p_2 + p_3 + \dots) \\ &= 1 - p_1 \left(\frac{p_2}{p_1} + \frac{p_3}{p_1} + \dots \right) \\ &= 1 - p_1 (Q - 1) \\ p_1 &= \frac{1}{Q} \end{aligned}$$

- Therefore,

$$\begin{aligned} \langle E \rangle &= p_1 k_B T^2 \frac{\partial}{\partial T} \left(\frac{p_1}{p_1} + \frac{p_2}{p_1} + \dots \right) \\ &= \frac{1}{Q} k_B T^2 \frac{\partial Q}{\partial T} \\ \langle E \rangle &= k_B T^2 \frac{\partial \ln Q}{\partial T} \end{aligned}$$

- The above is an important result.

- Changing the origin of energy.
 - We know that

$$\begin{aligned} Q(E_0) &= Q(E'_0) e^{-(E'_0 - E_0)/k_B T} \\ \ln Q(E_0) &= \ln Q(E'_0) - \frac{E'_0 - E_0}{k_B T} \end{aligned}$$

– Thus,

$$\begin{aligned}\langle E \rangle_{E_0} &= k_B T^2 \frac{\partial}{\partial T} (\ln Q(E_0)) \\ &= k_B T^2 \left(\frac{\partial}{\partial T} (\ln Q(E'_0)) - \frac{\partial}{\partial T} \left(\frac{E'_0 - E_0}{k_B T} \right) \right) \\ &= \langle E \rangle_{E'_0} + (E'_0 - E_0) \\ \langle E \rangle_{E_0} + E_0 &= \langle E \rangle_{E'_0} + E'_0\end{aligned}$$

– So the change of the energy origin does indeed change the total energy by the same amount.

1.3 Calculating Average Energies

- 1/14:
- We derived that for an ideal gas, $\langle E \rangle = 3k_B T/2$. But this may change at higher pressures.
 - Calculating the average kinetic energy at higher temperatures.
 - Use the main result from last time, which gives us the energy in terms of the partition function.
 - We have different degrees of freedom since KE and PE are on different coordinates (KE is on speed and PE is on position).
 - When we write the Boltzmann factor, we'll have an exponential with the sum of the kinetic and potential energy.

$$Q = \sum_{ij} e^{-(E_{KE_i} - E_{PE_j})/k_B T} = \sum_{ij} e^{-E_{KE_i}/k_B T} e^{-E_{PE_j}/k_B T} = Q_{KE} Q_{PE}$$

- The second equality holds because KE depends on the velocity coordinates and PE depends on position coordinates; thus, they are independent.
- Kinetic energy partition function.

$$E_{KE} = \frac{1}{2} m v_x^2$$

■ Thus,

$$Q_{KE_{v_x}} = \int_{-\infty}^{\infty} e^{-\frac{1}{2} m v_x^2 / k_B T} dv_x = \sqrt{\frac{2\pi k_B T}{m}}$$

- This function doesn't depend on anything of significant import.
- It follows that

$$\langle KE_x \rangle = k_B T^2 \frac{\partial}{\partial T} (\ln Q_{KE_{v_x}}) = k_B T^2 \frac{\partial}{\partial T} \left(\ln \sqrt{\frac{2\pi k_B T}{m}} + \frac{1}{2} \ln T \right) = \frac{k_B T}{2}$$

and

$$\langle KE \rangle = \langle KE_x \rangle + \langle KE_y \rangle + \langle KE_z \rangle = \frac{3}{2} k_B T$$

- Therefore, this result holds beyond the specific case of an ideal gas!^[2]
- Now for the potential energy of a harmonic oscillator.
 - $PE = \frac{1}{2} k x^2$; calculate the partition function for the coordinate x .

$$Q_x = \int_{-\infty}^{\infty} e^{-\frac{1}{2} k x^2 / k_B T} dx = \sqrt{\frac{2\pi k_B T}{k}}$$

²This is why this derivation is preferable over the one presented in the book. Note that the “gas” here is assumed to be made of hard spheres.

- Thus,

$$\langle PE_x \rangle = \frac{k_B T}{2}$$

- For a 3D harmonic oscillator,

$$\langle PE \rangle = \frac{3}{2} k_B T$$

- Average potential energy of a gravitational potential.

- Apply the virial theorem (relates the average kinetic energy of a system in a conservative potential to the potential energy).
- Since we've shown that for any system, the average kinetic energy in one dimension is $k_B T/2$, the potential in any system will be related (i.e., have a factor of $k_B T$).

- What it means to cool something down, if KE always follows the same formula.

- Although the formula does not change, $\langle KE \rangle \propto T$, so decreasing the temperature decreases the kinetic energy.
- Similarly, as things change phase, more and more potentials take hold (e.g., in the gas phase, there is no potential energy, but there is significant potential energy in the solid and liquid phases).

- Rotational kinetic energy.

- Consider N_2 , with its two rotational degrees of freedom.
- Classically,

$$E_{\text{rot}} = \frac{1}{2} I \omega^2$$

- Thus, once again,

$$Q_\omega = \int_{-\infty}^{\infty} e^{-\frac{1}{2} I \omega^2 / k_B T} d\omega = \sqrt{\frac{2\pi k_B T}{I}}$$

making

$$\langle E_{\text{rot}} \rangle = \frac{k_B T}{2}$$

for one degree of freedom.

- **Law of Dulong and Petit:** The heat capacity of elemental solids is about $3nR$.

- Observed in 1819.
- A major result in an era where atomic structure was just emerging.
- Imagine an atom bound in a three-dimensional (octahedral) potential. It's energy is thus

$$\frac{1}{2} m v^2 + \frac{1}{2} k r^2$$

- Thus,

$$\begin{aligned} \langle E_{\text{atom}} \rangle &= \frac{3}{2} k_B T + \frac{3}{2} k_B T = 3k_B T \\ \langle E_{\text{solid}} \rangle &= 3N k_B T = 3n N_A k_B T = 3n R T \end{aligned}$$

- Some heat capacities are lower than $3nR$ (solids of rare gases that are heavier and need more heat to behave ideally), and some are higher (the potential is not a harmonic potential).
- As experiments got better, people realized that heat capacity, as a function of temperature, decreases as $T \rightarrow 0$ K, and was only asymptotic at $3nR$ at temperatures sufficiently close to room temperature.

- Quantum mechanics, especially the work of Einstein, solved this mystery.
- Atomic motion is quantized in units of energy.
 - If the temperature is much higher than the quantized energies, the system behaves classically.
 - If the temperature drops below the quantization energies of the vibration, we will not have equal population of energy levels (most will be in the ground state, making the energy 0; thus, there is no derivative of it and no heat capacity).
- Partition function of a quantum harmonic oscillator and the energy of the oscillator.
 - Recall that the energies are given by $(n + 1/2)h\nu$.
 - Taking our zero of energy to be $0.5 E_h$, the partition function of the vibration of the quantum harmonic oscillator is

$$\begin{aligned}
 Q &= 1 + e^{-h\nu/k_B T} + e^{-2h\nu/k_B T} + \dots \\
 Q &= (e^{-h\nu/k_B T})^0 + (e^{-h\nu/k_B T})^1 + (e^{-h\nu/k_B T})^2 + \dots \\
 Q - Qe^{-h\nu/k_B T} &= 1 \\
 Q &= \frac{1}{1 - e^{-h\nu/k_B T}}
 \end{aligned}$$

when we take the zero point energy as our zero of energy.

- It follows that

$$\begin{aligned}
 \langle E \rangle &= k_B T^2 \frac{\partial}{\partial T} \left[\ln \left(\frac{1}{1 - e^{-h\nu/k_B T}} \right) \right] \\
 &= \frac{h\nu}{e^{h\nu/k_B T} - 1}
 \end{aligned}$$

- As $T \rightarrow \infty$, $h\nu/k_B T$ gets very small. But since $e^x \approx 1 + x$ at small x , as $T \rightarrow \infty$, we have that

$$\langle E \rangle \approx \frac{h\nu}{(1 + h\nu/k_B T) - 1} = k_B T$$

- Therefore, as $T \rightarrow \infty$, we recover the energy of a classical harmonic oscillator.
- On the other hand, as $T \rightarrow 0$, $E \rightarrow 0$.
- Note that heat capacity $C = \partial E / \partial T$.

1.4 Chapter 17: The Boltzmann Factor and Partition Functions

From McQuarrie and Simon (1997).

1/23:

- Up to this point, we have established that all physical systems' energy states are quantized. Now, we address questions such as "what fraction of the molecules are to be found in the ground vibrational state, the first excited vibrational state, and so on" (McQuarrie & Simon, 1997, p. 693).
 - We will see how notions such as 'higher temperature systems should have more populated excited states' translate into precise mathematics.
 - Our two most important tools to address such questions are the **Boltzmann factor** and the **partition function**.
- **Boltzmann factor:** The relation between the probability that a system will be in a given state to the energy of that state. *Given by*

$$p_j \propto e^{-E_j/k_B T}$$

- **Partition function:** A function in terms of which we can express all of the macroscopic properties of a given system, such as energy, heat capacity, and pressure. Denoted by Q , $Q(N, V, \beta)$, $Q(N, V, T)$. Given by

$$Q = \sum_j e^{-E_j/k_B T}$$

- Determining on what quantities the energies of a macroscopic system (such as some volume of fluid, gas, or solid) depend.
 - Consider an ideal gas confined to a cubic box of side length a (we will not generalize this result explicitly even though it can be).
 - In such a system, the constituent particles do not interact, so the energy of the system will be a simple sum

$$E_j(N, V) = \epsilon_1 + \cdots + \epsilon_N$$

of the energies $\epsilon_1, \dots, \epsilon_N$ of the N particles with no higher-degree interaction terms necessary.

- Additionally, the confinement means that quantum mechanically, every particle exists in a potential of zero within the cubic box and is subject to infinite potential outside the box.
- Thus, if we consider only the translational energies of each particle, we may apply the particle in a 3D cubic box model from Chapter 3 of McQuarrie and Simon (1997) to learn that each

$$\epsilon_i = \frac{h^2}{8ma^2}(n_x^2 + n_y^2 + n_z^2)$$

- Notice that E_j depends on N and V in this system via the N terms in the summation and the dependence of each ϵ_i on $a = \sqrt[3]{V}$.
- These are the most important (and general) factors on which E_j depends, and hence we often denote the energy of the j^{th} state of the system by $E_j(N, V)$
- **Heat reservoir:** An essentially infinite heat bath.
- **Ensemble:** A huge collection of systems with identical values of N , V , and T in thermal contact with a heat reservoir at a temperature T .
 - For a given ensemble, we denote the number of systems in state j by a_j and the total number of systems by \mathcal{A} .
- Finding the relative number a_m/a_n of systems in the ensemble in states a_n and a_m .
 - a_m/a_n will depend on the energies E_n and E_m via some function f , i.e.,

$$\frac{a_m}{a_n} = f(E_n, E_m)$$

- Energies are given relative to some zero, but a_m/a_n will not depend on the arbitrary choice of this zero. Thus, the math must make said zero cancel, so we take

$$\frac{a_m}{a_n} = f(E_n - E_m)$$

- The above equation must hold for any two energy states, so we can also write $a_l/a_m = f(E_m - E_l)$ and $a_l/a_n = f(E_n - E_l)$ for instance. But this implies that

$$\frac{a_l}{a_n} = \frac{a_m}{a_n} \cdot \frac{a_l}{a_m}$$

$$f(E_n - E_l) = f(E_n - E_m)f(E_m - E_l)$$

- The above equation uniquely describes an exponential function, so we take $f(E) = e^{\beta E}$ ^[3].

■ To check that our definition of f satisfies the above equation, note that

$$e^{\beta(E_n - E_l)} = e^{\beta(E_n - E_m)} e^{\beta(E_m - E_l)}$$

- Deriving an expression for a_m .

- From the above result, we have that

$$\begin{aligned}\frac{a_m}{a_n} &= e^{\beta(E_n - E_m)} \\ a_m e^{\beta E_m} &= a_n e^{\beta E_n}\end{aligned}$$

i.e., that the value of $a_n e^{\beta E_n}$ is the same (hence constant) for any n since m is arbitrary. Thus, let $C = a_n e^{\beta E_n}$.

- It follows that

$$\begin{aligned}\frac{a_m}{a_n} &= e^{\beta(E_n - E_m)} \\ a_m &= a_n e^{\beta E_n} e^{-\beta E_m} \\ a_m &= C e^{-\beta E_m}\end{aligned}$$

- Determining C .

- We have that

$$\begin{aligned}C \sum_j e^{-\beta E_j} &= \sum_j a_j = \mathcal{A} \\ C &= \frac{\mathcal{A}}{\sum_j e^{-\beta E_j}}\end{aligned}$$

- Thus,

$$\frac{a_j}{\mathcal{A}} = \frac{1}{\sum_j e^{-\beta E_j}} e^{-\beta E_j(N, V)}$$

- Taking the limit as the number of systems in the ensemble goes to infinity makes $a_j/\mathcal{A} \rightarrow p_j$, where p_j is the *probability* that a system will be in state j (see MathChapter B).
- Recognizing that the denominator above is the partition function (and a function of N , V , and β), we have that

$$p_j(N, V, \beta) = \frac{1}{Q(N, V, \beta)} e^{\beta E_j(N, V)}$$

- We will later show that $\beta = 1/k_B T$.

- Note, however, that from a theoretical point of view, β can be just as useful as T .

- Expressions for the average energy $\langle E \rangle$ of a system.

- From the definition of $\langle E \rangle$ in MathChapter B, we have that

$$\langle E \rangle = \sum_j p_j E_j = \sum_j \frac{E_j e^{-\beta E_j}}{Q}$$

³Note that the base need not be e , but we can take it to be e WLOG since $e^{\beta E} = (e^\beta)^E$ and we may take β such that e^β equals any positive real number.

- We can also express $\langle E \rangle$ entirely in terms of Q since

$$\frac{\partial \ln Q}{\partial \beta} = - \sum_j \frac{E_j e^{-\beta E_j}}{Q} = - \langle E \rangle$$

$$\langle E \rangle = - \frac{\partial \ln Q}{\partial \beta}$$

- Substituting $\beta = 1/k_B T$ and applying the chain rule to $\partial \ln Q / \partial T$ yields

$$\frac{\partial f}{\partial T} = \frac{\partial f}{\partial \beta} \cdot \frac{\partial \beta}{\partial T} = \frac{\partial f}{\partial \beta} \cdot \frac{1}{k_B} \frac{\partial}{\partial T} \left(\frac{1}{T} \right) = \frac{\partial f}{\partial \beta} \cdot -\frac{1}{k_B T^2}$$

$$\frac{\partial f}{\partial \beta} = -k_B T^2 \frac{\partial f}{\partial T}$$

so

$$\langle E \rangle = k_B T^2 \frac{\partial \ln Q}{\partial T}$$

- Using results from Chapter 14, McQuarrie and Simon (1997) calculates the average energy of a bare proton in a magnetic field, concluding that at $T = 0$ (i.e., zero thermal energy), the proton orients itself in the direction of the magnetic field with certainty while as $T \rightarrow \infty$, the thermal energy is such that the proton becomes equally likely to be in either state.
- Calculating the average energy of a monatomic ideal gas.

- From Chapter 18,

$$Q(N, V, \beta) = \frac{[q(V, \beta)]^N}{N!} \quad q(V, \beta) = \left(\frac{2\pi m}{h^2 \beta} \right)^{3/2} V$$

- We have that

$$\begin{aligned} \ln Q &= N \ln q - \ln N! \\ &= -\frac{3N}{2} \ln \beta + \frac{3N}{2} \ln \left(\frac{2\pi m}{h^2} \right) + N \ln V - \ln N! \\ &= -\frac{3N}{2} \ln \beta + \text{terms not involving } \beta \end{aligned}$$

- Therefore,

$$\langle E \rangle = - \frac{\partial \ln Q}{\partial \beta} = \frac{3N}{2} \frac{d \ln \beta}{d \beta} = \frac{3N}{2\beta} = \frac{3}{2} N k_B T = \frac{3}{2} n R T$$

- The above result leads us to a fundamental postulate of physical chemistry: “The ensemble average of any quantity, as calculated using the probability distribution $[p_j = \frac{1}{Q} e^{-\beta E_j}]$, is the same as the experimentally observed value of that quantity” (McQuarrie & Simon, 1997, p. 700).
- The experimentally observed energy of a system is denoted by U .
- A molar quantity is denoted by an overbar (e.g., \bar{U} is the experimentally observed energy of one mole of a system).
- Calculating the average energy of a diatomic ideal gas.

- From Chapter 18,

$$Q(N, V, \beta) = \frac{[q(V, \beta)]^N}{N!} \quad q(V, \beta) = \left(\frac{2\pi m}{h^2 \beta} \right)^{3/2} V \cdot \frac{8\pi^2 I}{h^2 \beta} \cdot \frac{e^{-\beta h \nu / 2}}{1 - e^{-\beta h \nu}}$$

for the rigid rotator-harmonic oscillator model of an ideal diatomic gas.

- The first term in the expression for $q(V, \beta)$ is translational (and identical to that of a monatomic ideal gas), the second term is rotational, and the third term is vibrational.
- Using the same procedure as before, we can calculate that for one mole of a diatomic ideal gas,

$$\bar{U} = \frac{3}{2}RT + RT + \frac{N_A h \nu}{2} + \frac{N_A h \nu e^{-\beta h \nu}}{1 - e^{-\beta h \nu}}$$

- “The first term represents the average translational energy, the second term represents the average rotational energy, the third term represents the zero-point vibrational energy, and the fourth term represents the average vibrational energy” (McQuarrie & Simon, 1997, p. 701).
- Note that the fourth term becomes significant only at higher temperatures.
- **Constant-volume heat capacity:** A measure of how the energy of a system changes with temperature at constant amount and volume. Denoted by C_V . Given by

$$C_V = \frac{\partial \langle E \rangle}{\partial T} = \frac{\partial U}{\partial T}$$

- We can express C_V in terms of Q via our above expression for $\langle E \rangle$ as a function of Q .
- For an ideal monatomic gas,

$$\bar{C}_V = \frac{3}{2}R$$

- For an ideal diatomic gas,

$$\bar{C}_V = \frac{5}{2}R + R \left(\frac{h\nu}{k_B T} \right)^2 \frac{e^{-h\nu/k_B T}}{(1 - e^{-h\nu/k_B T})^2}$$

- Molar heat capacity of a crystal, as per the **Einstein model of atomic crystals**.
 - Because each lattice site is identical, assume further that all atoms vibrate with the same frequency.
 - The associated partition function is thus

$$Q = e^{-\beta U_0} \left(\frac{e^{-\beta h \nu / 2}}{1 - e^{-\beta h \nu}} \right)^{3N}$$

where ν is characteristic of the particular crystal and U_0 is the **sublimation energy** (at 0 K).

- It follows as before that

$$\bar{C}_V = 3R \left(\frac{h\nu}{k_B T} \right)^2 \frac{e^{-h\nu/k_B T}}{(1 - e^{-h\nu/k_B T})^2}$$

- **Einstein model of atomic crystals:** A model of a crystal as N atoms situated at lattice sites, with each atom vibrating as a three-dimensional harmonic oscillator.
- **Sublimation energy** (at T): The energy needed to separate all the atoms from one another at T .
- One important consequence of this result is that by experimentally measuring the heat capacity of a crystal at different temperatures, we can determine its fundamental frequency ν .
- Another is the **law of Dulong and Petit**.
- **Law of Dulong and Petit:** The molar heat capacities of atomic crystals should level off at a value of $3R = 24.9 \text{ J mol}^{-1} \text{ K}^{-1}$ at high temperatures.
- Expressions for the average pressure $\langle P \rangle$ of a system.

- From Chapter 19, the pressure of a macroscopic system in state J is

$$P_j(N, V) = -\frac{\partial E_j}{\partial V}$$

- From the definition of $\langle P \rangle$ in MathChapter B, we have that

$$\langle P \rangle = \sum_j -\frac{\partial E_j}{\partial V} \frac{e^{-\beta E_j}}{Q}$$

- We can make the above more compact since

$$\frac{\partial Q}{\partial V} = -\beta \sum_j \frac{\partial E_j}{\partial V} e^{-\beta E_j} = Q\beta \langle P \rangle$$

so

$$\langle P \rangle = \frac{1}{\beta} \frac{\partial \ln Q}{\partial V} \qquad \langle P \rangle = k_B T \frac{\partial \ln Q}{\partial V}$$

- Just like we did with energy, we equate the ensemble average pressure with the observed pressure via $P = \langle P \rangle$.
- Deriving the ideal-gas equation of state.
 - We restrict ourselves at first to the special case of a monatomic ideal gas.
 - As before, we have that

$$\begin{aligned} \ln Q &= \frac{3N}{2} \ln \left(\frac{2\pi m}{h^2 \beta} \right) + N \ln V - \ln N! \\ &= N \ln V + \text{terms not involving } V \end{aligned}$$

- Therefore,

$$\begin{aligned} \langle P \rangle &= k_B T \frac{\partial \ln Q}{\partial V} = \frac{N k_B T}{V} \\ PV &= N k_B T \end{aligned}$$

- Since only the terms not involving V change for diatomic and polyatomic gases, the above equation of state holds for all ideal gases.
- The partition function associated with the van der Waals equation is

$$Q(N, V, \beta) = \frac{1}{N!} \left(\frac{2\pi m}{h^2 \beta} \right)^{3N/2} (V - Nb)^N e^{\beta a N^2 / V}$$

- Indeed, going through the same process as above with this equation yields

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

- Since we lack the computational power to calculate the set $\{E_j\}$ of eigenvalues of the N -body Schrödinger equation, we often approximate each E_j as the sum of the energies of the constituent particles of a system.
- Consider a system of independent, *distinguishable* particles.

- A good example of one is the Einstein model of atomic crystals, since each atom is assumed to vibrate independently of the others and each atom is distinguishable by its position in the crystal lattice.
- Applying the summation approximation, we get

$$E_l(N, V) = \underbrace{\varepsilon_i^a(V) + \varepsilon_j^b(V) + \varepsilon_k^c(V) + \cdots}_{N \text{ times}}$$

where each ε_i^a denotes the energy of an individual particle (i being the energy state, and a being the index of the particle [they are distinguishable]).

- Under this approximation, the partition function of the system becomes

$$Q(N, V, T) = \sum_l e^{-\beta E_l} = \sum_{i,j,k,\dots} e^{-\beta(\varepsilon_i^a(V) + \varepsilon_j^b(V) + \varepsilon_k^c(V) + \cdots)}$$

- Since we can sum over the indices separately (i.e., one after another), the above summation can mathematically be rewritten

$$\begin{aligned} Q(N, V, T) &= \sum_i e^{-\beta \varepsilon_i^a} \sum_j e^{-\beta \varepsilon_j^b} \sum_k e^{-\beta \varepsilon_k^c} \cdots \\ &= q_a(V, T) q_b(V, T) q_c(V, T) \cdots \end{aligned}$$

where each $q(V, T)$ is a **molecular partition function**.

- **Molecular partition function:** A partition function pertaining to a particular molecule within a system. Denoted by $q(V, T)$. Given by

$$q(V, T) = \sum_j e^{-\varepsilon_j/k_B T}$$

- Often able to be evaluated since they only depend on the allowed energies of individual atoms or molecules.
- **Boson:** A particle whose wave function must be symmetric under the interchange of two identical particles.
 - Particles of integer spin (such as photons [spin 1] and deuterons [spin 0]) are bosons.
 - Two identical fermions *can* occupy the same single-particle energy state.
- **Fermions:** A particle whose wave function must be antisymmetric under the interchange of two identical particles.
 - Particles of half-integer spin (such as electrons, protons, and neutrons [all with spin 1/2]) are fermions.
 - Two identical fermions *cannot* the same single-particle energy state.
- Consider a system of independent, *indistinguishable* particles.
 - As before, we have that $E_{ijk\dots} = \varepsilon_i + \varepsilon_j + \varepsilon_k + \cdots$ for N terms, but we since we cannot distinguish between particles, we cannot sum over the indices separately.
 - Thus, our partition function is set at

$$Q(N, V, T) = \sum_{i,j,k,\dots} e^{-\beta(\varepsilon_i + \varepsilon_j + \varepsilon_k + \cdots)}$$

- If the particles in question are fermions, the indices are not independent of each other.

- In particular, we cannot have a $\epsilon_1 + \epsilon_1 + \epsilon_2 + \epsilon_3 + \dots$ term in the summation because no two fermions can occupy the same single-particle energy state.
- This restriction omits all terms with more than one particle in the same energy state from the summation.
- If the particles in question are bosons, then we must avoid summing identical terms.
 - In particular, terms such as $\epsilon_1 + \epsilon_2 + \epsilon_2 + \epsilon_2 + \dots$ and $\epsilon_2 + \epsilon_1 + \epsilon_2 + \epsilon_2 + \dots$ represent the same state, and thus should be included only once in the summation. However, an unrestricted summation would include N such terms.
 - On the other end of the spectrum, there are $N!$ terms that include $\epsilon_1 + \epsilon_2 + \epsilon_3 + \dots + \epsilon_N$ in some order.
- In either case, the terms causing the problems are those with two or more identical indices.
- If it were not for such terms, we could carry out the summation in an unrestricted manner, obtaining $[q(V, T)]^N$ as with distinguishable particles (forcing each particle to have a different state is mathematically equivalent to making them distinguishable), and then divide by $N!$, obtaining $[q(V, T)]^N / N!$, to account for the over-counting.
- But such terms do exist in the summation. However, there are times when their presence is negligible.
 - If, for example, the number of quantum states available to any particle is significantly greater than the number of particles, it is very unlikely that two particles will ever be in the same state.
 - Indeed, most quantum-mechanical systems have an infinite number of energy states. However, at any given temperature, many of these states will be energetically inaccessible.
 - Thus, we can only invoke the $[q(V, T)]^N / N!$ approximation if the number of quantum states with energies less than $k_B T$ (which is roughly the average energy of a molecule) is much larger than the number of particles.
- In particular, if

$$\frac{N}{V} \left(\frac{h^2}{8mk_B T} \right)^{3/2} \ll 1$$

then we may approximate

$$Q(N, V, T) = \frac{[q(V, T)]^N}{N!}$$

- This criterion favors large particle mass, high temperature, and low density.
- It is easily satisfied in most conventional cases.
- Quantum systems that do not satisfy this criterion must be treated by special methods beyond the scope of McQuarrie and Simon (1997).
- **Boltzmann statistics:** Statistics in which the number of available molecular states is much greater than the number of particles.
 - Favored by high temperatures.
- Relating the average energy $\langle E \rangle$ of a system to the average energy $\langle \epsilon \rangle$ of a constituent molecule.

$$\begin{aligned} \langle E \rangle &= k_B T^2 \frac{\partial \ln Q}{\partial T} \\ &= k_B T^2 \frac{\partial}{\partial T} \left(\ln \left(\frac{q^N}{N!} \right) \right) \\ &= N k_B T^2 \frac{\partial \ln q}{\partial T} \\ \langle E \rangle &= N \langle \epsilon \rangle \end{aligned}$$

- The probability that a molecule is in its j^{th} molecular energy state is denoted by π_j .
- Assuming $\varepsilon = \varepsilon_i^{\text{trans}} + \varepsilon_i^{\text{rot}} + \varepsilon_i^{\text{vib}} + \varepsilon_i^{\text{elec}}$ allows us to write $q = q_{\text{trans}}q_{\text{rot}}q_{\text{vib}}q_{\text{elec}}$ since the various energy terms are distinguishable here (hence we can sum over the indices separately).
- The probability that a molecule is in its i^{th} translational, j^{th} rotational, k^{th} vibrational, and l^{th} electronic state is given by

$$\pi_{ijkl} = \frac{e^{-\varepsilon_i^{\text{trans}}/k_B T} e^{-\varepsilon_j^{\text{rot}}/k_B T} e^{-\varepsilon_k^{\text{vib}}/k_B T} e^{-\varepsilon_l^{\text{elec}}/k_B T}}{q_{\text{trans}}q_{\text{rot}}q_{\text{vib}}q_{\text{elec}}}$$

- Since the total probability π_k^{vib} that a molecule is in its k^{th} vibrational state (for example) encompasses all probabilities of it being in any translational, rotational, or vibrational state, we have by summation that

$$\pi_k^{\text{vib}} = \sum_{i,j,l} \pi_{ijkl} = \frac{\sum_i \left(e^{-\varepsilon_i^{\text{trans}}/k_B T} \right) \sum_j \left(e^{-\varepsilon_j^{\text{rot}}/k_B T} \right) \sum_l \left(e^{-\varepsilon_l^{\text{elec}}/k_B T} \right) e^{-\varepsilon_k^{\text{vib}}/k_B T}}{q_{\text{trans}}q_{\text{rot}}q_{\text{vib}}q_{\text{elec}}} = \frac{e^{-\varepsilon_k^{\text{vib}}/k_B T}}{q_{\text{vib}}}$$

- It follows that

$$\langle \varepsilon^{\text{vib}} \rangle = \sum_k \varepsilon_k^{\text{vib}} \frac{e^{-\varepsilon_k^{\text{vib}}/k_B T}}{q_{\text{vib}}} = -\frac{\partial \ln q_{\text{vib}}}{\partial \beta} = k_B T^2 \frac{\partial \ln q_{\text{vib}}}{\partial T}$$

– Analogous results hold for $\langle \varepsilon^{\text{trans}} \rangle$, $\langle \varepsilon^{\text{rot}} \rangle$, and $\langle \varepsilon^{\text{elec}} \rangle$.

- Although we have written partition functions as sums over energy *states* up to this point, we can also sum over energy *levels* by including the degeneracy g_J of the level.

– For example, since the energy and degeneracy of a rigid rotator are, respectively,

$$\varepsilon_J = \frac{\hbar^2}{2I} J(J+1) \qquad g_J = 2J+1$$

we can write

$$q_{\text{rot}}(T) = \sum_{J=0}^{\infty} (2J+1) e^{-\hbar^2 J(J+1)/2Ik_B T}$$

Week 2

Partition Functions and Ideal Gases

2.1 System Partition Functions

1/19:

- Decomposing the partition function of a molecule into the product of separate sums as partitioned by degrees of freedom (e.g., translation, rotation, vibration, and electronic).
- The partition functions of **independent**, distinguishable/indistinguishable molecules.
 - We should not double count the same states.
 - The $N!$ in $Q = q^N/N!$ is not important when calculating energy (because of the properties of the natural log), but it is very important when calculating quantities such as entropy.
- **Independent** (particles): A set of particles that do not interact with one another.
- Discusses bosons and fermions.
 - We can have a two fermions in the state $|1, 1\rangle$ because it is a symmetric state.
- Recall the Fermi level, the boundary between the filled and unfilled electronic states in a solid.
 - If T is small, this level is a hard boundary.
 - If T is large, electrons can easily be excited and the Fermi level is a soft boundary.
- Does the 3D particle in a box derivation for the translation molecular partition function.
 - Note that since the de Broglie wavelength $\lambda_{\text{DB}} = \sqrt{h^2/2mk_B T}$, we may write

$$q_x = \sum_{n_x} e^{-h^2/8mk_B T L_x^2} = \sum_{n_x} e^{-\lambda_{\text{DB}}^2 n_x^2 / 4L_x^2}$$

- The number of states are occupied/have energy within $k_B T$ of the ground state.
 - $-\lambda_{\text{DB}}^2 n_x^2 / 4L_x^2$ is on the order of 1, implying that n_x is on the order of $2L/\lambda_{\text{DB}}$.
 - It follows if L is on a macroscopic scale (e.g., $L \approx 1$ m) and λ_{DB} is on a sub-angstrom scale that n_x is on the order of 10^{10} . When n_x is at such a scale, $e^{-\lambda_{\text{DB}}^2 n_x^2 / 4L_x^2} \approx 1/e$.
 - It follows that in a 1 m^3 box, we will have about 10^{30} states, so we really are in a regime where the number of states is larger than the number of molecules.
- More precisely, we want

$$N \ll n_x n_y n_z = \left(\frac{8mk_B T}{h^2} \right)^{3/2} L_x L_y L_z$$

where the middle term approximates the number of states so that

$$\frac{N}{V} \ll \left(\frac{8mk_B T}{h^2} \right)^{3/2}$$

- Approximating the translational energy with an integral.
 - Concludes with the translational partition function.
 - Since we can approach this problem from a classical perspective (as we did last Friday) or quantum mechanically (as we did today) to achieve the same result, this system again demonstrates the relation between quantum and classical mechanics.

2.2 Molecular Partition Functions

- 1/21: • We approximate the total molecular energy as

$$q = q_{\text{elec}} q_{\text{trans}} q_{\text{vib}} q_{\text{rot}}$$

- The heat capacity in the very high temperature limit where translations, rotations, and vibrations are classical.
 - Translational: $\frac{3}{2}k_B$.
 - Vibrational: Each degree of freedom ($3N - 5$ for a linear molecule and $3N - 6$ for a nonlinear molecule) contributes k_B .
 - Rotational: Each degree of freedom (2 for a linear molecule and 3 for a nonlinear molecule) contributes $\frac{1}{2}k_B$.
- We can use the above to calculate the heat capacity of various molecules at very high temperatures (note, however, that at such temperatures, molecules would likely dissociate; we're simply theoretically considering the classical limit here).
 - Ne: $\frac{3}{2}k_B$.
 - H₂O: $\frac{3}{2}k_B + 3 \cdot k_B + 3 \cdot \frac{1}{2}k_B = 6k_B$.
 - O₂: $\frac{3}{2} + 1 \cdot k_B + 2 \cdot \frac{1}{2}k_B = \frac{7}{2}k_B$.
 - CO₂: $\frac{3}{2}k_B + 4 \cdot k_B + 2 \cdot \frac{1}{2}k_B = \frac{13}{2}k_B$.
 - CHCl₃: $\frac{3}{2}k_B + 9 \cdot k_B + 3 \cdot \frac{1}{2}k_B = 12k_B$.
- Electronic partition function.
 - Consider the bottom D_e of the potential well of a diatomic.
 - D_0 is the ionization energy from the bottom state ($D_e \neq D_0$, but relations can be obtained via spectroscopy).
 - It follows that

$$q_{\text{elec}} = g_1 e^{-(D_e/k_B T)} + g_2 e^{-E_2/k_B T}$$
 - If $dT \ll (E_2 + D_e)$, then $q_{\text{elec}} = g_1 e^{D_e/k_B T}$.
- Vibrational partition function.
 - As before with the law of Dulong and Petit.
 - It's a special point where $T = h\nu/k_B$.
- Rotational partition function.
 - Almost always classical.
 - The rotational energy of a polyatomic molecule will almost always be $\frac{3}{2}k_B$.
 - Let's look at a heteronuclear diatomic, such as CO. Derives

$$q_{\text{rot}} = \sum_{J=0}^{\infty} (2J+1) e^{-\hbar^2 J(J+1)/2Ik_B T}$$

- The **rotational temperature** leads to

$$q_{\text{rot}} = \sum_{J=0}^{\infty} (2J+1) e^{\Theta_{\text{rot}}/T} = \frac{T}{\Theta_{\text{rot}}}$$

- Thus, at the temperature at which we exist, rotation is equivalent classically to quantum mechanically.

- **Rotational temperature:** The following quantity. Denoted by Θ_{rot} . Given by

$$\Theta_{\text{rot}} = \hbar^2 / 2Ik_B$$

- PGS will not specify whether we need a quantum vs. classical model.

- Homonuclear diatomic (e.g., H_2).

- The vibrational differences in energy become visible with spectroscopy.

- $q_{\text{rot}} = T/2\Theta_{\text{rot}}$.

- Partition functions:

- If the molecule is linear, it's of the form T/Θ_{rot} .

- If the molecule is nonlinear, it's of the form $T/2\Theta_{\text{rot}}$.

- Spherical top (e.g., CH_4):

$$\frac{\sqrt{\pi}}{\sigma} \left(\frac{T}{\Theta_{\text{rot}}} \right)^{3/2}$$

- Symmetric top (e.g., NH_3):

$$\frac{\sqrt{\pi}}{\sigma} \sqrt{\frac{T^3}{\Theta_{\text{rot},a}^2 \Theta_{\text{rot},b}}}$$

- a and b are the two different symmetry axes.

- Asymmetric top (e.g., H_2O):

$$\frac{\sqrt{\pi}}{\sigma} \sqrt{\frac{T^3}{\Theta_{\text{rot},a} \Theta_{\text{rot},b} \Theta_{\text{rot},c}}}$$

- Application to total energy and heat capacity of a molecule.

- We have that

$$q = \left(\frac{2\pi m k_B T}{h^2} \right)^{3/2} \cdot V \cdot \sqrt{\frac{T^2}{\sigma \Theta_{\text{rot},a} \Theta_{\text{rot},b} \Theta_{\text{rot},c}}} \cdot \sum_1^{3N-6} \frac{e^{-\Theta_{\text{vib}}/2T}}{1 - e^{-\Theta_{\text{vib}}/T}} \cdot g_1 e^{D_e/k_B T}$$

- Thus,

$$\langle E \rangle = k_B T^2 \frac{\partial \ln q}{\partial T} = k_B T^2 \frac{\partial}{\partial T} \left(\frac{3}{2} \ln T + \text{constant} + \frac{3}{2} \ln T + \text{vibration} - D_e \right)$$

- The energy of the vibration is $E = k_B \Theta_{\text{vib}} / (e^{\Theta_{\text{vib}}/k_B T} - 1) + k_B \Theta_{\text{vib}}/2$. It follows that

$$C = \frac{\partial E}{\partial T} = k_B \frac{\Theta_{\text{vib}}^2}{T^2} \frac{e^{-\Theta_{\text{vib}}/T}}{(1 - e^{-\Theta_{\text{vib}}/T})}$$

2.3 Chapter 18: Partition Functions and Ideal Gases

From McQuarrie and Simon (1997).

- 1/23:
- Herein, we will calculate the partition functions and heat capacities of ideal gases.
 - We heavily rely on the expression of the partition function for a system of independent, indistinguishable particles, which ideal gases are likely to satisfy because of their low density.
 - Deriving the translational molecular partition function of an atom in a monatomic ideal gas.
 - As mentioned in Chapter 17, if we let the container be cubic, then

$$\varepsilon(n_x, n_y, n_z) = \frac{h^2}{8ma^2}(n_x^2 + n_y^2 + n_z^2)$$

- It follows that

$$\begin{aligned} q_{\text{trans}} &= \sum_{n_x, n_y, n_z=1}^{\infty} e^{-\beta \varepsilon(n_x, n_y, n_z)} \\ &= \sum_{n_x=1}^{\infty} \exp\left(-\frac{\beta h^2 n_x^2}{8ma^2}\right) \sum_{n_y=1}^{\infty} \exp\left(-\frac{\beta h^2 n_y^2}{8ma^2}\right) \sum_{n_z=1}^{\infty} \exp\left(-\frac{\beta h^2 n_z^2}{8ma^2}\right) \\ &= \left[\sum_{n=1}^{\infty} \exp\left(-\frac{\beta h^2 n^2}{8ma^2}\right) \right]^3 \end{aligned}$$

- The above sum cannot be evaluated in closed form. However, since later terms in the summation get very small, it is an excellent approximation to replace the summation with an integral, i.e.,

$$\begin{aligned} q_{\text{trans}} &= \left(\int_0^{\infty} e^{-\beta h^2 n^2 / 8ma^2} dn \right)^3 \\ &= \left(\sqrt{\frac{\pi}{4\beta h^2 / 8ma^2}} \right)^3 \\ &= \left(\sqrt{\frac{2\pi m}{\beta h^2}} \right)^3 a^3 \\ &= \left(\frac{2\pi m k_B T}{h^2} \right)^{3/2} V \end{aligned}$$

- 1/24:
- Deriving the electronic molecular partition function of an atom in a monatomic ideal gas.
 - We express the partition function here in terms of levels, i.e., by

$$q_{\text{elec}} = \sum_i g_{ei} e^{-\beta \varepsilon_{ei}}$$

where g_{ei} is the degeneracy and ε_{ei} is the energy of the i^{th} electronic level.

- Taking $\varepsilon_{e1} = 0$ to be the zero of energy yields

$$q_{\text{elec}} = g_{e1} + g_{e2} e^{-\beta \varepsilon_{e2}} + \dots$$

- Note that since ε 's are usually on the order of tens of thousands of wavenumbers, $e^{-\beta \varepsilon_{e2}}$ is around 10^{-5} for most atoms at ordinary temperatures, so only the first term in the summation is significantly different from zero.
- For some gases such as halogens, other terms may be important, but even there the sum converges very rapidly.

- Using spectroscopic data, we can show that the fraction of atoms of most gases in the first excited state is very small.
 - For example, the fraction of helium atoms at 300 K in the first excited state is 10^{-334} .
 - For fluorine, however, the fraction is on the order of 10^{-2} , which is significant. In this case, we need to approximate q_{elec} with more than one or two terms.
- McQuarrie and Simon (1997) recalculates the average energy, heat capacity, and pressure of a monatomic ideal gas using the above results.
- Diatomics.
 - The translational partition function is

$$q_{\text{trans}}(V, T) = \left[\frac{2\pi(m_1 + m_2)k_B T}{h^2} \right]^{3/2} V$$

- We take the zero of rotational energy to be the $J = 0$ state.
- We take the zero of vibrational energy to be the bottom of the internuclear potential well of the lowest electronic state (so that the energy of the ground vibrational state is $h\nu/2$).
- We take the zero of electronic energy to be the energy of the separated atoms at rest in their ground electronic state (so that the energy of the ground electronic state is $-D_e^{[1]}$).
- **Vibrational temperature:** The following quantity. Denoted by Θ_{vib} . Given by

$$\Theta_{\text{vib}} = \frac{h\nu}{k_B}$$

- Deriving the vibrational molecular partition function of a molecule in a diatomic ideal gas.

$$\begin{aligned} q_{\text{vib}}(T) &= \sum_{v=0}^{\infty} e^{-\beta(v+1/2)h\nu} \\ &= e^{-\beta h\nu/2} \sum_{v=0}^{\infty} e^{-\beta h\nu v} \\ &= e^{-\beta h\nu/2} \frac{1}{1 - e^{-\beta h\nu}} \\ &= \frac{e^{-\beta h\nu/2}}{1 - e^{-\beta h\nu}} \end{aligned}$$

- In terms of Θ_{vib} ,

$$\begin{aligned} q_{\text{vib}}(T) &= \frac{e^{-\Theta_{\text{vib}}/2T}}{1 - e^{-\Theta_{\text{vib}}/T}} \\ \langle E_{\text{vib}} \rangle &= Nk_B \left(\frac{\Theta_{\text{vib}}}{2} + \frac{\Theta_{\text{vib}}}{e^{\Theta_{\text{vib}}/T} - 1} \right) \\ \bar{C}_{\text{V,vib}} &= R \left(\frac{\Theta_{\text{vib}}}{T} \right)^2 \frac{e^{-\Theta_{\text{vib}}/T}}{(1 - e^{-\Theta_{\text{vib}}/T})^2} \end{aligned}$$

- Note that the high temperature limit of $\bar{C}_{\text{V,vib}}$ is R , and $\bar{C}_{\text{V,vib}}$ attains $R/2$ at $T = 0.34 \Theta_{\text{vib}}$.
- Calculating the fraction of molecules in the ground vibrational state reveals that generally, most molecules are in the ground vibrational state.

¹See Figure 9.7 of Labalme (2021c).

- Exceptions include Br_2 , the smaller force constant and larger mass of which lead to a smaller value of Θ_{vib} .

- **Rotational temperature:** The following quantity. Denoted by Θ_{rot} . Given by

$$\Theta_{\text{rot}} = \frac{\hbar^2}{2Ik_B} = \frac{hB}{k_B}$$

- B is the rotational constant (see Chapter 5) in the above equation.

- Deriving the rotational molecular partition function of a *heteronuclear* molecule in a diatomic ideal gas.

- We have

$$q_{\text{rot}}(T) = \sum_{J=0}^{\infty} (2J+1)e^{-\Theta_{\text{rot}}J(J+1)/T}$$

- As with the translational partition function, for $\Theta_{\text{rot}} \ll T$ (which is true for normal temperatures), we may approximate the above sum via an integral. This approximation is known as the high-temperature limit, and under it,

$$\begin{aligned} q_{\text{rot}}(T) &= \int_0^{\infty} (2J+1)e^{-\Theta_{\text{rot}}J(J+1)/T} dJ \\ &= \int_0^{\infty} e^{-\Theta_{\text{rot}}x/T} dx \\ &= \frac{T}{\Theta_{\text{rot}}} = \frac{8\pi^2 Ik_B T}{h^2} \end{aligned}$$

- For low temperatures or molecules with large values of Θ_{rot} we evaluate some number of terms of the sum directly, but we will not consider these cases further.

- It follows from the above that

$$\langle E_{\text{rot}} \rangle = Nk_B T \quad \quad \quad \overline{C}_{V,\text{rot}} = R$$

- Each of the two rotational degrees of freedom of a diatomic contributes $R/2$ to $\overline{C}_{V,\text{rot}}$.

- Contrary to the other component parts of energy, higher energy rotational states are significantly occupied.

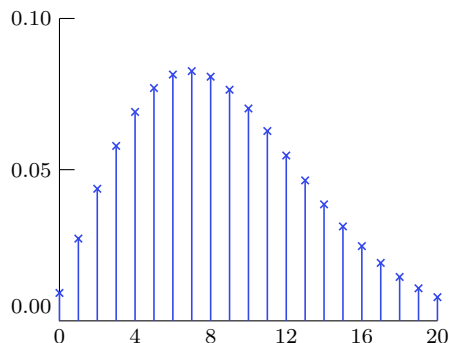


Figure 2.1: The fraction of molecules in the J^{th} rotational level for CO at 300 K.

- We have that the fraction f_J of molecules in the J^{th} vibrational state is

$$f_J = \frac{(2J+1)e^{-\Theta_{\text{rot}}J(J+1)/T}}{q_{\text{rot}}} = (2J+1)\left(\frac{\Theta_{\text{rot}}}{T}\right)e^{-\Theta_{\text{rot}}J(J+1)/T}$$

- We can estimate the most probable value of J by solving $\partial f_J / \partial J = 0$, which gives $J = 7$ in agreement with Figure 2.1.

1/25:

- We now address the rotational molecular partition function for a *homonuclear* diatomic ideal gas molecule.
 - Because of the additional perpendicular C_2 axes of symmetry in a homonuclear diatomic compared to a heteronuclear diatomic, the diatomic's constituent atoms are 'more' indistinguishable, i.e., only nuclear spin can distinguish them.
 - "In particular, if the two nuclei have integral spins (bosons), the molecular wave function must be symmetric with respect to an interchange of the two nuclei; if the nuclei have half odd integer spin (fermions), the molecular wave function must be antisymmetric" (McQuarrie & Simon, 1997, p. 747).
 - This symmetry affects the population of the rotational energy levels in a way that *can* be derived, but we will just state the important result, which is that for a homonuclear diatomic molecule,

$$q_{\text{rot}}(T) = \frac{T}{2\Theta_{\text{rot}}}$$

- To unify the two rotational molecular partition functions, we let

$$q_{\text{rot}}(T) = \frac{T}{\sigma\Theta_{\text{rot}}}$$

in general, where σ is the **symmetry number**.

- **Symmetry number:** The number of different ways a given molecule can be rotated into a configuration indistinguishable from the original. *Denoted by σ . Given by*

$$\sigma = \begin{cases} 1 & \text{heteronuclear} \\ 2 & \text{homonuclear} \end{cases}$$

- Taking the energy of an ideal diatomic gas molecule to be a simple sum of its translational, rotational, vibrational, and electronic energies yields the molecular partition function

$$q(V, T) = \left(\frac{2\pi M k_B T}{h^2} \right)^{3/2} V \cdot \frac{T}{\sigma\Theta_{\text{rot}}} \cdot \frac{e^{-\Theta_{\text{vib}}/2T}}{1 - e^{-\Theta_{\text{vib}}/T}} \cdot g_e e^{D_e/k_B T}$$

where we require $\Theta_{\text{rot}} \ll T$, that the only populated electronic state is the ground state, that the zero of electronic energy is the separated atoms at rest in their ground electronic states, and that the zero of vibrational energy is the bottom of the internuclear potential well of the lowest electronic state.

- McQuarrie and Simon (1997) derives the molar energy and heat capacity of a diatomic ideal gas one more time using the above equation.
 - The only difference is that the newly added electronic factor in the partition function adds a term of $-N_A D_e$ to the Chapter 17 formula for \bar{U} .
 - Also note that we can greatly improve the agreement of the harmonic oscillator-rigid rotator model with even small first-order corrections, such as including centrifugal distortion and anharmonicity.
- The translational and electronic molecular partition functions of an ideal polyatomic molecule are the same as those of an ideal monatomic or diatomic molecule.
- On the vibrational molecular partition function of an ideal polyatomic molecule.
 - Recall from Chapter 13 that the vibrational motion of a polyatomic molecule can be expressed in terms of normal coordinates.

- Thus, the vibrational energy of a polyatomic molecule in state $v_j = 0, 1, 2, \dots$ is

$$\varepsilon_{\text{vib}} = \sum_{j=1}^{N_{\text{vib}}} \left(v_j + \frac{1}{2}\right) h\nu_j$$

where ν_j is the frequency of the j^{th} normal mode.

- It follows that for a polyatomic molecule,

$$q_{\text{vib}} = \prod_{j=1}^{N_{\text{vib}}} \frac{e^{-\Theta_{\text{vib},j}/2T}}{1 - e^{-\Theta_{\text{vib},j}/T}}$$

$$E_{\text{vib}} = Nk_B \sum_{j=1}^{N_{\text{vib}}} \left(\frac{\Theta_{\text{vib},j}}{2} + \frac{\Theta_{\text{vib},j} e^{-\Theta_{\text{vib},j}/T}}{1 - e^{-\Theta_{\text{vib},j}/T}} \right)$$

$$C_{V,\text{vib}} = Nk_B \sum_{j=1}^{N_{\text{vib}}} \left[\left(\frac{\Theta_{\text{vib},j}}{T} \right)^2 \frac{e^{-\Theta_{\text{vib},j}/T}}{(1 - e^{-\Theta_{\text{vib},j}/T})^2} \right]$$

- Rotational molecular partition functions for linear molecules.

- We can still apply the rigid-rotator approximation, but with

$$I = \sum_{j=1}^n m_j d_j^2$$

where d_j is the distance of the j^{th} nucleus from the center of mass of the molecule.

- Doing so yields

$$q_{\text{rot}}(T) = \frac{T}{\sigma \Theta_{\text{rot}}}$$

where $\sigma = 1$ for unsymmetrical molecules such as N_2O and COS and $\sigma = 2$ for symmetrical molecules such as CO_2 and C_2H_2 .

- Note that the symmetry number of NH_3 is three.

- Rotational molecular partition functions for nonlinear molecules.

- Recall the discussion surrounding the principal moments of inertia in Chapter 13.
- We define three characteristic rotational temperatures, namely $\Theta_{\text{rot},j} = \hbar^2/2I_j k_B$ for $j = A, B, C$.
- Spherical top.
 - In this case, $\Theta_{\text{rot},A} = \Theta_{\text{rot},B} = \Theta_{\text{rot},C} = \Theta_{\text{rot}}$.
 - The quantum-mechanical spherical top can be solved exactly to give

$$\varepsilon_J = \frac{\hbar^2}{2I} J(J+1) \qquad g_J = (2J+1)^2$$

- Now $\Theta_{\text{rot}} \ll T$ for almost all spherical top molecules at ordinary temperatures, and this has two important consequence. First, we can approximate the partition function with an integral. Second, we can neglect 1 in comparison with J since the important values of J are large. Thus, we have that

$$q_{\text{rot}}(T) = \frac{1}{\sigma} \sum_{J=0}^{\infty} (2J+1)^2 e^{-\hbar^2 J(J+1)/2Ik_B T}$$

$$= \frac{1}{\sigma} \int_0^{\infty} (2J+1)^2 e^{-\Theta_{\text{rot}} J(J+1)/T} dJ$$

$$\begin{aligned}
&= \frac{1}{\sigma} \int_0^\infty 4J^2 e^{-\Theta_{\text{rot}} J^2 / T} dJ \\
&= \frac{4}{\sigma} \int_0^\infty J^2 e^{-a J^2} dJ \\
&= \frac{4}{\sigma} \cdot \frac{1}{4a} \sqrt{\frac{\pi}{a}} \\
q_{\text{rot}}(T) &= \frac{\sqrt{\pi}}{\sigma} \left(\frac{T}{\Theta_{\text{rot}}} \right)^{3/2}
\end{aligned}$$

– Similarly, we have respectively for a symmetric top and an asymmetric top that

$$q_{\text{rot}}(T) = \frac{\sqrt{\pi}}{\sigma} \left(\frac{T}{\Theta_{\text{rot},A}} \right) \sqrt{\frac{T}{\Theta_{\text{rot},C}}} \quad q_{\text{rot}}(T) = \frac{\sqrt{\pi}}{\sigma} \sqrt{\frac{T}{\Theta_{\text{rot},A} \Theta_{\text{rot},B} \Theta_{\text{rot},C}}}$$

- It follows that

$$\bar{U}_{\text{rot}} = \frac{3RT}{2} \quad \bar{C}_{V,\text{rot}} = \frac{3R}{2}$$

- Linear molecule equations.

$$\begin{aligned}
q(V, T) &= \left(\frac{2\pi M k_B T}{h^2} \right)^{3/2} V \cdot \frac{T}{\sigma \Theta_{\text{rot}}} \cdot \prod_{j=1}^{3n-5} \frac{e^{-\Theta_{\text{vib},j}/2T}}{1 - e^{-\Theta_{\text{vib},j}/T}} \cdot g_{e1} e^{D_e/k_B T} \\
\frac{U}{N k_B T} &= \frac{3}{2} + \frac{2}{2} + \sum_{j=1}^{3n-5} \left(\frac{\Theta_{\text{vib},j}}{2T} + \frac{\Theta_{\text{vib},j}/T}{e^{\Theta_{\text{vib},j}/T} - 1} \right) - \frac{D_e}{k_B T} \\
\frac{C_V}{N k_B} &= \frac{3}{2} + \frac{2}{2} + \sum_{j=1}^{3n-5} \left(\frac{\Theta_{\text{vib},j}}{T} \right)^2 \frac{e^{-\Theta_{\text{vib},j}/T}}{(1 - e^{-\Theta_{\text{vib},j}/T})^2}
\end{aligned}$$

- Nonlinear molecule equations.

$$\begin{aligned}
q(V, T) &= \left(\frac{2\pi M k_B T}{h^2} \right)^{3/2} V \cdot \frac{\sqrt{\pi}}{\sigma} \sqrt{\frac{T^3}{\Theta_{\text{rot},A} \Theta_{\text{rot},B} \Theta_{\text{rot},C}}} \cdot \prod_{j=1}^{3n-6} \frac{e^{-\Theta_{\text{vib},j}/2T}}{1 - e^{-\Theta_{\text{vib},j}/T}} \cdot g_{e1} e^{D_e/k_B T} \\
\frac{U}{N k_B T} &= \frac{3}{2} + \frac{3}{2} + \sum_{j=1}^{3n-6} \left(\frac{\Theta_{\text{vib},j}}{2T} + \frac{\Theta_{\text{vib},j}/T}{e^{\Theta_{\text{vib},j}/T} - 1} \right) - \frac{D_e}{k_B T} \\
\frac{C_V}{N k_B} &= \frac{3}{2} + \frac{3}{2} + \sum_{j=1}^{3n-6} \left(\frac{\Theta_{\text{vib},j}}{T} \right)^2 \frac{e^{-\Theta_{\text{vib},j}/T}}{(1 - e^{-\Theta_{\text{vib},j}/T})^2}
\end{aligned}$$

Part II

Thermodynamics

Week 3

Kinetic Theory of Gases / The First Law of Thermodynamics

3.1 Maxwell-Boltzmann Distribution

- 1/24:
- Applying the molecular partition function to the heat capacity of a water molecule.
 - A water molecule has three vibrational modes, which we will denote by ν_1, ν_2, ν_3 (corresponding to symmetric stretch, antisymmetric stretch, and bend).
 - Main takeaway: Heat capacity can change with temperature.
 - After a while (at several thousand kelvin), it will level off (see Figure 18.7).
 - Considers CO₂'s vibrational modes, too.
 - The infrared absorption of the bending mode is what's associated with the Greenhouse Effect.
 - The symmetric stretch is IR inactive due to its lack of change of dipole moment.
 - Raman active: Change in the polarizability of the molecule.
 - The Maxwell-Boltzmann distribution.
 - Maxwell derived it long before Boltzmann, but Boltzmann's thermodynamic derivation is much easier.
 - We know from the Boltzmann factor that $p(E) \propto e^{-E/k_B T}$.
 - Thus, to get the probability $p(v)$ of some speed v , we should have $p(v) \propto e^{-mv^2/2k_B T}$ times a constant giving the number of molecules of each speed? This yields

$$p(v) = A4\pi v^2 e^{-mv^2/2k_B T}$$

where A is a normalization constant.

- The Maxwell-Boltzmann distribution is such that

$$\begin{aligned} 1 &= \int_0^\infty p(v) \, dv \\ &= A \int_0^\infty 4\pi v^2 e^{-mv^2/2k_B T} \, dv \\ &= A \int_0^\infty 4\pi \left(\frac{2k_B T}{m} \right)^{3/2} u^2 e^{-u^2} \, du \\ &= A4\pi \left(\frac{2k_B T}{m} \right)^{3/2} \int_0^\infty u^2 e^{-u^2} \, du \end{aligned}$$

$$= A4\pi \left(\frac{2k_B T}{m} \right)^{3/2} \frac{\sqrt{\pi}}{4}$$

$$A = \left(\frac{m}{2\pi k_B T} \right)^{3/2}$$

– Therefore,

$$p(v) = 4\pi \left(\frac{m}{2\pi k_B T} \right)^{3/2} v^2 e^{-mv^2/2k_B T}$$

– Any distribution that doesn't look like this isn't in thermal equilibrium.

- A system with all particles having $v = 0$ is at thermal equilibrium with $T = 0$ K.
- A system with all particles having constant velocity in the same direction is at thermal equilibrium with $T = 0$ K.
 - Think relativity; if you're moving with them, it looks like they're not moving and thus this case is the same as the last one because your movement doesn't affect the thermodynamics of that system.
- A system with all particles having constant velocity in different directions is not at thermal equilibrium since it does not fit the bell curve but is rather a spike.

3.2 The First Law of Thermodynamics

1/26:

- See Labalme (2021d) for background on/content of today's lecture.
- Joule best quantified how we think about work and energy.
- **System:** Part of the world being investigated. It can contain energy, a number of particles, etc.
- The Newtonian way to change the energy into a system is to do work (mechanical, electrical, etc.) on the system. In chemistry, δw is positive if work is done on the system.
 - We have that

$$\delta w = -P dV$$

- A system at thermal equilibrium has a given temperature characteristic of the system. Some property of the system indicates how hot or cold it is (e.g., volume of mercury, etc.)
- Measure heat transfer using a calorimeter and a thermometer.
 - Convention: Heat put into a system δq is positive and this heat is transferred if the temperature is lower than another system or surroundings.
 - The heat capacity of a system times the change in temperature is equal to the heat put into the system. It is always positive as heat put into the system raises the temperature.
- Example molar heat capacities.
 - For water vapor at low pressure and 20 °C, $\overline{C}_V = 3R = 25 \text{ J mol}^{-1} \text{ K}^{-1}$.
 - For liquid water, it's higher (hydrogen bonding).
 - For ice, it's lower.
- **First law of thermodynamics:** The internal energy of a system changes with heat put into the system and work done on the system.

$$dU = \delta q + \delta w$$

– Note that in engineering, $dU = \delta q - \delta w$.

- **State variable:** A property that describes the system.
 - For example, a system of gas molecules has a state defined by the state variables T , P , V , and n .
- **State function:** A property that depends only upon the state of the system.
 - For example, some equations of state for an ideal gas are $PV = nRT$ or $PV = 2U/3$.
 - The internal energy is a state function.
 - Heat and work are not state functions because they do not depend uniquely on the values at equilibrium.
 - They also depend on the way you do something.
- **Reversible process:** A process that can be represented as a path along state variables, e.g., a line on a PV diagram. This implies that it is also a path where all state variables are known, and is therefore a path where the system is always in quasi-equilibrium.
 - Isothermal, isochoric, isobaric, and adiabatic changes are reversible.
 - All of these processes are analyzed exactly as in Labalme (2021d).
- **Irreversible process:** A process that cannot be drawn on a PV diagram.
- Experiment to measure γ (the ratio of specific heats):
 1. Let sit at $P_0 T_0$.
 2. Pump in a little gas (add Δn) and let sit, measure $P_0 + \Delta P_1, T_0$.
 3. Open the valve to air quickly to P_0 . Adiabatic expansion (cools down).
 4. Let sit to measure the new pressure $P + \Delta P_2$ when T is back at T_0 .
 5. γ is determined from $\Delta P_1, \Delta P_2$ (this will be a homework problem).
 - In the second step, we add some molecules into the container. We can show that $\Delta P_1/P_0 = \Delta N_1/n_0$.
 - In the third step, we let out the air, and we can show that $\Delta n_2/n_0 = \gamma \Delta P_1/P_0$.
 - In step 4, we have in the container $(n_0 + \Delta n_1 - \Delta n_2)RT_0 = (P_0 + \Delta P_2)V_0$.
 - This implies that $\Delta P_2/\Delta P_1 = 1 - 1/\gamma$.

3.3 Enthalpy

- 1/28:
- Thermodynamic derivation of the formula for $\langle P \rangle$ in terms of Q .
 - We have that

$$\begin{aligned}
 U &= \sum p_j E_j \\
 dU &= \sum (dp_j E_j + p_j dE_j) \\
 &= \underbrace{\sum dp_j E_j}_{\delta q} + \underbrace{\sum p_j \frac{\partial E_j}{\partial V} dV}_{-P}
 \end{aligned}$$

where the last part follows by analogy with $dU = \delta q - P dV$.

- It follows that

$$P = - \sum p_j \frac{\partial E_j}{\partial V} = - \left\langle \frac{\partial E}{\partial V} \right\rangle$$

- Thus, we have that

$$\begin{aligned}
 P &= - \sum \frac{e^{-E_j/k_B T}}{Q} \frac{\partial E_j}{\partial V} \\
 &= \frac{1}{Q} \sum k_B T \cdot - \frac{1}{k_B T} e^{-E_j/k_B T} \frac{\partial E_j}{\partial V} \\
 &= k_B T \frac{1}{Q} \sum \frac{\partial}{\partial E_j} \left(e^{-E_j/k_B T} \right) \frac{\partial E_j}{\partial V} \\
 &= k_B T \frac{1}{Q} \sum \frac{\partial}{\partial V} \left(e^{-E_j/k_B T} \right) \\
 &= k_B T \frac{1}{Q} \frac{\partial Q}{\partial V} \\
 P &= k_B T \frac{\partial \ln Q}{\partial V}
 \end{aligned}$$

- Applies the formula to an ideal gas of independent, indistinguishable particles to derive the ideal gas law.
- **Enthalpy:** A state function representign the heat put into the system at constant pressure. *Denoted by H . Given by*

$$H = U + PV$$

- We have that

$$\begin{aligned}
 dH &= dU + P dV + V dP \\
 &= \delta q - P dV + P dV + V dP \\
 &= \delta q + V dP
 \end{aligned}$$

- At constant pressure ($dP = 0$), we have that $dH = \delta q$.
- At constant volume, we have that $dH = \delta q$ as well?

- **Constant-volume heat capacity.** The following expression. *Denoted by C_V . Given by*

$$C_V = \left(\frac{\partial U}{\partial T} \right)_{V,N}$$

- **Constant-pressure heat capacity.** The following expression. *Denoted by C_P . Given by*

$$C_P = \left(\frac{\partial H}{\partial T} \right)_{P,N}$$

- For an ideal gas,

$$\begin{aligned}
 dH &= dU + d(PV) \\
 &= nC_V dT + nR dT \\
 &= n(C_V + R) dT
 \end{aligned}$$

- Recall this result from Labalme (2021d).

- Considers heat diagrams.

- Recall the enthalpy of phase changes ΔH_{fus} , ΔH_{vap} , and ΔH_{sub} .
- It follows that

$$H(T) - H(T_0) = \int_{T_0}^T C_p dT + \sum \Delta H_{\text{phase changes}}$$

- **Hess's Law:** $\Delta H = 0$ around a closed loop.
 - This is because H is a state function.
- **Standard enthalpy of formation.** The enthalpy of formation of a substance, as calculated from its constituent elements in their standard state (i.e., 1 bar, 298.15 K). Denoted by ΔH_f° . Units **kJ/mol**.
- We have, for example, that the $\Delta H_{\text{vap}}^\circ$ of a substance is the difference of its ΔH_f° in its gaseous state and its ΔH_f° in its liquid state.
- With the standard enthalpy of formation and the heat capacity $C_P(T)$, one gets the enthalpy of formation at nonstandard temperatures.
- To get the enthalpy of formation at non-standard pressures of chemical interest, most of the effect is from the gas components because solids and liquid enthalpy vary little with pressure.
- The direction of change is sometimes in the direction of *positive* enthalpy change.
 - This change is driven by the fact that in these cases, the direction of change is toward the most probable state.
- In a reversible process, $dU = \delta q_{\text{rev}} - P dV$. In this case

$$\delta q_{\text{rev}} = dU + P dV = nC_V dT + P dV \neq d(nC_V T + PV)$$

so δq_{rev} is not a state function.

- However,

$$\begin{aligned} \frac{\delta q_{\text{rev}}}{T} &= nC_V \frac{dT}{T} + \frac{P dV}{T} \\ &= nC_V \frac{dT}{T} + nR \frac{dV}{V} \\ &= d(nC_V \ln T + nR \ln V) \end{aligned}$$

is a state function.

3.4 Chapter 25: The Kinetic Theory of Gases

From McQuarrie and Simon (1997).

1/30:

- **Kinetic theory of gases:** A simple model of gases in which the molecules (pictured as hard spheres) are assumed to be in constant, incessant motion, colliding with each other and with the walls of the container.
- McQuarrie and Simon (1997) does the KMT derivation of the ideal gas law from Labalme (2021a). Some important notes follow.
 - McQuarrie and Simon (1997) emphasizes the importance of

$$PV = \frac{1}{3}Nm \langle u^2 \rangle$$

as a fundamental equation of KMT, as it relates a macroscopic property PV to a microscopic property $m \langle u^2 \rangle$.

- In Chapter 17-18, we derived quantum mechanically, and then from the partition function, that the average translational energy $\langle E_{\text{trans}} \rangle$ for a single particle of an ideal gas is $\frac{3}{2}k_B T$. From classical mechanics, we also have that $\langle E_{\text{trans}} \rangle = \frac{1}{2}m \langle u^2 \rangle$. *This* is why we may let

$$\frac{1}{2}m \langle u^2 \rangle = \frac{3}{2}k_B T$$

recovering that the average translational kinetic energy of the molecules in a gas is directly proportional to the Kelvin temperature.

- **Isotropic** (entity): An object or substance that has the same properties in any direction.
 - For example, a homogeneous gas is isotropic, and this is what allows us to state that $\langle u_x^2 \rangle = \langle u_y^2 \rangle = \langle u_z^2 \rangle$.
- McQuarrie and Simon (1997) derives

$$u_{\text{rms}} = \sqrt{\frac{3RT}{M}}$$

- u_{rms} is an estimate of the average speed since $\langle u^2 \rangle \neq \langle u \rangle^2$ in general.
- McQuarrie and Simon (1997) states without proof that the speed of sound u_{sound} in a monatomic ideal gas is given by

$$u_{\text{sound}} = \sqrt{\frac{5RT}{3M}}$$

- Assumptions of the kinetic theory of gases.
 - Particles collide elastically with the wall.
 - Justified because although each collision will not be elastic (the particles in the wall are moving too), the average collision will be elastic.
 - Particles do not collide with each other.
 - Justified because “if the gas is in equilibrium, on the average, any collision that deflects the path of a molecule... will be balanced by a collision that replaces the molecule” (McQuarrie & Simon, 1997, p. 1015).
- Note that we can do the kinetic derivation at many levels of rigor, but more rigorous derivations offer results that differ only by constant factors on the order of unity.
- Deriving a theoretical equation for the distribution of the *components* of molecular velocities.
 - Let $h(u_x, u_y, u_z) du_x du_y du_z$ be the fraction of molecules with velocity components between u_j and $u_j + du_j$ for $j = x, y, z$.
 - Assume that the each component of the velocity of a molecule is independent of the values of the other two components¹. It follows statistically that

$$h(u_x, u_y, u_z) = f(u_x)f(u_y)f(u_z)$$

- Note that we use just one function f for the probability distribution in each direction because the gas is isotropic.
- We can use the isotropic condition to an even greater degree. Indeed, it implies that any information conveyed by u_x is necessarily and sufficiently conveyed by u_y , u_z , and u . Thus, we may take

$$h(u) = h(u_x, u_y, u_z) = f(u_x)f(u_y)f(u_z)$$

¹This can be proven.

- It follows that

$$\frac{\partial \ln h(u)}{\partial u_x} = \frac{\partial}{\partial u_x} (\ln f(u_x) + \text{terms not involving } u_x) = \frac{d \ln f(u_x)}{du_x}$$

- Since

$$\begin{aligned} u^2 &= u_x^2 + u_y^2 + u_z^2 \\ \frac{\partial}{\partial u_x} (u^2) &= \frac{\partial}{\partial u_x} (u_x^2 + u_y^2 + u_z^2) \\ 2u \frac{\partial u}{\partial u_x} &= 2u_x \\ \frac{\partial u}{\partial u_x} &= \frac{u_x}{u} \end{aligned}$$

we have that

$$\begin{aligned} \frac{\partial \ln h}{\partial u_x} &= \frac{d \ln h}{du} \frac{\partial u}{\partial u_x} = \frac{u_x}{u} \frac{d \ln h}{du} \\ \frac{d \ln h(u)}{u du} &= \frac{d \ln f(u_x)}{u_x du_x} \end{aligned}$$

which generalizes to

$$\frac{d \ln h(u)}{u du} = \frac{d \ln f(u_x)}{u_x du_x} = \frac{d \ln f(u_y)}{u_y du_y} = \frac{d \ln f(u_z)}{u_z du_z}$$

- Since u_x, u_y, u_z are independent, we know that the above equation is equal to a constant, which we may call $-\gamma$. It follows that for any $j = x, y, z$, we have that

$$\begin{aligned} \frac{d \ln f(u_j)}{u_j du_j} &= -\gamma \\ \frac{1}{f} \frac{df}{du_j} &= -\gamma u_j \\ \int \frac{df}{f} &= \int -\gamma u_j du_j \\ \ln f &= -\frac{\gamma}{2} u_j^2 + C \\ f(u_j) &= A e^{-\gamma u_j^2} \end{aligned}$$

where we have incorporated the $1/2$ into γ .

- To determine A and γ , we let arbitrarily let $j = x$. Since f is a continuous probability distribution, we may apply the normalization requirement.

$$\begin{aligned} 1 &= \int_{-\infty}^{\infty} f(u_x) du_x \\ &= 2A \int_0^{\infty} e^{-\gamma u_x^2} du_x \\ &= 2A \sqrt{\frac{\pi}{4\gamma}} \\ A &= \sqrt{\frac{\gamma}{\pi}} \end{aligned}$$

- Additionally, since we have that $\langle u_x^2 \rangle = \frac{1}{3} \langle u^2 \rangle$ and $\langle u^2 \rangle = 3RT/M$, we know that $\langle u_x^2 \rangle = RT/M$. This combined with the definition of $\langle u_x^2 \rangle$ as a continuous probability distribution yields

$$\begin{aligned} \frac{RT}{M} &= \langle u_x^2 \rangle \\ &= \int_{-\infty}^{\infty} u_x^2 f(u_x) \, du_x \\ &= 2\sqrt{\frac{\gamma}{\pi}} \int_0^{\infty} u_x^2 e^{-\gamma u_x^2} \, du_x \\ &= 2\sqrt{\frac{\gamma}{\pi}} \cdot \frac{1}{4\gamma} \sqrt{\frac{\pi}{\gamma}} \\ &= \frac{1}{2\gamma} \\ \gamma &= \frac{M}{2RT} \end{aligned}$$

- Therefore,

$$f(u_x) = \sqrt{\frac{M}{2\pi RT}} e^{-Mu_x^2/2RT}$$

- It is common to rewrite the above in terms of molecular quantities m and k_B .
- It follows that as temperature increases, more molecules are likely to be found with higher component velocity values.
- We can use the above result to show that

$$\langle u_x \rangle = \int_{-\infty}^{\infty} u_x f(u_x) \, du_x = 0$$

- We can also calculate that $\langle u_x^2 \rangle = RT/M$ and $m \langle u_x \rangle^2 / 2 = k_B T / 2$ from the above result^[2].
 - An important consequence is that the total kinetic energy is divided equally into the x -, y -, and z -components. This fact was also demonstrated in Week 1, Lecture 3.
- **Doppler broadening:** The broadening of spectral lines due to the distribution of molecular velocities.
 - Ideally, spectral lines will be very narrow.
 - However, due to the Doppler effect, if an atom or molecule emits radiation of frequency ν_0 while moving away or toward the observer with speed u_x , then the observed frequency will be

$$\nu \approx \nu_0 \left(1 + \frac{u_x}{c} \right)$$

- Indeed, “if one observes the radiation emitted from a gas at temperature T , then it is found that the spectral line at ν_0 will be spread out by the Maxwell distribution of u_x of the molecule emitting the radiation” (McQuarrie & Simon, 1997, p. 1021).
- It follows by the definition of $f(u_x)$ and the above that

$$I(\nu) \propto e^{-mc^2(\nu-\nu_0)^2/2\nu_0^2 k_B T}$$

i.e., that $I(\nu)$ is of the form of a Gaussian centered at ν_0 with variance $\sigma^2 = \nu_0^2 k_B T / mc^2$.

- **Deriving Maxwell-Boltzmann distribution.**

²See the equipartition of energy theorem from Labalme (2021d).

- Let the probability that a molecule has speed between u and $u + du$ be defined by a continuous probability distribution $F(u) du$. In particular, we have from the above isotropic condition that

$$\begin{aligned} F(u) du &= f(u_x) du_x f(u_y) du_y f(u_z) du_z \\ &= \left(\frac{m}{2\pi k_B T} \right)^{3/2} e^{-m(u_x^2 + u_y^2 + u_z^2)/2k_B T} du_x du_y du_z \end{aligned}$$

- Considering F over a **velocity space**, we realize that we may express the probability distribution F as a function of u via $u^2 = u_x^2 + u_y^2 + u_z^2$ and the differential volume element in every direction over the sphere of equal velocities (a sphere by the isotropic condition) by $4\pi u^2 du = du_x du_y du_z$.
- Thus, the Maxwell-Boltzmann distribution in terms of speed is

$$F(u) du = 4\pi \left(\frac{m}{2\pi k_B T} \right)^{3/2} u^2 e^{-mu^2/2k_B T} du$$

- **Maxwell-Boltzmann distribution:** The distribution of molecular speeds.



Figure 3.1: The Maxwell-Boltzmann distribution.

- **Velocity space:** A rectangular coordinate system in which the distances along the axes are u_x, u_y, u_z .
- We may use the above result to calculate that

$$\langle u \rangle = \sqrt{\frac{8RT}{\pi m}}$$

which only differs from u_{rms} by a factor of 0.92.

- **Most probable speed:** The most probable speed of a gas molecule in a sample that obeys the Maxwell-Boltzmann distribution. Denoted by u_{mp} . Given by

$$u_{\text{mp}} = \sqrt{\frac{2RT}{M}}$$

- Derived by setting $dF/du = 0$.

- We may also express the Maxwell-Boltzmann distribution in terms of energy via $u = \sqrt{2\varepsilon/m}$ and $du = d\varepsilon / \sqrt{2m\varepsilon}$ to give

$$F(\varepsilon) d\varepsilon = \frac{2\pi}{(\pi k_B T)^{3/2}} \sqrt{\varepsilon} e^{-\varepsilon/k_B T} d\varepsilon$$

- We can also confirm our previously calculated values for $\langle u^2 \rangle$ and $\langle \varepsilon \rangle$.
- McQuarrie and Simon (1997) does a higher-level derivation of the ideal gas law that is rather analogous to the one done in class (i.e., via its flux perspective).
- McQuarrie and Simon (1997) discusses a simple and Nobel-prize winning experiment that verified the Maxwell-Boltzmann distribution.

3.5 Chapter 19: The First Law of Thermodynamics

From McQuarrie and Simon (1997).

- 1/31:
- **Thermodynamics:** The study of various properties and, particularly, the relations between the various properties of systems in equilibrium.
 - Primarily an experimental science that is still of great practical value to the fields of today.
 - “All the results of thermodynamics are based on three fundamental laws. These laws summarize an enormous body of experimental data, and there are absolutely no known exceptions” (McQuarrie & Simon, 1997, p. 765).
 - **Classical thermodynamics:** The development of thermodynamics before the atomic theory of matter.
 - Since thermodynamics was not developed in concert with the atomic theory, we can rest assured that its results will not need to be modified. However, it provides limited insight into what is going on at the molecular level.
 - **Statistical thermodynamics:** The molecular interpretation of thermodynamics developed since the atomic theory of matter became generally accepted.
 - Chapters 17-18 are an elementary treatment of statistical thermodynamics.
 - Since atomic structure is still being determined, these results are not on as solid of a footing as classical thermodynamics.
 - **First Law of Thermodynamics:** The law of conservation of energy applied to a macroscopic system.
 - **System:** The part of the world we are investigating.
 - **Surroundings:** Everything else.
 - **Heat:** The manner of energy transfer that results from a temperature difference between the system and its surroundings. *Denoted by q .*
 - Sign convention: Heat input into a system is positive; heat evolved by a system is negative.
 - **Work:** The transfer of energy between the system of interest and its surroundings as a result of the existence of unbalanced forces between the two. *Denoted by w .*
 - Sign convention: Work done *on* the system (i.e., that increases the energy of the system) is positive; work done *by* the system (i.e., that increases the energy of the surroundings) is negative.
 - Work can be related to raising a mass.
 - If a pressurized gas is capped by a piston with a mass m on top and then it pushes the piston upwards a distance h , it does $w = -mgh$ of work.
 - Knowing that the external pressure $P_{\text{ext}} = F/A = mg/A$ and $Ah = \Delta V$, we recover

$$w = -P_{\text{ext}}\Delta V$$

- If P_{ext} is not constant,

$$w = - \int_{V_i}^{V_f} P_{\text{ext}} dV$$

- **Definite state:** A state of a system in which all of the variables needed to describe the system completely are defined.

- For example, the state of one mole of an ideal gas can be described completely via P , V , and T . In fact, we need only specify two of these since $PV = RT$ for one mole of gas (in particular, specifying any two specifies the third).
- **State function:** A property that depends only upon the state of the system, not upon how the system was brought to that state.
 - State functions can be integrated in a normal way, i.e., $\Delta U = U_2 - U_1 = \int_1^2 dU$. In particular, we need not worry about the *path* from state 1 to state 2, only that we got from U_1 to U_2 .
 - Energy is a state function, but work and heat are not state functions.
- **Reversible process:** An expansion or compression in which P_{ext} and P differ only infinitesimally.
 - Technically, a reversible process would take infinite time, but it serves as a useful idealized limit regardless.
- To calculate w_{rev} (reversible work) for the compression of an ideal gas isothermally, we may replace P_{ext} by the pressure of the gas P to obtain

$$w_{\text{rev}} = - \int_{V_1}^{V_2} P dV = - \int_{V_1}^{V_2} \frac{nRT}{V} dV = -nRT \ln \frac{V_2}{V_1}$$

- Isothermal compression/expansion of a gas.



Figure 3.2: Isothermal manipulation of an ideal gas.

- Imagine a gas at pressure P , volume V , and temperature T in a container with a moveable piston at the top.
- Isothermal compression (Figure 3.2a).
 - Let the external pressure be held constant at P_{ext} . If the volume of the gas is initially at V_2 and temperature T , its pressure will be P_i . Thus, to equilibrate with the external pressure, it will compress to volume V_1 and final pressure $P_f = P_{\text{ext}}$ along the isotherm. As the force doing this work is the constant external pressure, the work done will be encapsulated by the red box in Figure 3.2a. Note that the gas necessarily releases an amount of heat equivalent to the work in the red box during the course of the compression to maintain isothermal conditions.
- Isothermal expansion (Figure 3.2a).
 - Let the external pressure be held constant at P_{ext} . If the volume of the gas is initially at V_1 and temperature T , its pressure will be P_i . Thus, to equilibrate with the external pressure, it will expand to volume V_1 and final pressure $P_f = P_{\text{ext}}$ along the isotherm. As the force doing this work is the constant external pressure, the work done will be encapsulated by the red box in Figure 3.2a. Note that the gas necessarily absorbs an amount of heat equivalent to the work in the red box during the course of the compression to maintain isothermal conditions.
- Reversible compression/expansion (Figure 3.2c).

- In a very slow manner, incrementally increase (resp. decrease) P_{ext} so as to allow the gas to reversibly compress (resp. expand).
- Compressing a gas reversibly and isothermally does the minimum amount of work on the gas. Expanding a gas reversibly and isothermally requires the gas to do the maximum amount of work.
- This work done on the gas raises the internal energy of the system, right? So shouldn't that raise the temperature, making the process not isothermal?
- Where does the extra energy above and below the isotherm in the irreversible processes go? Is it converted to heat?
- Does this mean that if you used a 1000 kg weight to compress a gas to half its original volume vs. using a 10 kg weight to compress a gas to half its original volume, the gas would get 100 times hotter in the former case?
- **Path function:** A function whose value depends on the path from state 1 to state 2, not just the initial and final states.
 - Path functions cannot be integrated in the normal way. Mathematically, they have **inexact differentials**, i.e., we write $\int_1^2 \delta w = w$. This is because it makes no sense to write w_1 , w_2 , $w_2 - w_1$, or Δw , for example.
 - Work and heat are path functions.
- The First Law of Thermodynamics says that $dU = \delta q + \delta w$ (in differential form) and $\Delta U = q + w$ (in integrated form).
 - An important consequence is that even though δq and δw are separately path functions/inexact differentials, their sum is a state function/exact differential.
- **Adiabatic process:** A process in which no heat is transferred between the system and its surroundings.
- Work during an adiabatic process.
 - For an adiabatic process, $\delta q = 0$.
 - Thus, $w = \Delta U$.
 - But since ΔU is entirely dependent on temperature, we have that

$$w = \Delta U = \int_{T_1}^{T_2} \left(\frac{\partial U}{\partial T} \right)_V dT = \int_{T_1}^{T_2} C_V(T) dT$$

- Why isn't this an integral of P with respect to V ?

- Temperature during an adiabatic process.
 - We have that

$$\begin{aligned} dU &= dw \\ C_V(T) dT &= -P dV \\ n\bar{C}_V(T) dT &= -\frac{nRT}{V} dV \\ \int_{T_1}^{T_2} \frac{\bar{C}_V(T)}{T} dT &= -R \int_{V_1}^{V_2} \frac{dV}{V} \end{aligned}$$

- In the specific case of a monatomic ideal gas, $C_V(T) = 3/2$. Thus, continuing, we have

$$\begin{aligned} \frac{3}{2} \ln \frac{T_2}{T_1} &= \ln \frac{V_1}{V_2} \\ \left(\frac{T_2}{T_1} \right)^{3/2} &= \frac{V_1}{V_2} \end{aligned}$$

- We can also express the above in terms of pressure

$$\left(\frac{P_2 V_2 / nR}{P_1 V_2 / nR}\right)^{3/2} = \frac{V_1}{V_2}$$

$$P_1 V_1^{5/3} = P_2 V_2^{5/3}$$

- For a diatomic gas, we end up with

$$P_1 V_1^{7/5} = P_2 V_2^{7/5}$$

- Note that for an isothermal expansion, Boyle's law applies: $P_1 V_1 = P_2 V_2$.

- Relating work and heat to molecular properties.

- By comparing recently derived equations with previously derived equations, we have that

$$U = \sum_j p_j E_j$$

$$dU = \sum_j p_j dE_j + \sum_j E_j dp_j$$

$$= \sum_j p_j \left(\frac{\partial E_j}{\partial V}\right)_N dV + \sum_j E_j dp_j$$

- The above equation suggests that we can interpret the first term as the average change in energy of a system caused by a small change in its volume, i.e., the average work.
- It follows by the First Law of Thermodynamics that we can interpret the second term as the average heat.
- This expresses the important but subtle notion that work results from “an infinitesimal change in the allowed energies of a system, without changing the probability distribution of its states” while heat results from “a change in the probability distribution of the states of a system, without changing the allowed energies” (McQuarrie & Simon, 1997, p. 780).
- In particular, if we take the process under study to be reversible, we have

$$dU = \underbrace{\sum_j p_j \left(\frac{\partial E_j}{\partial V}\right)_N dV}_{\delta w_{\text{rev}}} + \underbrace{\sum_j E_j dp_j}_{\delta q_{\text{rev}}} = \underbrace{\sum_j p_j \left(\frac{\partial E_j}{\partial V}\right)_N dV}_{-P} + \sum_j E_j dp_j$$

- The second equality above expresses the fact that

$$P = -\left\langle \frac{\partial E}{\partial V} \right\rangle$$

which we previously used in Chapter 17.

- For a constant-volume process, $w = 0$, so we know that the heat evolved in the process $q_V = \Delta U$.
- Defining a state function analogous to U for constant-pressure processes.

- We have from the First Law that

$$\Delta U = q + w = q - \int_{V_1}^{V_2} P dV$$

- Thus, at constant pressure,

$$q_P = \Delta U + P_{\text{ext}} \int_{V_1}^{V_2} dV = \Delta U + P \Delta V$$

– The above equation suggests how to define our new state function.

- **Enthalpy:** The state function describing the heat put into a system at constant pressure. *Denoted by H . Given by*

$$H = U + PV$$

– ΔH can be determined experimentally as the heat associated with a constant-pressure process.

- Examples.

– For the melting of ice, ΔV is small, so $\Delta U \approx \Delta H$.

– For the vaporization of water, ΔV is large, so $\Delta U < \Delta H$.

- We interpret the excess by the fact that most of the energy goes into raising the internal energy of the water (i.e., breaking the hydrogen bonding), but some of it must go into increasing the volume of the system against the atmospheric pressure.

- For reactions or processes that involve ideal gases,

$$\Delta H = \Delta U + RT\Delta n_{\text{gas}}$$

where Δn_{gas} is the difference in the number of moles of gaseous products vs. reactants.

- **Extensive quantity:** A quantity that depends on the amount of substance.

– Heat capacity is an extensive quantity.

- Heat capacity is a path function, as it depends on whether we heat the substance at constant volume or constant pressure.

- We have

$$C_V = \left(\frac{\partial U}{\partial T} \right)_V \qquad C_P = \left(\frac{\partial H}{\partial T} \right)_P$$

– We expect $C_P > C_V$ since we also have to work against atmospheric pressure.

– In fact, for a monatomic ideal gas,

$$\begin{aligned} H &= U + PV \\ &= U + nRT \\ \frac{dH}{dT} &= \frac{dU}{dT} + nR \\ C_P - C_V &= nR \\ C_P &= \frac{3}{2}R + nR \\ &= \frac{5}{2}R \end{aligned}$$

– It follows that the difference between C_P and C_V is significant for gases, but not for solids and liquids.

– Note that we can also prove a general expression for $C_P - C_V$ (see Chapter 22).

- Note that

$$R = 0.08314 \frac{\text{L bar}}{\text{mol K}}$$

- Relative enthalpies can be determined from heat capacity data and heats of transition.

– Integrate $C_P(T)$ from T_1 to T_2 , adding in $\Delta_{\text{fus}}H$ and $\Delta_{\text{vap}}H$ as necessary.

- **Thermochemistry:** The branch of thermodynamics which concerns the measurement of the evolution or absorption of energy as heat associated with chemical reactions.
- McQuarrie and Simon (1997) reviews exothermic/endothermic reactions, and $\Delta_r H = H_{\text{prod}} - H_{\text{react}}$.
- **Hess's Law:** The additivity property of $\Delta_r H$ values.
- **Standard reaction enthalpy:** The enthalpy change associated with one mole of a specified reagent when all reactants and products are in their standard states. *Denoted by $\Delta_r H^\circ$.*
 - An intensive quantity.
- **Standard molar enthalpy of formation:** The standard reaction enthalpy for the formation of one mole of a molecule from its constituent elements. *Denoted by $\Delta_f H^\circ$.*
 - We can obtain such values even if a compound cannot be formed directly from its elements via several related reactions and Hess's Law.
- To calculate $\Delta_r H$ for the general chemistry equation



using standard heats of formation, we decompose the reactants into their component elements in their normal states and then reassemble them into the products.

– Mathematically,

$$\Delta_r H = y\Delta_f H^\circ[\text{Y}] + z\Delta_f H^\circ[\text{Z}] - a\Delta_f H^\circ[\text{A}] - b\Delta_f H^\circ[\text{B}]$$

- The enthalpies of reaction at temperatures T_1, T_2 are related by

$$\Delta_r H(T_2) - \Delta_r H(T_1) = \int_{T_1}^{T_2} \Delta C_P(T) \, dT$$

where

$$\Delta C_P(T) = yC_{P,Y}(T) + zC_{P,Z}(T) - aC_{P,A}(T) - bC_{P,B}(T)$$

Week 4

Entropy and the Second Law of Thermodynamics

4.1 Entropy Equations

- 1/31:
- We define a new state function S by $dS = \delta q_{\text{rev}}/T$ and call it **entropy**.
 - See notes from last time for why this is a state function.
 - Verify that the same definition of entropy is a state function for any system.
 - Consider an ideal gas system in thermal equilibrium with an arbitrary system and drive the ideal gas system along a loop.
 - Around the cycle: $\Delta S_{\text{total}} = 0$.
 - Ideal gas:

$$\begin{aligned}\Delta S_{\text{total}} &= \Delta S_1 + \Delta S_2 \\ &= \int \frac{\delta q_{\text{rev}_1}}{T} - \int \frac{\delta q_{\text{rev}_1}}{T} \\ &= \int \frac{\delta q_{\text{rev}_1}}{T} + \int \frac{\delta q_{\text{rev}_2}}{T}\end{aligned}$$

- We must devise a reversible process to calculate the entropy changes for an irreversible process leading to the same final state.



Figure 4.1: Two linked containers.

- Imagine two linked containers, one filled with n moles of gas and the other vacuumed.
- Opening the two containers to each other results in an adiabatic expansion. All vibrational/rotational energy of the molecules is consumed and used for translation.
- Measuring the temperature with spectroscopy (the Maxwell-Boltzmann distribution of each spectral line, plus only the ground rovibrational states are occupied now) shows a drastic drop in temperature.
- We have $\delta q = 0$ and $\delta w = 0$ so that $dU = 0$ and $\Delta T = 0$ overall?
- An isothermal expansion is a reversible process leading to the same final state.

- $dU = 0$ implies $\delta q_{\text{rev}} = -\delta w = P dV$.
- We have that

$$\Delta S = \int \frac{\delta q_{\text{rev}}}{T} = \int_{V_0}^{2V_0} \frac{P dV}{T} = \int_{V_0}^{2V_0} \frac{nRT}{V} \frac{1}{T} dV = nR \ln 2$$

- Using entropy as a state function to predict the vapor pressure in equilibrium with its liquid, from the enthalpy at boiling and the boiling temperature.



Figure 4.2: Vapor pressure thermodynamic loop.

- Consider the above thermodynamic loop, where T is the temperature of the water and P is the pressure above the water.
- We have that

$$\Delta S_1 = \int_T^{T_b} \frac{C_{P_l}}{T} dT \quad \Delta S_2 = \frac{\Delta H_{\text{vap}}}{T_b} \quad \Delta S_3 = nR \ln \frac{P_0}{P} \quad \Delta S_4 = \int_{T_b}^T \frac{C_{P_g}}{T} dT$$

and that

$$\Delta S_0 = \frac{\Delta H_{\text{vap}}}{T}$$

- We know that ΔS around the loop is zero since S is a state function. We neglect the heat capacity effect. Thus,

$$\begin{aligned} \frac{\Delta H_{\text{vap}}}{T_b} + nR \ln \frac{P_0}{P} - \frac{\Delta H_{\text{vap}}}{T} &= 0 \\ \ln \frac{P_0}{P} &= \frac{\Delta H_{\text{vap}}}{nR} \left(\frac{1}{T} - \frac{1}{T_b} \right) \\ P &= P_0 e^{-\Delta H_{\text{vap}}/nR(1/T - 1/T_b)} \end{aligned}$$

- The above equation gives the vapor pressure at T in terms of the vapor pressure P_0 at T_b .

- **Trouton's rule:** The statement that

$$\frac{\Delta H_{\text{vap}}}{T_b} \approx 85 \pm 5 \text{ J mol}^{-1} \text{ K}^{-1}$$

- Discovered this rule as an undergrad after an afternoon's manipulation of data from a book of tables.
- This rule reflects the fact that

$$\frac{\Delta H_{\text{vap}}}{T_b} = \Delta S_{\text{vap}}$$

and implies that ΔS_{vap} is approximately a constant.

- Example of entropy change: The direction of heat flow between two systems (1 and 2) only in thermal contact.

- We have

$$\begin{aligned}\delta q_{\text{rev}_1} &= \delta q_{\text{rev}_2} \\ C_{V_1} dT_1 &= -C_{V_2} dT_2\end{aligned}$$

- Thus,

$$\begin{aligned}dS &= dS_1 + dS_2 \\ &= \frac{\delta q_{\text{rev}_1}}{T_1} + \frac{\delta q_{\text{rev}_2}}{T_2} \\ &= \frac{C_V dT_1}{T_1} - \frac{C_V dT_1}{T_2} \\ &= C_V dT_1 \left(\frac{1}{T_1} - \frac{1}{T_2} \right)\end{aligned}$$

- The conclusion is that if $dT_1 > 0$, then $dS > 0$. This is the spontaneous direction, the direction that nature chooses, the one in which entropy increases.
- The maximum of S is the equilibrium temperature between the two systems.
- Entropy change of the isothermal mixing of two ideal gases at the same temperature.
 - Consider the same two-container setup from Figure 4.1.
 - We have that

$$\begin{aligned}\Delta S &= Rn_1 \ln \frac{V_1 + V_2}{V_1} + Rn_2 \ln \frac{V_1 + V_2}{V_2} \\ &= R(n_1 + n_2) \left(\frac{n_1}{n_1 + n_2} \ln \frac{V_1 + V_2}{V_1} + \frac{n_2}{n_1 + n_2} \ln \frac{V_1 + V_2}{V_2} \right) \\ &= R(n_1 + n_2)(-y_1 \ln y_1 - y_2 \ln y_2) \\ &= R(n_1 + n_2)[-y_1 \ln y_1 - (1 - y_1) \ln(1 - y_1)]\end{aligned}$$

- Note that $y_1 = n_1/(n_1 + n_2) = V_1/(V_1 + V_2)$ is the mole fraction, and similarly for y_2 .
- The conclusion is that $\Delta S > 0$.
- The maximum of ΔS is at $y_1 = y_2 = 1/2$.
- **Gibb's paradox:** Suppose you have the same gas on both sides of the containers. Then $\Delta S = nR \ln 2$ for an indistinguishable gas.
 - This is wrong.
 - Resolved by knowing that the gases *must* be distinguishable.

4.2 Statistical Entropy in Various Systems

2/2:

- For an isolated system, energy is conserved but the entropy keeps on increasing until the system reaches thermal equilibrium.
- Thermal equilibrium is reached when entropy is maximum for a constant energy.
- The sign of the entropy change in a spontaneous process for an isolated system is positive.
- Entropy is the only macroscopic physical quantity that requires a particular direction for time, sometimes called an **arrow of time**.
- **Second law of thermodynamics:** The entropy of an isolated system can only increase.

- **Clasuius inequality:** The following inequality, where equality holds iff the process is reversible.

$$\Delta S \geq \int \frac{\delta q}{T}$$

- Considers the isolated system to justify.
- Statistical entropy: $S = k_B \ln W$ where W is the number of microstates of the system (i.e., the number of possible ways the system can be arranged).
 - Shows additivity of the log.
 - When doubling the volume available to a gas, $\Delta S = Nk_B \ln 2$. $W_{\text{after}} = 2^N W_{\text{before}}$.
 - The statistical definition of entropy avoids the Gibbs paradox since at a molecular level, we can differentiate between particles.
- Goes over calculating $W(n_1, n_2)$.
- The ways we can distinguish the number of molecules in the container becomes smaller and smaller as we increase the number of particles.
- Consider two identical containers at fixed temperature with N non-interacting indistinguishable molecules $n_1 + n_2 = N$.
 - $W(n_1) = W(n_1, n_2)$ is the number of ways to arrange the molecules between containers 1 and 2.
 - We have

$$\begin{aligned} \ln W(n_1, n_2) &= \ln N! - \ln n_1! - \ln(N - n_1)! \\ &= N \ln N - N - [n_1 \ln n_1 - n_1 + (N - n_1) \ln(N - n_1) - (N - n_1)] \\ &= N \ln N - n_1 \ln n_1 - (N - n_1) \ln(N - n_1) \\ &= (n_1 + n_2) \ln N - n_1 \ln n_1 - n_2 \ln n_2 \\ &= -n_1 \ln \frac{n_1}{N} - n_2 \ln \frac{n_2}{N} \\ &= N \left(-\frac{n_1}{N} \ln \frac{n_1}{N} - \frac{n_2}{N} \ln \frac{n_2}{N} \right) \end{aligned}$$

- Therefore,

$$S = Nk_B(-p_1 \ln p_1 - p_2 \ln p_2)$$

- Entropy for a set of systems expressed in terms of the probability for these systems to be in a certain state.
 - Covers $W(n_1, \dots, n_r)$.
 - We have

$$\begin{aligned} \ln W &= \ln A! - \sum_i \ln a_i! \\ &= A \ln A - A - \sum_i (a_i \ln a_i - a_i) \\ &= A \ln A - \sum_i a_i \ln a_i \\ &= \left(\sum_i a_i \right) \ln A - \sum_i a_i \ln a_i \\ &= \sum_i \left(-a_i \ln \frac{a_i}{A} \right) \\ &= A \sum_i \left(-\frac{a_i}{A} \ln \frac{a_i}{A} \right) \\ &= A \sum_i (-p_i \ln p_i) \end{aligned}$$

- Therefore,

$$S = Ak_B \sum_i (-p_i \ln p_i)$$

- We will use this result to derive the Boltzmann Factor.

4.3 Further Entropy Relations and Phenomena

- 2/4: • Verifying that the entropy is at a maximum when all the states have equal probability.

- In the case of just two states,

$$\begin{aligned} S &= Nk_B(-p_1 \ln p_1 - p_2 \ln p_2) \\ dS &= Nk_B[-d(p_1 \ln p_1) - d(p_2 \ln p_2)] \\ &= Nk_B[-dp_1 \cdot \ln p_1 - p_1 \cdot \frac{dp_1}{p_1} - dp_2 \cdot \ln p_2 - p_2 \cdot \frac{dp_2}{p_2}] \\ &= Nk_B(-dp_1)(\ln p_1 - \ln p_2) \\ &= 0 \end{aligned}$$

if $p_1 = p_2$.

- Thus, since $dS/dp_1|_{p_1=1/2} = 0$, we know that a graph of entropy vs. p_1 has a maximum at $p_1 = 1/2$.

- In the case of many buckets,

$$dS = -Nk_B \sum_i \ln p_i dp_i = 0$$

iff all p_i are the same, because then we could pull out the $\ln p_i$ and take $\sum dp_i = 0$.

- Fluctuations from the equilibrium state are very unlikely in a macroscopic system. Entropy measures the likelihood.

- At equilibrium, we have a number of configurations W_{eq} . Outside of equilibrium, we have a number of configurations W .

- Thus,

$$\begin{aligned} \Delta S &= S - S_{eq} = k_B \ln \frac{W}{W_{eq}} \\ \frac{W}{W_{eq}} &= e^{\Delta S/k_B} \end{aligned}$$

- For $\Delta S = R = 8.31 \text{ J mol}^{-1} \text{ K}^{-1}$, we have that $W = W_{eq}e^{N_A}$, which will never happen?

- Deriving the Boltzmann factor by identifying the thermodynamic and statistical entropies.

- Consider two states E_1, E_2 . Then $\delta q_{rev} = (E_1 - E_2)$.

- We have

$$\begin{aligned} dS &= \frac{\delta q_{rev}}{T} & dS &= -k_B N d(p_1 \ln p_1 + p_2 \ln p_2) \\ & & &= -k_B N (\ln p_1 - \ln p_2) dp_1 \end{aligned}$$

- Setting them equal to each other, we get

$$\begin{aligned}\frac{dp_1(E_1 - E_2)}{T} &= -k_B \ln \frac{p_1}{p_2} dp_1 \\ \frac{E_1 - E_2}{k_B T} &= -\ln \frac{p_1}{p_2} \\ \frac{p_1}{p_2} &= e^{-(E_1 - E_2)/k_B T}\end{aligned}$$

which is the Boltzmann factor, as desired.

- Expressing the entropy in terms of the partition function.

- Derives

$$S = k_B T \frac{\partial \ln Q}{\partial T} + k_B \ln Q$$

from $S = -Nk_B \sum_i p_i \ln p_i$ as in McQuarrie and Simon (1997).

- Entropy of an ideal monatomic gas.

- Derives

$$S = \frac{5}{2} Nk_B + Nk_B \ln \left[\left(\frac{2\pi m k_B T}{h^2} \right)^{3/2} \frac{V}{N} g_{e1} \right]$$

as in McQuarrie and Simon (1997).

- Note that

$$\left(\frac{2\pi m k_B T}{h^2} \right)^{3/2} \approx \left(\frac{1}{\lambda_{dB}} \right)^3$$

where λ_{dB} is the de Broglie wavelength.

- The Carnot engine. The most efficient thermal engine between two temperature reservoirs.
 - Recall that the area inside the cycle on a P/V diagram is the work produced by the cycle.
 - It is adiabatic expansion, isothermal expansion, adiabatic compression, isothermal compression.
 - Does the same efficiency analysis as in McQuarrie and Simon (1997).
 - PGS says it's ok to treat the processes as reversible because they're isothermal by definition.

4.4 Office Hours (PGS)

- Week 1, Lecture 3: When we calculate the average kinetic energy to be $3k_B T/2$, what constraints do we have on the system. Is it truly any system that can be described by a partition function? And if so, why does this result differ from the one presented in McQuarrie and Simon (1997) for a diatomic ideal gas?
 - There aren't constraints on the system under study.
 - $PV = \frac{2}{3} \langle KE \rangle$ is true without ideality.
 - $\langle E \rangle = 3k_B T/2$ is only true with a monatomic ideal gas.
 - DOFs contribute to $\langle E \rangle$ (at very high temperatures).
 - But in the equation from Chapter 17 for a diatomic, some terms disappear as $T \rightarrow \infty$.
- Week 1, Lecture 3: What does calculating the expected energy of a harmonic oscillator have to do with anything? Are you just doing it to show that we can for the quantized system? Also, are we integrating because this is a “continuous” partition function, and if it is a continuous partition function, is it so because the harmonic oscillator can, in theory, assume any bond distance even though its *energies* are quantized?

- Week 1, Lecture 3: How is $k_B T$ the average energy of a classical harmonic oscillator? Is it the sum of the kinetic energy and potential energy, both equal to $k_B T/2$? Is this the virial theorem again?
 - Also a high-temperature limit thing.
- What was Week 1 about? It mostly seemed like a lot of examples to sort of motivate the course.
 - Yes, this was basically just to motivate the course.
 - Take away: $k_B T$ determines the range of energies you'll be in for everything.
 - Things much higher in energy won't be **thermally populated**.
 - Things much lower are equally populated.
- Quiz 1, Questions 5-6. Plus will things like surface tension be on the midterm?
 - Surface tension is the same thing as pressure in two dimensions.
 - We can do $\gamma A = nRT$.
 - $A\gamma$ becomes 2 of the kinetic energy $1/2 k_B T$.
 - Work this out for myself!
- Week 2, Lecture 2: The vibrational partition function at $T = h\nu/k_B$ equals zero. What is the significance of this? Is it just that, that the average vibrational energy is equal to zero here?
- Do the rotational and vibrational temperatures have a nice physical interpretation?
 - They are the energies of vibrations/rotations on a scale of temperature instead of energy (the two are directly proportional).
 - The theoretical rotational/vibrational temperature/energy.
 - It's the same temperature as the speeds. When molecules collide, they exchange energy.
 - After enough collisions (as a system tends toward thermal equilibrium), the rotational temperature becomes the same as vibrational.
- Week 3, Lecture 1: In your derivation of the Maxwell-Boltzmann distribution, where does the $4\pi v^2$ come from? Is the sphere you referred to in class the same one as in the textbook, this "velocity space?" If so, could you explain a bit more what a velocity space is?
- The internal energy of a system is only a function of temperature because in the expression for $\langle E \rangle$ in terms of Q , T is the only variable that is algebraically in the equation and the partial derivative is held constant with respect to all variables except T .
- I'm very confused on the distinction between a reversible and irreversible process.
 - Reversible means that at every point, you are in thermal equilibrium, so you can specify the state of the system.
 - Remember that T is not always specified for a system (only for systems in thermal equilibrium, i.e., that obey the Maxwell-Boltzmann distribution)! The same goes for P and V .
 - If the process happens very suddenly or you don't know what the temperature/pressure are during the process, that's irreversible. Something burning in a closed container is irreversible, but at the beginning and end you are in thermal equilibrium, so you can imagine a reversible path between the two.
 - Energy is the same regardless of path, entropy is the same regardless of path.
 - Temperature is not uniform throughout the system during an expansion of a gas from a filled container into a vacuumed container (the system is not in thermal equilibrium). The particles in the filled container will still be at the initial temperature, but those that have left are expending their KE to do so and cooling drastically.

- *Any* line on a PV diagram describes a reversible process. The processes may differ in terms of non-state functions such as heat exchange or work, but the net ΔU , ΔH , and ΔS will be the same.
 - The internal energy will be different *along* different paths, but it will make the same net change.
- A reversible process need not be isothermal. We can merely consider a piecewise path that contains isothermal components because they are easy to work with mathematically, if the quantity we're interested in calculating is a state function.
- If we want to calculate the work for a reversible process, we need $w = -\int_1^2 P dV$. If we want to calculate ΔU , we need the state variables T_1 and T_2 but that's it (since U is a state function of T).
- Could you also speak to exact vs. inexact differentials and how to tell if something is a state function by its differential?
 - If you have a state function, this means that you can express it in terms of the state variables, but you need to know them.
 - Examples of state variables are energy, temperature, volume, and number of moles.
 - Once you have a relation between those variables, the state function is uniquely defined between those variables. Entropy is absolute (with physical meaning); enthalpy and energy have defined values once we choose a zero. These latter values, however, have no physical meaning; they only allow us to calculate ΔU and ΔH , which do have physical meaning.
 - We find that $\delta q_{\text{rev}} = C_V dT$ and $\delta w = -P dV$. When we take $\delta q_{\text{rev}} = dU + P dV$, this differential does not describe a state function because it is not the derivative of any function (P is a state variable, but it's also a function of V and T [the latter being the problem since it appears in a term with only dV]). However, if we divide by T for an ideal gas, then we can get the form $M(T) dT + N(V) dV$, where $M(T) = C_V(T)/T$ and $N(V) = nR/V$.
 - Note that this differential satisfies $dN/dT = dM/dV = 0$ as per Labalme (2021b). In fact, this directly expresses the fact that M is constant with respect to y and N is constant with respect to x .
 - Note that according to this definition, we don't theoretically need M, N to only be functions of one variable.
 - Maxwell's relation: A function $f(x, y)$ such that $df/dx dy = df/dy dx$.
 - Allows us to relate variables that we might not generally be able to.
- Week 3, Lecture 3: I know you said when you were working with the differential of entropy that $dH = \delta q$ at constant pressure ($dP = 0$). This makes sense. Did you say that at constant volume, we have that $dH = \delta q$ as well?
- Week 3, Lecture 3: How do we obtain the enthalpies of formation at nonstandard temperatures?
- Do we need to know anything about Doppler broadening? I believe you alluded to it once.
- How are P_{ext} and P related, especially in the context of a PV graph? What exactly is the " P " in $w = -\int_1^2 P dV$?
 - In a reversible process, the pressures are the same. This extends to *any* path along a PV diagram, not just an isothermal one, as per our previous discussion.
- What is with the isothermal approximation thing? Do we just approximate everything as isothermal? When can and can't we do this?
- Page 39 of my notes questions.

- What gives us the right to express work during an adiabatic process as an integral with respect to temperature? Is there an equivalent formulation in terms of an integral with respect to V ?
- Quiz 2, Question 4.
- Week 4, Lecture 1: You calculated that $\Delta S = nR \ln 2$ for n moles of a gas expanding to twice its initial volume. I assume this doesn't relate to the Gibbs Paradox?
 - The Gibbs paradox says that the final and initial states of two gases mixing are the same, and thus there is no change in entropy.
 - However, we are implicitly assuming that the gases are expanding by virtue of the fact that they're initially separated and thus distinguishable.
 - A paradox doesn't mean that something isn't true, it just means that our understanding is incomplete.
 - The $nR \ln 2$ isn't characteristic of the Gibbs paradox; it's characteristic of a doubling in volume, in whatever context.
- Week 4, Lecture 3: How do you get to $W = W_{\text{eq}} e^{-\Delta S/T}$?
 - It is $+\Delta S$; it's just that ΔS is a negative quantity when you're moving away from equilibrium.
- In an irreversible process vs. a reversible process, is there a difference in entropy?
 - See the Clausius inequality. If you want more heat, do reversible. If you want more work, do irreversible.
- The experiment in Homework 2, Question 1.
 - The Clément-Desormes experiment.
- Do we need another variable for HW2, Q2?
 - You do need an initial temperature T_h .
 - PGS will send a note.

4.5 MathChapter J: The Binomial Distribution and Stirling's Approximation

From McQuarrie and Simon (1997).

- Counting the number of ways to arrange N distinguishable objects into two groups of size N_1, N_2 where $N_1 + N_2 = N$.
 - There are $N!$ ways to arrange N distinguishable objects, $N!/(N - N_1)!$ ways to arrange the objects in group 1, and $N_2!$ ways to arrange the objects in group 2. Thus, there are

$$\frac{N!}{(N - N_1)!} \cdot N_2!$$

permutations of the N objects in two groups.

- For example, $N = 4$, $N_1 = 3$, and $N_2 = 1$, we are currently counting both $abc : d$ and $bac : d$ as different ways of arranging the four objects into two groups, when clearly such ordering does not matter.
- Dividing the above by the number of ways to arrange N_1 objects in the first group ($N_1!$) and the number of ways to arrange N_2 objects in the second group ($N_2!$) gives the desired result.

$$W(N_1, N_2) = \frac{N!}{N_1! N_2!}$$

- Now we have a result that, as per the previous example, allows us to count only $abc : d$, $bcd : a$, $cda : b$, and $dab : c$.

- McQuarrie and Simon (1997) reviews the binomial expansion in light of the above result's status as a binomial coefficient.
- Counting the number of ways to arrange N distinguishable objects into r groups of size N_1, \dots, N_r where $N_1 + \dots + N_r = N$.

$$W(N_1, \dots, N_r) = \frac{N!}{N_1! \dots N_r!}$$

- Note that this quantity is called a **multinomial coefficient** because it occurs in the multinomial expansion $(x_1 + \dots + x_r)^N$.
- **Asymptotic approximation:** An approximation to a function that gets better as the argument of the function increases.
- **Stirling's approximation:** An asymptotic approximation to $\ln N!$. *Given by*

$$\ln N! = N \ln N - N$$

- Proof: We have that

$$\ln N! = \sum_{n=1}^N \ln n$$

- For N large, this sum behaves more and more like the integral $\int_1^N \ln x \, dx$. Thus, we take

$$\ln N! = \sum_{n=1}^N \ln n \approx \int_1^N \ln x \, dx = N \ln N - N$$

- A refinement of the approximation is the following.

$$\ln N! = N \ln N - N + \frac{1}{2} \ln(2\pi N)$$

4.6 Chapter 20: Entropy and the Second Law of Thermodynamics

From McQuarrie and Simon (1997).

- 2/1:
- The change of energy alone is not sufficient to determine the direction of a spontaneous process.
 - Although mechanical and chemical systems tend to evolve in such a way as to minimize their energy, we can find examples of spontaneous chemical processes that are not exothermic.
 - Examples include the mixing of two gases and the highly endothermic (and spontaneous) reaction of $\text{Ba}(\text{OH})_2$ and NH_4NO_3 .
 - Such processes obey the First Law of Thermodynamics, but their spontaneous direction cannot be explained by it.
 - Each of these “special cases” involves an increase in the disorder of the system.
 - For example, in the mixing of gases, we can show quantum mechanically that increasing the volume of the container increases the number of accessible translational states.
 - Competition between the drive to lower energy and the drive to increase disorder.
 - Simple mechanical systems can't become that much more disordered; thus, energy considerations dominate.
 - The mixing of gases doesn't change the energy that much; thus, disorder considerations dominate.

- Defining a quantitative state function describing disorder.

- Note that

$$\begin{aligned}\delta q_{\text{rev}} &= dU - \delta w_{\text{rev}} \\ &= C_V(T) dT + P dV \\ &= C_V(T) dT + \frac{nRT}{V} dV\end{aligned}$$

is an inexact differential since the second term cannot be written as a derivative of some function of T and V (because T depends on V). In particular, the integral depends on what path through T and V we take.

- However, if we divide both sides of the above by T , we get an exact differential, i.e., a state function.
- Note that we can show that this result holds for all systems, not just an ideal gas.
- **Entropy:** The state function describing the disorder of a system. *Denoted by S . Given by*

$$dS = \frac{\delta q_{\text{rev}}}{T}$$

- **Integrating factor:** A term that converts an inexact differential to an exact (integrable) differential.
 - $1/T$ is an integrating factor of δq_{rev} .
- Since entropy is a state function, $\Delta S = 0$ for a cyclic process, i.e.,

$$\oint dS = 0$$

- McQuarrie and Simon (1997) calculates ΔS for a process that proceeds from state 1 to state 2 isothermally, and adiabatically/isochorically, to show that the quantity is the same in both cases.
- Justifying $dS = \delta q_{\text{rev}}/T$ qualitatively:
 - Increase in heat means increase in disorder (check).
 - Same increase in heat at a lower temperature increases disorder more since there is more order at lower temperatures (check).
- **Isolated** (system): A system that is separated from its surroundings by rigid walls that do not allow matter or energy to pass through them.
- Unlike energy, entropy is not necessarily conserved; it can increase within an isolated system if a spontaneous process takes place therein.
- The entropy of a system is at its maximum when the system is equilibrium; at this point, $dS = 0$.
- Consider an isolated system consisting of two compartments. One compartment holds large, one-component system A , and other holds B . They are separated by a heat-conducting wall.
 - Because of isolation,

$$U_A + U_B = \text{constant}$$

$$V_A = \text{constant}$$

$$S = S_A + S_B$$

$$V_B = \text{constant}$$

- Since V_A, V_B are fixed, $dV = 0$, meaning that $dU = \delta q_{\text{rev}} + 0$. It follows that

$$\begin{aligned} dS &= dS_A + dS_B \\ &= \frac{dU_A}{T_A} + \frac{dU_B}{T_B} \\ &= dU_B \left(\frac{1}{T_B} - \frac{1}{T_A} \right) \end{aligned}$$

- Since the gases A and B can still mix without absorbing energy, we define dS_{prod} as the entropy produced by the system and redefine $\delta q/T$ as dS_{exch} (the entropy exchanged with the surroundings via a transfer of heat).
- It follows that for an reversible process ($dS_{\text{prod}} = 0$), we have

$$dS = \frac{\delta q_{\text{rev}}}{T}$$

while for an irreversible process ($dS_{\text{prod}} > 0$), we have

$$dS = dS_{\text{prod}} + \frac{\delta q_{\text{irr}}}{T} > \frac{\delta q_{\text{irr}}}{T}$$

- **Inequality of Clausius:** The following inequality. *Given by*

$$\Delta S \geq \int \frac{\delta q}{T}$$

- **Second Law of Thermodynamics:** There is a thermodynamic state function of a system called the entropy S such that for any change in the thermodynamic state of the system, $dS \geq \delta q/T$, where equality holds iff the change is carried out reversibly.
- “Because the universe itself may be considered to be an isolated system and all naturally occurring processes are irreversible, one statement of the Second Law of Thermodynamics says that the entropy of the universe is constantly increasing. In fact, Clausius summarized the first two laws of thermodynamics by, ‘The energy of the Universe is constant; the entropy is tending to a maximum’” (McQuarrie & Simon, 1997, p. 829).
- Relating entropy, a thermodynamic quantity, to a statistical quantity.

- Consider an ensemble of \mathcal{A} isolated systems, each with number of particles N , volume V , and energy $E(N, V)$.
- Let $\Omega(E)$ be the degeneracy of E , i.e., the number of quantum states with energy E ^[1]. Label the $\Omega(E)$ quantum states by $j = 1, 2, \dots, \Omega(E)$.
- Let a_j be the number of systems in state j .
- It follows that the number of ways of having a_1 systems in state 1, a_2 systems in state 2, etc. is given by

$$W(a_1, \dots, a_{\Omega(E)}) = \frac{\mathcal{A}!}{a_1! \cdots a_{\Omega(E)}!} = \frac{\mathcal{A}!}{\prod_j (a_j!)}$$

with $\sum_j a_j = \mathcal{A}$.

- If every system is in one totally ordered state (i.e., $a_j = \mathcal{A}$ for some j), $W = 1$. On the other end of the spectrum, W can be massive for disorder.
- As W is a measure of entropy, we are now free to relate S and W , in particular via

$$S = k_B \ln W$$

¹Note that for systems relatively far from the ground state, $\Omega(E) \approx e^N$.

- We choose a log because we want to be able to split S into $S_A + S_B$ and have the math reflect that. In particular, for two systems A, B , $W_{AB} = W_A W_B$, which nicely works out such that

$$S_{AB} = k_B \ln W_{AB} = k_B \ln W_A + k_B \ln W_B = S_A + S_B$$

- McQuarrie and Simon (1997) goes over an alternate “derivation” of the above in terms of the degeneracy to get $S = k_B \ln \Omega$.

2/3: • Since entropy is a state function, we calculate entropy changes via a reversible process.

- Imagine a gas expanding from V_1 to V_2 in a non-isolated system.
- Although this is an adiabatic process, since entropy is a state function, we may perform the easier isothermal calculation for the “equivalent” reversible process.

$$\Delta S_{\text{sys}} = \int_1^2 \frac{\delta q_{\text{rev}}}{T} = - \int_1^2 \frac{\delta w_{\text{rev}}}{T} = \int_{V_1}^{V_2} \frac{P}{T} dV = nR \int_{V_1}^{V_2} \frac{dV}{V} = nR \ln \frac{V_2}{V_1}$$

- Note that it is the change to an isothermal integral that allows us to assume $dU = 0$ (internal energy is a function of only temperature).
- However, there is still a difference between reversible and irreversible processes.
 - In the reversible, isothermal process, $q_{\text{rev}} = -w_{\text{rev}} = -nRT \ln(V_2/V_1)$. Thus,

$$\Delta S_{\text{univ}} = \Delta S_{\text{sys}} + \Delta S_{\text{surr}} = nR \ln \frac{V_2}{V_1} - nR \ln \frac{V_2}{V_1} = 0$$

as we would expect for an reversible process.

- In the irreversible, adiabatic process, however, $\Delta S_{\text{surr}} = 0^{[2]}$. Thus,

$$\Delta S_{\text{univ}} = nR \ln \frac{V_2}{V_1} > 0$$

as we would expect for an irreversible process.

- The isothermal mixing of two ideal gases (say N_2 and Br_2).

- Because the two gases are ideal, they act independently of each other. Thus,

$$\Delta S_{N_2} = n_{N_2} R \ln \frac{V_{N_2} + V_{Br_2}}{V_{N_2}} \quad \Delta S_{Br_2} = n_{Br_2} R \ln \frac{V_{N_2} + V_{Br_2}}{V_{Br_2}}$$

- It follows that

$$\begin{aligned} \Delta S &= \Delta S_{N_2} + \Delta S_{Br_2} \\ &= -n_{N_2} R \ln \frac{V_{N_2}}{V_{N_2} + V_{Br_2}} - n_{Br_2} R \ln \frac{V_{Br_2}}{V_{N_2} + V_{Br_2}} \\ &= -n_{N_2} R \ln \frac{n_{N_2}}{n_{N_2} + n_{Br_2}} - n_{Br_2} R \ln \frac{n_{Br_2}}{n_{N_2} + n_{Br_2}} \\ \Delta \bar{S} &= -\frac{n_{N_2}}{n_{N_2} + n_{Br_2}} R \ln \frac{n_{N_2}}{n_{N_2} + n_{Br_2}} - \frac{n_{Br_2}}{n_{N_2} + n_{Br_2}} R \ln \frac{n_{Br_2}}{n_{N_2} + n_{Br_2}} \\ \frac{\Delta_{\text{mix}} \bar{S}}{R} &= -y_{N_2} \ln y_{N_2} - y_{Br_2} \ln y_{Br_2} \end{aligned}$$

where $V \propto n$ by the ideal gas law, $\Delta \bar{S}$ is the *molar* change in entropy, and y_{N_2} is the mole fraction of N_2 (same for bromine), and $\Delta_{\text{mix}} \bar{S}$ indicates that this is the molar change in entropy for the *mixing* of two gases.

²McQuarrie and Simon (1997) justify $\Delta S_{\text{surr}} = 0$ for an irreversible *isothermal* process by $\Delta U = 0$ and $P_{\text{ext}} = 0$ imply $w_{\text{irr}} = 0$ and therefore $q_{\text{irr}} = 0$.

- For the isothermal mixing of N ideal gases, we have

$$\Delta_{\text{mix}} \bar{S} = -R \sum_{j=1}^N y_j \ln y_j$$

- ΔS when two equal sized pieces of the same metal at different temperatures (T_h, T_c) are brought into thermal contact and then isolated from the surroundings.

- Both pieces of metal will approach the same final temperature T as per

$$C_V(T_h - T) = C_V(T - T_c)$$

$$T = \frac{T_h + T_c}{2}$$

- There is essentially no work done, so $dU = \delta q_{\text{rev}}$.
- Thus, taking C_V to be constant from T_c to T_h yields

$$\Delta S = \int_{T_1}^{T_2} \frac{\delta q_{\text{rev}}}{T} = \int_{T_1}^{T_2} \frac{C_V dT}{T} = C_V \ln \frac{T_2}{T_1}$$

- Therefore, we have

$$\Delta S_h = C_V \ln \frac{T_h + T_c}{2T_h} \qquad \Delta S_c = C_V \ln \frac{T_h + T_c}{2T_c}$$

- It follows that

$$\Delta S = \Delta S_h + \Delta S_c = C_V \ln \frac{(T_h + T_c)^2}{4T_h T_c}$$

where

$$(T_h - T_c)^2 = T_h^2 - 2T_h T_c + T_c^2 > 0$$

$$T_h^2 + 2T_h T_c + T_c^2 = (T_h + T_c)^2 > 4T_h T_c$$

implies that $\Delta S > 0$, as desired for an irreversible process.

- The Carnot cycle (for a steam engine).

- Each cycle, the engine (system) “withdraws energy [q_h] as heat from some high-temperature thermal reservoir, uses some of the energy to do work [w], and then discharges the rest of the energy [q_c] as heat to a lower-temperature thermal reservoir” (McQuarrie & Simon, 1997, p. 838).
- Treating the process as reversible (since both internal energy and entropy as state functions) gives us the following analysis.
- Since this is a cycle, we have that

$$\Delta U_{\text{engine}} = w + q_{\text{rev,h}} + q_{\text{rev,c}} = 0 \qquad \Delta S_{\text{engine}} = \frac{\delta q_{\text{rev,h}}}{T_h} + \frac{\delta q_{\text{rev,c}}}{T_c}$$

- It follows if we define the maximum efficiency of the engine to be the quotient of work done by the engine ($-w$) and heat input ($q_{\text{rev,h}}$) that

$$\begin{aligned} \text{maximum efficiency} &= \frac{-w}{q_{\text{rev,h}}} \\ &= \frac{q_{\text{rev,h}} + q_{\text{rev,c}}}{q_{\text{rev,h}}} \\ &= 1 + \frac{-T_c q_{\text{rev,h}}/T_h}{q_{\text{rev,h}}} \\ &= \frac{T_h - T_c}{T_h} \end{aligned}$$

- It follows for typical values of $T_h = 573\text{ K}$ and $T_c = 373\text{ K}$ that maximum efficiency $\approx 35\%$.
- Moreover, it implies that engines run with higher-temperature heat reservoirs and lower-temperature cold reservoirs are more efficient, regardless of design.
- Note that $T_h = T_c$ implies that maximum efficiency = 0%, i.e., no net work can be obtained from an isothermal cyclic process.
- The above result leads to **Kelvin's statement of the Second Law**.
- **Kelvin's statement of the Second Law:** A closed system operating in an isothermal cyclic manner cannot convert heat into work without some accompanying change in the surroundings.
- Expressing entropy in terms of a partition function.

$$\begin{aligned}
 S_{\text{ensemble}} &= k_B \ln W \\
 &= k_B \ln \frac{\mathcal{A}!}{\prod_j a_j!} \\
 &= k_B \ln \mathcal{A}! - k_B \sum_j \ln a_j! \\
 &= k_B \mathcal{A} \ln \mathcal{A} - k_B \mathcal{A} - k_B \sum_j a_j \ln a_j + k_B \sum_j a_j && \text{(Stirling's approximation)} \\
 &= k_B \mathcal{A} \ln \mathcal{A} - k_B \sum_j a_j \ln a_j && (\sum_j a_j = \mathcal{A}) \\
 &= k_B \mathcal{A} \ln \mathcal{A} - k_B \sum_j p_j \mathcal{A} \ln p_j \mathcal{A} && (p_j = a_j / \mathcal{A}) \\
 &= k_B \mathcal{A} \ln \mathcal{A} - k_B \mathcal{A} \sum_j p_j \ln p_j - k_B \mathcal{A} \ln \mathcal{A} \sum_j p_j \\
 \frac{S_{\text{ensemble}}}{\mathcal{A}} &= -k_B \sum_j p_j \ln p_j && (\sum_j p_j = 1) \\
 S_{\text{system}} &= -k_B \sum_j p_j \ln p_j \\
 &= -k_B \sum_j \frac{e^{-\beta E_j}}{Q} (-\beta E_j - \ln Q) \\
 &= \beta k_B \sum_j \frac{E_j e^{-\beta E_j}}{Q} + k_B \ln Q \sum_j \frac{e^{-\beta E_j}}{Q} \\
 &= \frac{1}{T} \cdot \langle E \rangle + k_B \ln Q \\
 S_{\text{system}} &= k_B T \frac{\partial \ln Q}{\partial T} + k_B \ln Q
 \end{aligned}$$

- For a monatomic ideal gas where all atoms are in their ground electronic state, we have

$$\bar{S} = \frac{5}{2}R + R \ln \left[\left(\frac{2\pi m k_B T}{h^2} \right)^{3/2} \frac{\bar{V} g_{e1}}{N_A} \right]$$

- The mixing of two ideal gases (say N_2 and Br_2) from a molecular perspective.
 - Since the natural log in \bar{S} for a monatomic ideal gas contains the same number of terms involving V as for a diatomic ideal gas, we have that

$$\begin{aligned}
 S &= N k_B \ln V + \text{terms not involving } V \\
 &= n R \ln V + \text{terms not involving } V
 \end{aligned}$$

for both N_2 and Br_2 .

- Thus, the initial state is given by

$$\begin{aligned} S_1 &= S_{1,N_2} + S_{1,Br_2} \\ &= n_{N_2} R \ln V_{N_2} + n_{Br_2} R \ln V_{Br_2} + \text{terms not involving } V \end{aligned}$$

and the final state is given by

$$\begin{aligned} S_2 &= S_{2,N_2} + S_{2,Br_2} \\ &= n_{N_2} R \ln(V_{N_2} + V_{Br_2}) + n_{Br_2} R \ln(V_{N_2} + V_{Br_2}) + \text{terms not involving } V \end{aligned}$$

- Therefore,

$$\begin{aligned} \Delta_{\text{mix}} S &= S_2 - S_1 \\ &= n_{N_2} R \ln \frac{V_{N_2} + V_{Br_2}}{V_{N_2}} + n_{Br_2} R \ln \frac{V_{N_2} + V_{Br_2}}{V_{Br_2}} \\ \frac{\Delta_{\text{mix}} \bar{S}}{R} &= -y_{N_2} \ln y_{N_2} - y_{Br_2} \ln y_{Br_2} \end{aligned}$$

as expected.

- Relating $S = k_B \ln W$ to $dS = \delta q_{\text{rev}}/T$ (and proving that $\beta = 1/k_B T$!).

- We have that

$$\begin{aligned} S &= -k_B \sum_j p_j \ln p_j \\ \frac{dS}{dp_j} &= -k_B \sum_j \left(p_j \cdot \frac{1}{p_j} + 1 \cdot \ln p_j \right) \\ dS &= -k_B \sum_j (dp_j + \ln p_j dp_j) \\ &= -k_B \sum_j (-\beta E_j - \ln Q) dp_j \\ &= \beta k_B \sum_j E_j dp_j + \ln Q \sum_j dp_j \\ &= \beta k_B \delta q_{\text{rev}} \quad (\sum_j p_j = 1 \Rightarrow \sum dp_j = 0) \end{aligned}$$

as desired.

- Additionally, the above result implies that βk_B is an integrating factor of δq_{rev} , i.e.,

$$\begin{aligned} \beta k_B &= \frac{1}{T} \\ \beta &= \frac{1}{k_B T} \end{aligned}$$

Week 5

Entropy and the Third Law of Thermodynamics

5.1 Entropy and the Third Law

2/7: • An experimental determination of entropy.

- We have that

$$\begin{aligned}dU &= \delta q_{\text{rev}} + \delta w_{\text{rev}} \\ &= T dS - P dV\end{aligned}$$

- At constant V , $dV = 0$, so $dS = dU / T = C_V dT / T$.
- More rigorously, we can do the rest of the derivation as in McQuarrie and Simon (1997) to get

$$\left(\frac{\partial S}{\partial T}\right)_V = \frac{C_V}{T} \qquad \left(\frac{\partial S}{\partial V}\right)_T = \frac{1}{T} \left[P + \left(\frac{\partial U}{\partial V}\right)_T \right]$$

- It follows that $\Delta S = \int_{T_1}^{T_2} C_V / T dT$.
 - This is one way to measure the change in entropy.
 - However, it's not very practical since it's very hard to do constant volume chemistry.

- Derives

$$\left(\frac{\partial S}{\partial T}\right)_P = \frac{C_P}{T} \qquad \left(\frac{\partial S}{\partial P}\right)_T = \frac{1}{T} \left[\left(\frac{\partial H}{\partial P}\right) - V \right]$$

as in McQuarrie and Simon (1997).

- It follows that $\Delta S = \int_{T_1}^{T_2} C_P / T dT$, too.
- However, since this is ΔS at constant pressure, we expect it to be bigger than ΔS at constant volume because work is done.

- **Third Law:** The entropy of pure ordered crystals is zero at 0 K.

- Solid CO is *not* ordered, as discussed in McQuarrie and Simon (1997).
- Derive with $S = k_B \ln W$.

- Since we can't actually achieve 0 K and measure properties such as entropy there, we appeal to the Debye model to determine entropy there and in the vicinity.

- **Debye model:** The statement that $C_P(T) \propto T^3 = AT^3$ for insulating crystals.

- It follows that for T_0 sufficiently close to 0 K,

$$\Delta S = \int_0^{T_0} \frac{AT^3}{T} dT = \frac{1}{3}AT_0^3 = \frac{C_P(T_0)}{3}$$

- Justifying the Debye model.

- According to the Einstein model, each atom is a harmonic oscillator of frequency $h\nu$. This implies, however, that $C_P(T) \approx 1/(e^{h\nu/k_B T} - 1)$ drops too fast at low T .
 - Debye says that sound waves are harmonic oscillators with very low frequencies. If you take $h\nu \ll k_B T$, or $hc/\lambda \ll k_B T$, you are concerned with all λ sufficiently large (or wavevectors $k = 2\pi/\lambda$ sufficiently small). The wavevectors are within a sphere in a wavevector space with basis k_x, k_y, k_z . The radius of this sphere is proportional to T . Thus, its volume, which contains what we're interested in, is $4/3\pi r^3$ and thus the sum of all the wavevectors is proportional to T^3 . Thus, the number of states of energy less than $k_B T$ is proportional to T^3 and each has $k_B T$ of energy (thus, if you were trying to find a scaling for the energy, it would be $U(T) \approx T^4$). Note also $h\nu = hc/\lambda = 2\pi\hbar c/\lambda = \hbar ck$.
- Comparing experimental and theoretical values of entropy of gases to the theory value from the partition function.
 - Ideal gas phase S is accurately calculated with $Q = q^N/N!$.
 - Table 21.4 compares the agreement for seven substances and finds it accurate to within 0.1%.
 - The discrepancies come from disorder at 0 K, as described in McQuarrie and Simon (1997).

5.2 Introduction to Free Energies

2/9:

- Discusses the entropy trends from Chapter 21.
 - More atoms ($\text{CO} < \text{CO}_2$) increases entropy.
 - Heavier atoms increases entropy.
 - Floppier molecules (pentane $>$ cyclopentane) have greater entropy.
 - Gases have greater entropy (this rule is king).
- Example (21-42):
 - When determining which chemical reaction has the greatest increase in entropy, look at the change in number of moles of gas as the first tie-breaker.
- Consider two systems: Water and ice in a container in thermal contact with a surrounding at zero celsius, and water at room temperature and ice at zero celsius in a thermally isolated container.
 - The first process is reversible since we can melt by raising the temperature slightly (allowing heat flow into the system) and vice versa for freezing.
 - The second process is irreversible because it tends toward thermal equilibrium and there is no way to undo the final equilibrium.
 - We can consider reversible paths for both processes though to calculate state functions.
 - Goes over how to calculate the final temperature T in the second process, which is necessary to get the entropy componentwise.
- Free energies decrease in a spontaneous process.
 - We know that $dS \geq 0$ for an isolated system, and that $dS \geq \delta q/T$ for any system (equality holds for reversible processes).

- Rearranging, we have that

$$0 \geq \delta q - T dS$$

- It follows since $dU = \delta q + \delta w$ that

$$\begin{aligned} dU - T dS &= (\delta q - T dS) + \delta w \\ &\leq \delta w \end{aligned}$$

- But since $A = U - TS$ is clearly a state function (as a combination of state functions and state variables),

$$dA = dU - T dS - S dT$$

is an exact differential for a state function. In particular, at constant temperature,

$$dA = dU - T dS$$

- Therefore, $dA \leq \delta w$ at constant temperature.
- Thus, at constant T, V, w , $dA \leq 0$. In other words, A is monotonically decreasing.

- **Helmholtz free energy:** The state function A defined above.

- We use this because it's more “natural” to think about finding the lowest free *energy* state than the largest *entropy* state.
- The Helmholtz free energy is most often used in physics where there isn't often a “pressure bath.” Chemists prefer the **Gibbs free energy**, which is constructed exactly the same way.

- **Gibbs free energy:** The state function describing the free energy of a system at constant pressure. Denoted by G . Given by

$$G = H - TS$$

- It follows from our prior results that at constant temperature,

$$\begin{aligned} dG &= dH - T dS \\ &= dU + d(PV) - T dS \\ &= \delta q + \delta w + P dV + V dP - T dS \\ &= \delta q - P dV + \delta w_{\text{non-rev } PV} + P dV + V dP - T dS \\ &= \delta q - T dS + \delta w_{\text{non-rev } PV} + V dP \\ &\leq 0 + \delta w_{\text{non-rev } PV} + V dP \end{aligned}$$

- Thus, at constant T, P and with no other w done on the system, $dG \leq 0$.
- In a reversible process (constant T, P), $dG = \delta w_{\text{non-rev } PV}$.

- Note that δw is the work put into the system, so $\delta w \geq \delta w_{\text{rev}} = dA$.
- The system produces more work when operated reversibly.
- Gibbs free energy and electrochemical work: The case of hydrogen fuel cells vs. thermal engines.
 - The hydrogen fuel cell is based off of the reaction



with $\Delta H = 285.76 \text{ kJ/mol}$ and $\Delta G = 237.1 \text{ kJ/mol}$.

- ΔH can be used to produce heat $q = \Delta H$.
- If we want to produce work, we have to run a thermal engine that will have efficiency bounded above by the Carnot cycle's $(T_h - T_c)/T_h$.

- With new materials, people have been able to use hotter thermal reservoirs and achieve efficiencies up to 50%!
- ΔG can be converted to work directly (as current and voltage in the fuel cell). Thus, in theory, running a reversible fuel cell produces more work than burning H_2 in a reversible engine.
- The limitation is overpotential, though, and we would need a better hydrogen burning catalyst.
- Maxwell relations:

$$\frac{\partial^2 f}{\partial x \partial y} = \frac{\partial^2 f}{\partial y \partial x}$$

- If we want to show that $(\partial P / \partial T)_V = (\partial S / \partial V)_T$, we take

$$\begin{aligned} dA &= \delta q_{\text{rev}} - T dS - S dT - P dV \\ &= -S dT - P dV \\ &= \left(\frac{\partial A}{\partial T} \right)_V + \left(\frac{\partial A}{\partial V} \right)_T \end{aligned}$$

so that

$$\frac{\partial P}{\partial T} = \frac{\partial}{\partial T} \left(- \frac{\partial A}{\partial V} \right) = - \frac{\partial^2 A}{\partial T \partial V} = - \frac{\partial^2 A}{\partial V \partial T} = \frac{\partial}{\partial V} \left(- \frac{\partial A}{\partial T} \right) = \frac{\partial S}{\partial V}$$

- Facts like these help us show things such as

$$\Delta S = \int_{V_1}^{V_2} \frac{\partial S}{\partial V} dV = \int_{V_1}^{V_2} \frac{\partial P}{\partial T} dV = \int_{V_1}^{V_2} \frac{nR}{V} dV = nR \ln \frac{V_2}{V_1}$$

so we don't need to measure the heat flow to measure entropy when we have a more complex equation of state than $PV = nRT$.

5.3 Differential Relations

- 2/11: • A useful equation for the change in energy per unit volume at constant temperature, and equation of state.

- We have that

$$dU = \left(\frac{\partial U}{\partial V} \right)_T dV + \left(\frac{\partial U}{\partial T} \right)_V dT$$

- It follows that

$$\begin{aligned} dU &= T dS - P dV \\ &= T \left(\frac{\partial S}{\partial T} \right)_V dT + \left(\frac{\partial S}{\partial V} \right)_T dV - P dV \\ &= \underbrace{T \left(\frac{\partial S}{\partial T} \right)_V}_{(\partial U / \partial T)_V} dT + \underbrace{\left[\left(\frac{\partial S}{\partial V} \right)_T - P \right]}_{(\partial U / \partial V)_T} dV \end{aligned}$$

- Thus, by Maxwell relations,

$$\begin{aligned} \left(\frac{\partial U}{\partial V} \right)_T &= T \left(\frac{\partial S}{\partial V} \right)_T - P \\ &= T \left(\frac{\partial P}{\partial T} \right)_V - P \end{aligned}$$

- This equation can be evaluated given the equation of state.
- For an ideal gas, we get

$$\left(\frac{\partial U}{\partial V}\right)_T = T \frac{nR}{V} - P = 0$$

- For a van der Waals gas (just the excluded volume b part, not the a term), we get

$$\left(\frac{\partial U}{\partial V}\right)_T = T \frac{\partial}{\partial T} \left(\frac{nRT}{V - nb} \right) - P = \frac{nRT}{V - nb} - P = 0$$

- Now for small enough volume, we know that a gas will compress into a liquid. In other words, $T(\frac{\partial P}{\partial T})_V \geq P$ always, where equality is only a good approximation at sufficiently large volumes.
- A similar method can be applied to enthalpy changes with pressure.
 - We can derive in an analogous method that

$$\left(\frac{\partial H}{\partial P}\right)_T = V + T \left(\frac{\partial S}{\partial P}\right)_T$$

- You can only apply the chain rule to partial derivatives held constant with respect to the same variable(s).
- **Natural variables** (of a state function): A set of state variables for which the partial derivatives of the state function with respect to said variables are (simple expressions of) state variables or state functions.

- For example, S, P are the natural variables of U since $\partial U / \partial S = T$ and $\partial U / \partial V = -P$ as shown by

$$dU = T dS - P dV$$

- V, T are not natural variables of U , as we showed earlier, since the partial derivatives of U with respect to them are complicated expressions.
- For the other state functions, we have

$$\begin{aligned} dH &= T dS + V dP & dA &= -S dT - P dV \\ dS &= \frac{1}{T} dU + \frac{P}{T} dV & dG &= -S dT + V dP \end{aligned}$$

- Note that natural variables are not unique, as we also have

$$dS = \frac{1}{T} dH - \frac{V}{T} dP$$

- An example of using Maxwell relations to getting C_V to C_P (19-27 and 22-11).

- We have that

$$C_P - C_V = T \left(\frac{\partial P}{\partial T}\right)_V \left(\frac{\partial V}{\partial T}\right)_P$$

- We also have that

$$dV = \left(\frac{\partial V}{\partial T}\right)_P dT + \left(\frac{\partial V}{\partial P}\right)_T dP$$

- It follows if we let $dV = 0$ in the above expression that

$$\left(\frac{\partial P}{\partial T}\right)_V = - \frac{(\partial V / \partial T)_P}{(\partial V / \partial P)_T}$$

– Therefore,

$$\begin{aligned}
 C_P - C_V &= T \left[-\frac{(\partial V/\partial T)_P}{(\partial V/\partial P)_T} \right] \left(\frac{\partial V}{\partial T} \right)_P \\
 &= -T \frac{(\partial V/\partial T)_P^2}{(\partial V/\partial P)_T} \\
 &= -T \frac{V^2 \alpha^2}{-V \kappa} \\
 &= \frac{TV \alpha^2}{\kappa}
 \end{aligned}$$

where α is the **thermal expansion coefficient** and κ is the **isothermal compressibility**.

- **Thermal expansion coefficient:** The following constant. Denoted by α . Given by

$$\alpha = \frac{1}{V} \left(\frac{\partial V}{\partial T} \right)_P$$

- **Isothermal compressibility:** The following constant. Denoted by κ . Given by

$$\kappa = -\frac{1}{V} \left(\frac{\partial V}{\partial P} \right)_T$$

5.4 Office Hours (PGS)

- So is the content of the midterm Chapters 17-22 plus parts of Chapter 25 (the Maxwell-Boltzmann distribution)?
- Are there going to be any surprises I should look out for (a la surface tension)?
- Will questions be like homework questions, quiz questions?
- Are we going to have to know derivations or just formulas?

5.5 Chapter 21: Entropy and the Third Law of Thermodynamics

From McQuarrie and Simon (1997).

- 2/3:
- In this chapter, we will learn how to calculate absolute (as opposed to relative) values of entropy.
 - Relating thermodynamic quantities to entropy.
 - It follows from the First Law of Thermodynamics that

$$dU = \underbrace{T dS}_{\delta q_{\text{rev}}} - \underbrace{P dV}_{\delta w_{\text{rev}}}$$

- We have that the total differential of $U(V, T)$ is

$$dU = \frac{\partial U}{\partial T} dT + \frac{\partial U}{\partial V} dV$$

- Combining the two above equations, we have that

$$\begin{aligned}
 T dS - P dV &= \frac{\partial U}{\partial T} dT + \frac{\partial U}{\partial V} dV \\
 dS &= \frac{1}{T} \frac{\partial U}{\partial T} dT + \frac{1}{T} \left(P + \frac{\partial U}{\partial V} \right) dV \\
 &= \frac{C_V dT}{T} + \frac{1}{T} \left(P + \frac{\partial U}{\partial V} \right) dV
 \end{aligned}$$

- It follows by comparing the above with the total differential of $S(V, T)$ that

$$\left(\frac{\partial S}{\partial T}\right)_V = \frac{C_V}{T} \qquad \left(\frac{\partial S}{\partial V}\right)_T = \frac{1}{T} \left[P + \left(\frac{\partial U}{\partial V}\right)_T \right]$$

- Since

$$dH = dU + P dV + V dP = T dS + V dP$$

we can proceed in a similar manner to the above to obtain

$$\left(\frac{\partial S}{\partial T}\right)_P = \frac{C_P}{T} \qquad \left(\frac{\partial S}{\partial P}\right)_T = \frac{1}{T} \left[\left(\frac{\partial H}{\partial P}\right)_T - V \right]$$

- **Third Law of Thermodynamics:** Every substance has a finite positive entropy, but at zero kelvin, the entropy may become zero, and does so in the case of a perfectly crystalline substance.
 - While the first and second laws provide state functions (internal energy and entropy, respectively), the third law provides a numerical scale for entropy.
- The third law, although formulated before quantum mechanics, follows nicely from it: At absolute zero, every system in an ensemble will be in the same energy state, so $W = 1$; it follows that $S = k_B \ln 1 = 0$.
 - If, however, the ground state has a degeneracy of n , then

$$\begin{aligned} S(0 \text{ K}) &= -k_B \sum_j p_j \ln p_j \\ &= -k_B \sum_{j=1}^n \frac{1}{n} \ln \frac{1}{n} \\ &= k_B \ln n \end{aligned}$$

- Nevertheless, even for a degeneracy of N_A , we will only have $S(0 \text{ K}) \approx 7.56 \times 10^{-22} \text{ J mol}^{-1} \text{ K}^{-1}$, which is far less than any measurable value.
- Because of the third law, we can define entropy absolutely (assuming no phase change between 0 and T).

$$\begin{aligned} S(T) &= S(0) + S(T) - S(0) \\ &= S(0) + \Delta S \\ &= S(0) + \int_0^T \frac{C_P(t)}{t} dt \\ &= \int_0^T \frac{C_P(t)}{t} dt \end{aligned}$$

- Accounting for phase changes.
 - A phase change is a great example of a reversible process since we only need the temperature to be slightly above or slightly below the transition temperature T_{trs} to accomplish it.
 - Thus,

$$\Delta_{\text{trs}} S = \frac{q_{\text{rev}}}{T_{\text{trs}}} = \frac{\Delta_{\text{trs}} H}{T_{\text{trs}}}$$

- It follows that, as applicable,

$$S(T) = \int_0^{T_{\text{fus}}} \frac{C_P^s(t)}{t} dt + \frac{\Delta_{\text{fus}} H}{T_{\text{fus}}} + \int_{T_{\text{fus}}}^{T_{\text{vap}}} \frac{C_P^l(t)}{t} dt + \frac{\Delta_{\text{vap}} H}{T_{\text{vap}}} + \int_{T_{\text{vap}}}^T \frac{C_P^g(t)}{t} dt$$

- Note that with typical values plugged in, $\Delta_{\text{fus}}S \ll \Delta_{\text{vap}}S$.
- **Debye T^3 law:** As $T \rightarrow 0$ (i.e., for about $T = 0 - 15$ K), $C_P^s(T) \rightarrow T^3$ for most nonmetallic crystals and $C_P^s(T) \rightarrow aT + bT^3$ ($a, b \in \mathbb{R}_{\geq 0}$) for most metallic crystals.

– It follows by the absolute definition of S that

$$S(T) = \frac{C_P(T)}{3}$$

at low temperatures for nonmetallic solids.

- **Debye temperature:** A constant characteristic of the solid. *Denoted by Θ_D .*
- **Third-law entropy:** An absolute entropy value calculated according to the convention that $S(0 \text{ K}) = 0$. *Also known as **practical absolute entropy**.*
- McQuarrie and Simon (1997) calculates the third-law entropy of N_2 based on various thermodynamic data.
- **Standard entropy:** An entropy value of a gas as presented in the literature. *Denoted by S° . Units $\text{J mol}^{-1} \text{ K}^{-1}$.*
 - Standard entropies are, by convention, corrected for the nonideality of the gas at one bar (for how to calculate this correction, see Chapter 22).
- McQuarrie and Simon (1997) rederives $S(0 \text{ K}) = k_B \ln n$ from a partition function approach, and then gives a specific example for N_2 , noting how well this value correlates with the one from the previous example.
 - “This type of agreement is quite common, and in many cases the statistical thermodynamic value is more accurate than the calorimetric value. . . The accepted literature values are often a combination of statistical thermodynamic and calorimetric values” (McQuarrie & Simon, 1997, p. 863).
 - Also gives a linear symmetric example (CO_2).
- Entropy trends.
 1. “The standard molar entropies of the gaseous substances are the largest, and the standard molar entropies of the solid substances are the smallest” (McQuarrie & Simon, 1997, p. 865).
 2. “The increase in standard molar entropy of the noble gases is a consequence of their increasing mass as we move down the periodic table” (McQuarrie & Simon, 1997, p. 865).
 - More mass implies more translational energy levels are available implies more disorder.
 - This is a consequence of quantum mechanics — considering the formula for the energy of a particle in a 3D box, note that m is in the denominator. Thus, increasing m means that the levels are more closely spaced, and hence more are readily accessible.
 3. “Generally speaking, the more atoms of a given type in a molecule, the greater is the capacity of the molecule to take up energy and thus the greater is its entropy” (McQuarrie & Simon, 1997, p. 866).
 - More atoms implies more vibrational modes implies more disorder.
 4. “For molecules with approximately the same molecular masses, the more compact the molecule is, the smaller is its entropy” (McQuarrie & Simon, 1997, p. 867).
 - Between isomers, the one with more unrestricted motion will have greater entropy.
- We can calculate the entropy for substances that “don’t exist” via alternate paths.
 - For example, Br_2 is a liquid at 298 K, but we can still calculate $S^\circ[\text{Br}_{2(\text{g})}]$ by imagining the following path: Raise $\text{Br}_{2(\text{l})}$ to its boiling point; vaporize it; cool it back down to 298 K.

- This calculated value is in agreement with the spectroscopic value.
- **Residual entropy:** The difference between the calculated molar entropy of a substance and its experimental molar entropy. *Given by*

$$\bar{S}_{\text{calc}} - \bar{S}_{\text{exp}}$$
 - We define residual entropy this way because wherever there are discrepancies, it so happens that $S_{\text{calc}} > S_{\text{exp}}$.
- Large residual entropies are encountered for the linear molecules CO and N₂O.
 - This is because these molecules have small dipole moments, so upon crystallization, there is not a strong tendency for the molecule to align in the most energetically favorable way.
 - Thus, with molecules already locked in higher energy states at T_{fus} , as we cool to 0 K, we do not have a “perfect” crystal.
 - Since the ground state is two-fold degenerate in both of these cases (we have CO and OC, and NNO and ONN), $\bar{S}(0\text{ K}) = R \ln 2$ here.
- We can similarly account for the larger still residual entropy in H₃CD by noting that the ground state is four-fold degenerate, and thus $\bar{S}(0\text{ K}) = R \ln 4$ here.
- We can use standard molar entropies to calculate the entropy changes of chemical reactions.

5.6 Chapter 22: Helmholtz and Gibbs Energies

2/14:

- Since the $dS \geq 0$ criterion for isolated systems is quite restrictive (only applies to isolated systems), we introduce two new state functions that will help determine the direction of a spontaneous process in a non-isolated system.
- **Helmholtz energy:** A state function describing the energy of a system at constant volume and temperature. *Denoted by A . Given by*

$$A = U - TS$$

- Since $dV = dT = 0$, we know that

$$\begin{aligned} dU &= \delta q + \delta w \\ &\leq T dS - P dV \\ dU - T dS - S dT &\leq 0 \\ d(U - TS) &\leq 0 \end{aligned}$$

where equality holds for a reversible process and inequality holds for an irreversible process (by the properties of the Clausius inequality).

- Thus, the Helmholtz energy will decrease until all spontaneous processes have occurred. At this point, $dA = 0$ and A will be at a minimum.
- $dA \leq 0$ is the analog of $dS \geq 0$ for a spontaneous process to occur.
- “A process for which $\Delta A > 0$ cannot take place spontaneously in a system at constant T and V . Consequently, something, such as work, must be done on the system to effect the change” (McQuarrie & Simon, 1997, p. 883).
- “The Helmholtz energy represents this compromise between the tendency of a system to decrease its energy and to increase its entropy” (McQuarrie & Simon, 1997, p. 883).
- Note that the sign of U is more important at low T and vice versa for high T by $U - TS$.
- McQuarrie and Simon (1997) applies the definition of A to the mixing of two ideal gases, showing that $\Delta A < 0$ for it.

- A physical interpretation to ΔA .

- When a process is carried out along a reversible path,

$$\Delta A = \Delta U - q_{\text{rev}} = w_{\text{rev}}$$

meaning that ΔA is either the maximum work that can be pulled out of the system or the minimum work required to drive the process (depending on its sign).

- **Gibbs energy:** A state function describing the energy of a system at constant pressure and temperature. Denoted by G . Given by

$$G = H - TS$$

- Since $dP = dT = 0$, we know that

$$\begin{aligned} dU &= \delta q + \delta w \\ &\leq T dS - P dV \\ dU + P dV + V dP - T dS - S dT &\leq 0 \\ d(U + PV - TS) &\leq 0 \\ d(H - TS) &\leq 0 \end{aligned}$$

where equality holds for a reversible process and inequality holds for an irreversible process (by the properties of the Clausius inequality).

- The Gibbs energy is analogous in almost every way to the Helmholtz energy.
- We can also write $G = A + PV$ to relate G, A in the same manner that H, U are related, respectively.
- **Molar Gibbs energy of vaporization:** The following quantity. Denoted by $\Delta_{\text{vap}} \bar{G}$. Given by

$$\Delta_{\text{vap}} \bar{G} = \Delta_{\text{vap}} \bar{H} - T \Delta_{\text{vap}} \bar{S}$$

- Note that $\Delta_{\text{vap}} \bar{G} = 0$ at 1 atm and 100 °C. This means that the phases are in equilibrium under these conditions.
- A slight increase in temperature, however, makes $\Delta G < 0$ (i.e., makes vaporization spontaneous), while a slight decrease in temperature makes $\Delta G > 0$ (i.e., makes condensation spontaneous).
- McQuarrie and Simon (1997) covers that in a reversible process, $\Delta G = w_{\text{non PV}}$, i.e., the maximum non-pressure/volume work available from a process or minimum non-pressure/volume work required to drive a process.
 - $w_{\text{non PV}}$ can be electrical work, for example.
- **Maxwell relation:** An relation obtained by equating the second cross partial derivatives of a function.
- **Thermodynamic equation of state:** An equation that relates a thermodynamic function to functions of P , V , and T .
- McQuarrie and Simon (1997) notes that

$$\left(\frac{\partial U}{\partial V} \right)_T = -P + T \left(\frac{\partial P}{\partial T} \right)_V$$

meaning that energy does depend on volume somewhat (it's not entirely a function of temperature).

- We can show that the above derivative reduces to zero for an ideal gas, however.
- McQuarrie and Simon (1997) plays around with partial derivatives for a while.
 - Discusses the α, κ definition of $C_P - C_V$.

- We can also use Maxwell relations and differentials to derive the following results.

$$\begin{aligned}\left(\frac{\partial G}{\partial T}\right)_P &= -S & \left(\frac{\partial G}{\partial P}\right)_T &= V \\ -\left(\frac{\partial S}{\partial P}\right)_T &= \left(\frac{\partial V}{\partial T}\right)_P \\ \left(\frac{\partial H}{\partial P}\right)_T &= V - T\left(\frac{\partial V}{\partial T}\right)_P\end{aligned}$$

- Covers natural variables.

Thermodynamic energy	Differential expression	Corresponding Maxwell relations
U	$dU = T dS - P dV$	$\left(\frac{\partial T}{\partial V}\right)_S = -\left(\frac{\partial P}{\partial S}\right)_V$
H	$dH = T dS + V dP$	$\left(\frac{\partial T}{\partial P}\right)_S = \left(\frac{\partial V}{\partial S}\right)_P$
A	$dA = -S dT - P dV$	$\left(\frac{\partial S}{\partial V}\right)_T = \left(\frac{\partial P}{\partial T}\right)_V$
G	$dG = -S dT + V dP$	$\left(\frac{\partial S}{\partial P}\right)_T = -\left(\frac{\partial V}{\partial T}\right)_P$

Table 5.1: Thermodynamic energies, differentials, and Maxwell relations.

- Calculating corrections to the standard molar entropy of a gas.

- Basically, we want to find $S^\circ(1 \text{ bar}) - \bar{S}(1 \text{ bar})$. To do so, we consider

$$S^\circ(1 \text{ bar}) - \bar{S}(1 \text{ bar}) = [S^\circ(1 \text{ bar}) - \bar{S}(P_{\text{id}})] + [\bar{S}(P_{\text{id}}) - \bar{S}(1 \text{ bar})]$$

where P_{id} is a pressure sufficiently low such that the gas in question behaves ideally.

- We know from the definition of changes in entropy as an integral and Table 5.1 that

$$\begin{aligned}S^\circ(1 \text{ bar}) - \bar{S}(P_{\text{id}}) &= \int_{P_{\text{id}}}^{1 \text{ bar}} \left(\frac{\partial S}{\partial P}\right)_T dP & \bar{S}(P_{\text{id}}) - \bar{S}(1 \text{ bar}) &= \int_{1 \text{ bar}}^{P_{\text{id}}} \left(\frac{\partial S}{\partial P}\right)_T dP \\ &= - \int_{P_{\text{id}}}^{1 \text{ bar}} \left(\frac{\partial V}{\partial T}\right)_P dP & &= - \int_{1 \text{ bar}}^{P_{\text{id}}} \left(\frac{\partial V}{\partial T}\right)_P dP \\ &= - \int_{P_{\text{id}}}^{1 \text{ bar}} \frac{R}{P} dP & &= \int_{P_{\text{id}}}^{1 \text{ bar}} \left(\frac{\partial V}{\partial T}\right)_P dP\end{aligned}$$

- Note that we calculate the first change using the ideal gas law and the second change using the equation of state describing the real gas.

- Adding the above two equations yields

$$S^\circ(1 \text{ bar}) - \bar{S}(1 \text{ bar}) = \int_{P_{\text{id}}}^{1 \text{ bar}} \left[\left(\frac{\partial V}{\partial T}\right)_P - \frac{R}{P} \right] dP$$

- We can evaluate this equation using only the second virial coefficient (see Chapter 16).

- McQuarrie and Simon (1997) discusses $G = G^\circ + RT \ln Q$.

- **Gibbs-Helmholtz equation:** An equation for the temperature dependence of G . *Given by*

$$\left(\frac{\partial G/T}{\partial T}\right)_P = -\frac{H}{T^2}$$

– Derivation:

$$\begin{aligned} G &= H - TS \\ \frac{G}{T} &= \frac{H}{T} - S \\ \left(\frac{\partial G/T}{\partial T}\right)_P &= \left[\frac{\partial}{\partial T}\left(\frac{1}{T} \cdot H\right)\right]_P - \left(\frac{\partial S}{\partial T}\right)_P \\ &= \left[\frac{\partial}{\partial T}\left(\frac{1}{T}\right) \cdot H\right]_P + \frac{1}{T} \left(\frac{\partial H}{\partial T}\right)_P - \left(\frac{\partial S}{\partial T}\right)_P \\ &= -\frac{H}{T^2} + \frac{C_P}{T} - \left(\frac{\partial S}{\partial T}\right)_P \\ &= -\frac{H}{T^2} \end{aligned}$$

- McQuarrie and Simon (1997) discusses the graph of $G(T)$.

– It offers an alternate explanation for continuity, namely that since $\Delta_{\text{trs}}S = \Delta_{\text{trs}}H/T_{\text{trs}}$, then

$$0 = \Delta_{\text{trs}}H - T_{\text{trs}}\Delta_{\text{trs}}S = \Delta_{\text{trs}}G$$

- McQuarrie and Simon (1997) discusses **fugacity**^[1].

¹Come back after reading Chapter 16.

Part III

Applications of Thermodynamics and Statistical Mechanics

Week 6

Deviations from Ideality

6.1 Thermodynamics of the Rubber Band

2/14:

- Midterm:
 - Take for two hours, any two hours tomorrow.
 - Upload a single PDF file for your answer.
 - Do not discuss the questions with anybody.
- **Standard Gibbs free energy** (at T): The following energy, where $H^\circ(T)$ is the standard enthalpy at T and $S^\circ(T)$ is the standard entropy at T . Denoted by $G^\circ(T)$. Given by

$$G^\circ(T) = H^\circ(T) - TS^\circ(T)$$

- It follows that since the enthalpy is taken at constant pressure and the entropy $S(T, P)$ at non-standard pressure is given by $S(T, P) = S^\circ(T) + R \ln P/P_0$ that the Gibbs free energy $G(T, P)$ at nonstandard pressure is

$$\begin{aligned} G(T, P) &= H^\circ(T) - TS(T, P) \\ &= H^\circ(T) - T \left(S^\circ(T) + R \ln \frac{P}{P_0} \right) \\ &= G^\circ(T) - RT \ln \frac{P}{P_0} \end{aligned}$$

- Gibbs free energy and equilibrium at constant T, P .

- If



is in equilibrium where A, B, C, D are ideal gases, then $\Delta G = 0$.

- This implies that

$$a\Delta G_a + b\Delta G_b = c\Delta G_c + d\Delta G_d$$

which is the **law of mass action**.

- Example of phase equilibrium ($G(T)$ for solid/liquid/gas phases).

- We have from the total differential of G that

$$\left(\frac{\partial G}{\partial T} \right)_P = -S$$

- Since $S \geq 0$ always, G is monotonically decreasing.

- During the liquid phase, there is a relatively constant slight negative slope in the $G(T)$ graph.
- During the gas phase, S is much bigger, so there is a larger negative slope in the $G(T)$ graph.
- Additionally, at the heats of vaporization and fusion, the system is in equilibrium (hence the energies are the same), so the graph is continuous.
- Rubber band temperature analysis.
 - A rubber band heats up when stretched:
 - We have that $dU = T dS + f dL$.
 - We want to show that $(\partial U / \partial L)_T = T(\partial S / \partial L)_T + f$.

$$\begin{aligned} dU &= T \left(\frac{\partial S}{\partial T} \right)_L dT + \left(\frac{\partial S}{\partial L} \right)_T dL + f dL \\ &= T \left(\frac{\partial S}{\partial T} \right)_L dT + \left[\left(\frac{\partial S}{\partial L} \right)_T + f \right] dS \end{aligned}$$

- All that's left is to show that $-(\partial S / \partial L)_T = (\partial f / \partial T)_L$, which we can do using Maxwell relations.

$$dA = -S dT + f dL$$

$$\left(\frac{\partial U}{\partial L} \right)_T = -T \left(\frac{\partial f}{\partial T} \right)_L + f$$

- Stating the equation of state for the “ideal” rubber band.

$$f = T\phi(L)$$

- It follows that $(\partial f / \partial T)_L = f/T$, and $(\partial U / \partial L)_T = -T \cdot f/T + f = 0$.
- We are now ready to answer the question of does it cool down or heat up when stretched adiabatically.

$$\begin{aligned} dU &= \left(\frac{\partial U}{\partial L} \right)_T dL + \left(\frac{\partial U}{\partial T} \right)_L dT = \delta q + f dL \\ \left(\frac{\partial U}{\partial T} \right)_L dT &= f dL \\ C_L dT &= f dL \\ dT &= \frac{f}{C_L} dL > 0 \end{aligned}$$

so since $dT > 0$, the rubber band heats up as it stretches.

- For intuitive motivation, PGS discusses Figure 17.1 of Labalme (2021a).
- Building a statistical model of the rubber band.
 - Consider the rubber band to be made up of segments (you can think of each segment as part of a polymer). These segments can be oriented up or down. In a stretched rubber band, the polymers will be straight, i.e., all the segments will point the same way.
 - The difference in energy ΔE between a segment (of length ℓ_0) pointing up or down is $2f$.
 - Thus, the partition function for each segment is

$$q = e^{-f\ell_0/k_B T} + e^{f\ell_0/k_B T}$$

- Note that this is the same as the partition function for a paramagnet (which can also either be up or down) except that $f\ell_0$ becomes $\mu_B B_z$.

- When we sum the energies, we multiply the component partition functions. Thus,

$$Q = q^N$$

- We know that

$$L = N(p + p_0 + p - (-p_0)) = Np_0(p + (-p)) = N\ell_0 \tanh\left(\frac{f\ell_0}{k_B T}\right)$$

6.2 Van der Waals Phase Transitions

2/16:

- Wrapping up the rubber band analysis.
- Note that we can approximate $\tanh x \approx x$ for small x . Thus,

$$L = N\ell_0 \cdot \frac{f\ell_0}{k_B T}$$

for small f .

- Therefore, statistical mechanics predicts the rubber band “ideal” equation of state, $f = T\phi(L)$, where f is the stretching force. Note that this is not at all like a mechanical spring constant; it is purely an entropy effect.

$$- dS = dS_{\text{orientation}} + dS_{\text{thermal}} = dS_{\text{orientation}} + C/T dT.$$

$$- dU = C dT + T dS = 0, \text{ implying that } dT = -T/C dS.$$

- Adiabatic stretching decreases orientation entropy and increases the temperature.
- The same model predicts the Curie law of paramagnetism, $M = CB/T$.
 - We have

$$M = N\mu \tanh \frac{\mu B}{k_B T}$$

$$M = \frac{N\mu^2}{k_B} \frac{B}{T}$$

where the second equation only holds for small B .

- Adiabatic demagnetization allows you to go from 4 K to 1 K using He_4 to He_3 liquefaction. Adiabatic demagnetization increases spin entropy and reduces the temperature.
 - To achieve 4 K, we let He_4 adiabatically expand, which makes it very cold. However, at temperatures below 4 K, He_4 is a liquid and it can no longer adiabatically expand like a gas.
 - To achieve temperatures lower than 4 K, you stick copper salt in a cryostat, subject it to a magnetic field, and cool it to 4 K with the above method. The magnet aligns the spins. When you take the salt out of the magnetic field, the spins will randomize entropically, but this takes energy. You use the lattice energy to raise the spin energy.
- The challenge of protein folding and structure determination.
 - We have $\Delta G = \Delta(H - TS)$.
 - Protein folding necessitates $\Delta S < 0$ (you are creating order). Thus, we must have $\Delta H < \Delta(TS) = T\Delta S$, i.e., the protein must adopt a very, very stable conformation.
 - Estimating ΔS :

$$\Delta S = k_B \ln W = k_B \ln 3^{100} = 914 \text{ J K}^{-1} \text{ mol}^{-1}$$

- Using **Levinthal's paradox**, we can estimate the number of possible configurations of a protein.
- Each segment (essentially a C–C bond) has about three dihedral angle possibilities for the segments at either end.
- Thus, for a protein that is 100 segments long, $W \approx 3^{100}$.
- At 300 K, $T\Delta S = 274 \text{ kJ/mol}$.
 - It doesn't take much of a temperature difference to alter or prevent protein folding.
- About 15 hydrogen bonds or a single disulfide bond is about 274 kJ/mol of energy, so we can find ways to stabilize proteins over the entropic barriers.
- Just recently, a UChicago grad student (who had been working with a UChicago professor who's been studying the protein folding problem for a long time) headed a team at Google that has an AI that looks like it will be able to solve the protein folding problem.
- **Levinthal's paradox:** The observation that finding the native folded state of a protein by a random search among all possible configurations can take an enormously long time. Yet proteins can fold in seconds or less.
- Real gases deviate from ideality due to molecular interactions.
- **Compressibility factor:** Denoted by z . Given by

$$z = \frac{P\bar{V}}{RT}$$

- The van der Waals equation of state (where \bar{V} is the molar volume):

$$\left(P + \frac{a}{\bar{V}^2}\right)(\bar{V} - b) = RT$$

- We can also rewrite this as

$$P = \frac{RT}{\bar{V} - b} - \frac{a}{\bar{V}^2}$$

- Note that $a, b \geq 0$.
- The compressibility is

$$z = \frac{\bar{V}}{\bar{V} - b} - \frac{a}{RT\bar{V}}$$

- The van der Waals equation of state is cubic in \bar{V} and can predict two different molar volumes (gas and liquid) for one pressure.



Figure 6.1: The van der Waals isotherm of CO_2 at 0°C .

- We have that

$$\begin{aligned}
 RT &= \left(P + \frac{a}{\bar{V}^2} \right) (\bar{V} - b) \\
 &= P\bar{V} - bP + \frac{a}{\bar{V}} - \frac{ab}{\bar{V}^2} \\
 \frac{RT\bar{V}^2}{P} &= \bar{V}^3 - b\bar{V}^2 + \frac{a}{P}\bar{V} - \frac{ab}{P} \\
 0 &= \bar{V}^3 - \left(b + \frac{RT}{P} \right) \bar{V}^2 + \frac{a}{P}\bar{V} - \frac{ab}{P}
 \end{aligned}$$

- Since it is cubic in \bar{V} , this means we can have up to three different molar volumes for one pressure.
- This also reflects the fact that as we compress a gas, it behaves ideally for a while, and then pressure is constant as condensation takes hold, and then we must apply massive amounts of pressure to make the volume any smaller.
- On \overline{AG} , the system is a gas. On \overline{LD} , it is a liquid. On \overline{AD} , pressure is constant as we compress more and more because condensation takes hold, so highly compressed gas molecules become liquid, reducing the pressure.
- Drawing the line \overline{AD} : The points at which gas stops and liquid starts must be at the same pressure.
- We must also have the area above and below the line equal (**Maxwell equal area contraction**).
 - Phase equilibrium (like we have along \overline{AD}) means that free energy does not change. Mathematically, $\Delta G(A) = \Delta G(D)$.
 - If A and D represent gas and liquids at the same temperature, they are in equilibrium and they must have the same molar Gibbs free energy.
 - We have $dG = -S dT + V dP$.
 - Since T is constant (this is an isotherm), $dG = V dP$.
 - Thus, we can integrate along the curve \overline{DA} to find an appropriate \overline{DA} such that the integral is zero.

$$\begin{aligned}
 0 &= \Delta G(A) - \Delta G(D) \\
 &= \int_D^A dG \\
 &= \int_D^A V dP
 \end{aligned}$$

- See Problem 23-46 for the “Maxwell equal area construction rule.”

6.3 More van der Waals Phenomena

- 2/18:
- The cubic vdW equation of state predicts a critical point at T_c at and above which only one P for (V, T) solution is possible.
 - The behavior of isotherms around the critical temperature.
 - For isotherms below the critical temperature, there will be a range of volumes where the gas and liquid phases are in equilibrium.
 - For isotherms above the critical temperature, we only have the gas phase, so there is no region of constant pressure as we compress the system.
 - This means that at the critical temperature $T = T_c$, there will only be a single volume V_c and hence pressure P_c at which the gas and liquid phases are in equilibrium. Mathematically, this means that each of the three roots of the cubic van der Waals equation exist at the same point V_c , i.e., that the equation is of the form $(V - \bar{V}_c)^3$.

- Expanding, we have

$$\begin{aligned} 0 &= \bar{V}^3 - \left(b + \frac{RT}{P}\right) \bar{V}^2 + \frac{a}{P} \bar{V} - \frac{ab}{P} \\ &= \bar{V}^3 - 3\bar{V}_c \bar{V}^2 + 3\bar{V}_c^2 \bar{V} - \bar{V}_c^3 \end{aligned}$$

- Thus, at T_c ,

$$b + \frac{RT_c}{P_c} = 3\bar{V}_c \qquad \frac{a}{P_c} = 3\bar{V}_c^2 \qquad \frac{ab}{P_c} = \bar{V}_c^3$$

- One immediate consequence is that

$$\begin{aligned} \bar{V}_c^3 &= \frac{a}{P_c} \cdot b \\ \bar{V}_c^3 &= 3\bar{V}_c^2 b \\ \bar{V}_c &= 3b \end{aligned}$$

- Thus, the critical volume is on the order of magnitude of the molecular volume.
- Note that we can't manipulate the first equation into a different relation between \bar{V}_c and b using the ideal gas law substitution because this is a van der Waals gas.
- It follows that

$$\begin{aligned} 3\bar{V}_c^2 &= \frac{a}{P_c} & b + \frac{RT_c}{P_c} &= 3\bar{V}_c \\ P_c &= \frac{a}{3(3b)^2} & b + RT_c \cdot \frac{27b^2}{a} &= 3(3b) \\ P_c &= \frac{a}{27b^2} & T_c &= \frac{8a}{27Rb} \end{aligned}$$

- Thus,

$$z_c = \frac{P_c V_c}{RT_c} = \frac{3}{8}$$

- Indeed, the vdW equation predicts the compressibility at the critical point to be 3/8. The Redlich-Kwong predicts 1/3. The Peng-Robinson predicts 0.307. The experimental values are around 0.3. Water and ammonia deviate significantly: 0.23 and 0.24, respectively, due to their strong dipole moments/hydrogen bonding. Table 16.5 gives a lot of related data.
- Also, at T_c, V_c , we have that $(\partial P / \partial V)_T = 0$ and $\kappa \rightarrow \infty$ (recall that κ is the isothermal compressibility).
 - It follows from the vdW that $\kappa \propto (\bar{V} - \bar{V}_c)^{-1}$.
 - Experimental value: *Every* gas satisfies $\kappa \propto (\bar{V} - \bar{V}_c)^{-1.24}$.
 - The mystery was solved theoretically in the 1970s with renormalization group theory.
- Law of corresponding states.

- We define the relative pressure, volume, and temperature by

$$P_R = \frac{P}{P_c} \qquad V_R = \frac{V}{V_c} \qquad T_R = \frac{T}{T_c}$$

- Note that

$$\frac{RT}{P_c V_c} = \frac{RT_c}{P_c V_c} \frac{T}{T_c} = \frac{1}{z_c} \frac{T}{T_c} = \frac{8}{3} T_R$$

– Thus,

$$\begin{aligned} RT &= \left(P + \frac{a}{\bar{V}^2} \right) (\bar{V} - b) \\ &= \left(P + \frac{3\bar{V}_c^2 P_c}{\bar{V}_c^2} \right) \left(\bar{V} - \frac{\bar{V}_c}{3} \right) \\ \frac{RT}{P_c \bar{V}_c} &= \left(\frac{P}{P_c} + \frac{3\bar{V}_c^2}{\bar{V}} \right) \left(\frac{\bar{V}}{\bar{V}_c} - \frac{1}{3} \right) \\ \frac{8}{3} T_R &= \left(P_R + \frac{3}{\bar{V}_R^2} \right) \left(\bar{V}_R - \frac{1}{3} \right) \end{aligned}$$

- Virial expansion to experimentally determine the vdW coefficients a, b from the compressibility near ideal conditions.

– We let

$$z = 1 + \frac{B_{2V}(T)}{\bar{V}} + \frac{B_{3V}(T)}{\bar{V}^2} + \dots$$

where $B_{iV}(T)$ is the i^{th} **virial coefficient**.

– It follows that

$$\begin{aligned} \frac{P\bar{V}}{RT} &= \frac{\bar{V}}{\bar{V} - b} - \frac{a}{RT\bar{V}} \\ &= \frac{1}{1 - b/\bar{V}} - \frac{a}{RT\bar{V}} \\ &= 1 + \left(\frac{b}{\bar{V}} - \frac{a}{RT\bar{V}} \right) + \text{terms in } \frac{1}{\bar{V}^2} + \dots \end{aligned}$$

where we get from the second to the third line using the expansion

$$\frac{1}{1-x} = 1 + x + x^2 + \dots$$

– Thus,

$$B_{2V}(T) = b - \frac{a}{RT}$$

- Microscopic origin of the vdW coefficients from the interaction potential.

– Draws the interaction potential for a diatomic molecule.

■ The repulsion comes from the Fermi exclusion principle.

– Discusses dipole-induced dipole moments.

■ $-1/2 \propto E^2 \propto 1/r^6$.

– The origin of vdW (or London dispersion) interaction is quantum mechanical.

■ We use perturbation theory to calculate the interaction between two neighboring quantum dipoles.

$$\Delta E^{(1)} = \langle \psi_A^\circ \psi_B^\circ | H_{int} | \psi_A^\circ \psi_B^\circ \rangle = 0$$

■ Thus we need the second order correction:

$$\Delta E^{(2)} = - \sum_{ij} \frac{|\langle \psi_A^\circ \psi_B^\circ | H_{int} | \psi_A^i \psi_B^j \rangle|^2}{E_a^A + E_j^B - (E_0^A + E_0^B)} \approx - \frac{c}{r^6}$$

where $c \geq 0$.

6.4 Office Hours (PGS)

- Integrating along the curve in Figure 6.1?
 - Integrating along the curve with respect to P means (geometrically) that we take the area “under” (to the left of) the curve from A to C , then subtract the area under from C to B , and then add the area under from B to A . This calculation gets us overall the area of the bottom chunk as negative and the area of the top chunk as positive.

6.5 Chapter 16: The Properties of Gases

From McQuarrie and Simon (1997).

2/21:

- Outline:
 - Having studied ad nauseam the properties of individual atoms and molecules, we now turn our attention to systems consisting of large numbers of atoms and molecules.
 - The ideal gas equation is discussed, the van der Waals equation as the most famous extension of it, and then **virial expansions** as an even more systematic and accurate approach.
 - Relating virial coefficients to the intermolecular interaction energy reveals that deviations from ideal gas behavior provide great insight into molecular interactions.
- **Virial expansion:** An expression for the pressure of a gas as a polynomial in the density.
- **Ideal-gas equation of state:** An equation of state relating the pressure, volume, and temperature of a gas that is sufficiently dilute (i.e., with particles sufficiently far apart) for the intermolecular interactions to be negligible. *Given by*

$$PV = nRT$$

$$P\bar{V} = RT$$

- **Ideal gas:** A gas that obeys the ideal-gas equation of state.
 - All gases, regardless of the shape or size of the molecules or how the molecules interact with each other, obey behave ideally if they are sufficiently dilute.
 - “Experimentally, most gases satisfy [the ideal-gas equation of state] to approximately 1% at one atm and 0 °C” (McQuarrie & Simon, 1997, p. 638).
- **Extensive quantity:** A quantity that is directly proportional to the size of the system in question. *Also known as* **extensive variable**.
- **Intensive quantity:** A quantity that does not depend on the size of the system in question. *Also known as* **intensive variable**.
- “If we divide an extensive quantity by the number of particles or the number of moles in a system, we obtain an intensive quantity” (McQuarrie & Simon, 1997, p. 638).
- **Liter:** A unit of volume. *Denoted by* **L**. *Given by*

$$1 \text{ L} = 1 \text{ dm}^3$$

- The SI unit of volume is m^3 , but liters are an acceptable unit of volume in the IUPAC system.

- **Pascal:** The SI unit of pressure. *Denoted by* **Pa**. *Given by*

$$1 \text{ Pa} = 1 \frac{\text{N}}{\text{m}^2} = 1 \frac{\text{kg}}{\text{m s}^2}$$

- “Pressure can be measured experimentally by observing how high a column of liquid (usually mercury) is supported by the gas” (McQuarrie & Simon, 1997, p. 638).

- If m is the mass of the liquid, g is the **gravitational acceleration constant**, A is the base area of the column, ρ is the density of the liquid, and h is the height of the column, then

$$P = \frac{F}{A} = \frac{mg}{A} = \frac{\rho h A g}{A} = \rho h g$$

- **Gravitational acceleration constant:** The acceleration of an object due to gravity at the surface of the Earth. *Denoted by g . Given by*

$$g = 9.8067 \frac{\text{m}}{\text{s}^2}$$

- The pressure exerted by a 76.000 cm column of mercury ($\rho(\text{Hg}) = 13.596 \text{ g cm}^{-3}$) is

$$\begin{aligned} P &= \rho h g \\ &= 1.0133 \times 10^6 \frac{\text{g}}{\text{cm s}^2} \\ &= 1.0133 \times 10^5 \frac{\text{kg}}{\text{m s}^2} \\ &= 1.0133 \times 10^5 \text{ Pa} \\ &= 101.33 \text{ kPa} \end{aligned}$$

- **Atmosphere:** The pressure that supports a 76.0 cm column of mercury. *Denoted by **atm**. Given by*

$$1 \text{ atm} = 101.325 \text{ kPa}$$

- **Bar:** The SI standard of pressure. *Denoted by **bar**. Given by*

$$1 \text{ bar} = 1 \times 10^5 \text{ Pa}$$

- **Torr:** The pressure that supports a 1.00 mm column of mercury. *Denoted by **torr**. Given by*

$$1 \text{ torr} = \frac{1}{760} \text{ atm}$$

- “Because we are experiencing a transition period between the widespread use of atm and torr on the one hand and the future use of bar and kPa on the other hand, students of physical chemistry must be proficient in both sets of pressure units” (McQuarrie & Simon, 1997, p. 639).
- Temperatures as low as $1 \times 10^{-7} \text{ K}$ and as high as $1 \times 10^8 \text{ K}$ have been achieved in the laboratory.
- **Triple point:** The point (P, V, T) at which all three of a given substance’s phases are in equilibrium.
- **Kelvin:** The SI unit of temperature, defined as $1/273.16$ of the temperature of the triple point of water. *Denoted by **K**.*
- “We will use the lower case t for $^{\circ}\text{C}$ and the upper case T for K ” (McQuarrie & Simon, 1997, p. 640).
- **Room temperature:** The temperature $25^{\circ}\text{C} = 298.15 \text{ K}$.
- If we plot $P\bar{V}$ vs. P for several gases at $T = 273.15 \text{ K}$, all the data can be extrapolated to a common value of $P\bar{V} = 22.414 \text{ L atm}$ as $P \rightarrow 0$. Thus,

$$\begin{aligned} R &= \frac{P\bar{V}}{T} = \frac{22.414 \text{ L atm}}{273.15 \text{ K}} = 0.082058 \frac{\text{L atm}}{\text{mol K}} \\ &= 8.3145 \frac{\text{J}}{\text{mol K}} \\ &= 0.083145 \frac{\text{L bar}}{\text{mol K}} \end{aligned}$$

- **Compressibility factor:** The following quantity. *Denoted by **Z**. Given by*

$$Z = \frac{P\bar{V}}{RT}$$



(a) Changing the gas type.

(b) Changing the gas temperature.

Figure 6.2: Plots of the compressibility factor vs. pressure.

- We use the compressibility factor, which is constant at 1 for an ideal gas, to visualize deviations from ideality for different gases (Figure 6.2a) at different temperatures (Figure 6.2b).
- At lower temperatures, intermolecular attraction takes hold, reducing the true volume and making $Z < 1$.
- At higher temperatures, molecules are moving fast enough to make negligible their attractions. Here, only repulsions due to their nonzero volume take hold at higher pressures.
- “The closer the gas is to the point at which it begins to liquify, the larger the deviations from ideal behavior will be” (McQuarrie & Simon, 1997, p. 642).
- **Van der Waals equation:** The most well known of the gaseous equations of state which account for intermolecular interactions. *Given by*

$$\left(P + \frac{a}{\bar{V}^2}\right)(\bar{V} - b) = RT$$

- Note that as $\bar{V} \rightarrow \infty$, the van der Waals equation becomes the ideal-gas equation of state.
- The van der Waals equation is also commonly written in the following form.

$$P = \frac{RT}{\bar{V} - b} - \frac{a}{\bar{V}^2}$$

- **Van der Waals constants:** The constants a, b in the van der Waals equation, whose value depend upon the particular gas.
 - We will see later that “the value of a reflects how strongly the molecules of a gas attract each other and the value of b reflects the size of the molecules” (McQuarrie & Simon, 1997, p. 643).
 - Table 16.3 lists van der Waals constants for various substances.
- The van der Waals equation approximates the behavior depicted in Figure 6.2 since

$$Z = \frac{P\bar{V}}{RT} = \frac{\bar{V}}{\bar{V} - b} - \frac{a}{RT\bar{V}}$$

- At high pressures ($\bar{V} \rightarrow 0$), the first term above dominates because $\bar{V} - b$ gets really small.
- At low pressures, the second term dominates.
- Notice that this also relates to the previously mentioned roles of a, b (e.g., at high pressures, the size of the molecules [and b] becomes significant).

- To solve the van der Waals equation for a molar volume given the other state variables, we must find the roots of a cubic equation via Newton's method.

$$\bar{V}^3 - \left(b + \frac{RT}{P}\right)\bar{V}^2 + \frac{a}{P}\bar{V} - \frac{ab}{P} = 0$$

- **Redlich-Kwong equation:** A relatively simple equation of state that is much more accurate than the van der Waals equation. *Given by*

$$P = \frac{RT}{\bar{V} - B} - \frac{A}{\bar{V}(\bar{V} + B)\sqrt{T}}$$

- Table 16.4 lists Redlich-Kwong parameters for various substances.
- As a cubic equation (Problem 16-26):

$$\bar{V}^3 - \frac{RT}{P}\bar{V}^2 - \left(B^2 + \frac{BRT}{P} - \frac{A}{P\sqrt{T}}\right)\bar{V} - \frac{AB}{P\sqrt{T}} = 0$$

- **Peng-Robinson equation:** Another relatively simple equation of state that is much more accurate than the van der Waals equation. *Given by*

$$P = \frac{RT}{\bar{V} - \beta} - \frac{\alpha}{\bar{V}(\bar{V} + \beta) + \beta(\bar{V} - \beta)}$$

- α is a somewhat complicated function of temperature, so α, β values are not listed.
- Problem 16-28 gives the Peng-Robinson equation as a cubic.
- As a general rule, the van der Waals equation fails beyond 200 bar while the Redlich-Kwong and Peng-Robinson equations remain nearly quantitative all the way into regions where the gas liquifies.
 - Additionally, the Peng-Robinson equation is better in the liquid-vapor phase transition region, and the Redlich-Kwong equation is better at higher pressures (these equations were actually constructed for these purposes).
- There exist many more sophisticated equations of state, some containing 10+ parameters, that can quantitatively reproduce experimental data over a wide range of state variables.
- A note on the drawing of Figure 6.2.
 - The curves in the figure were plotted with the Redlich-Kwong equation of state (since it is the one that most accurately depicts reality at high pressures).
 - To plot said curves, we need $Z(P)$. Much like with the van der Waals equation, we can write that

$$Z = \frac{P\bar{V}}{RT} = \frac{\bar{V}}{\bar{V} - B} - \frac{A\bar{V}}{RT\bar{V}(\bar{V} + B)\sqrt{T}}$$

- However, the above equation expresses Z as a function of \bar{V} , not P . Fortunately, we can solve the Redlich-Kwong equation for P using Cardano's formula.
- Comparing the cubic form of the Redlich-Kwong equation to the general cubic $ax^3 + bx^2 + cx + d = 0$ yields

$$a = 1 \quad b = -\frac{RT}{P} \quad c = \frac{A}{P\sqrt{T}} - \frac{BRT}{P} - B^2 \quad d = -\frac{AB}{P\sqrt{T}}$$

- We now depress $ax^3 + bx^2 + cx + d = 0$ into $t^3 + pt + q = 0$ via the substitutions

$$t = x + \frac{b}{3a} \quad p = \frac{3ac - b^2}{3a^2} \quad q = \frac{2b^3 - 9abc + 27a^2d}{27a^3}$$

- Since $4p^3 + 27q^2 > 0$ for all $P \in (0, 1000]$ (as we can verify graphically), the depressed cubic $t^3 + pt + q = 0$ has the real root

$$t = \sqrt[3]{-\frac{q}{2} + \sqrt{\frac{q^2}{4} + \frac{p^3}{27}}} + \sqrt[3]{-\frac{q}{2} - \sqrt{\frac{q^2}{4} + \frac{p^3}{27}}}$$

- It follows by returning the substitution for x that the original cubic has the real root

$$x = -\frac{b}{3a} + \sqrt[3]{-\frac{q}{2} + \sqrt{\frac{q^2}{4} + \frac{p^3}{27}}} + \sqrt[3]{-\frac{q}{2} - \sqrt{\frac{q^2}{4} + \frac{p^3}{27}}}$$

i.e., that the above equation provides values \bar{V} that make the Redlich-Kwong equation true for arbitrary P . In other words, the above equation equals $\bar{V}(P)$.

- It is a simple matter then to consider $Z(\bar{V}(P))$.

Week 7

Phase Diagrams and Critical Points

7.1 Virial Coefficients and Fugacity

2/21: • Relation between the interaction potential and the first virial coefficient (Equation 16.25).

- Statistical mechanics is important because it gives us the relation

$$B_{2V}(T) = -2\pi N_A \int_0^\infty \left(e^{-u(r)/k_B T} - 1 \right) r^2 dr$$

- We can derive this with our knowledge of statistical mechanics, but PGS will not go through this.
- Now recall the Lennard-Jones potential

$$u(r) = 4\epsilon \left[\left(\frac{\sigma}{r} \right)^{12} - \left(\frac{\sigma}{r} \right)^6 \right]$$

- Note that the minimum is at $(2^{1/6}\sigma, -\epsilon)$.
- The limiting case of the Lennard-Jones potential is hard sphere repulsion (the repulsion as you approach a hard sphere, which is zero up until you're at the surface and then infinite repulsion). Thus, with no intermolecular attraction, $a = 0$, so in this case,

$$\begin{aligned} B_{2V}(T) &= -2\pi N_A \int_0^\sigma (-1) r^2 dr \\ &= \frac{2\pi N_A \sigma^3}{3} \\ &= b - \frac{0}{RT} \end{aligned}$$

- Now consider a potential that is van der Waals (c/r^6) up until a point and then hard sphere. In this case,

$$\begin{aligned} B_{2V}(T) &= \frac{2\pi N_A \sigma^3}{3} - 2\pi N_A \int_\sigma^\infty \left(e^{-c/r^6 k_B T} - 1 \right) r^2 dr \\ &= \frac{2\pi N_A \sigma^3}{3} - 2\pi N_A \int_\sigma^\infty \left(-\frac{c}{r^6 k_B T} \right) r^2 dr \\ &= b + \frac{2\pi N_A c}{k_B T} \cdot -\frac{1}{3\sigma^3} \end{aligned}$$

where we have used $e^x = 1 + x + \dots$ to get from the first line to the second.

■ Therefore,

$$a = \frac{2\pi N_A^2}{3} \frac{c}{\sigma^2}$$

- Derivation of the relation between $B_{2V}(T)$ and the interaction potential $u(r)$.

- Consider a system of independent, indistinguishable particles.
- The total Hamiltonian for the system has a kinetic energy part and an interaction part.

$$\hat{H}(p_i, r_i) = \sum_i \frac{\hat{p}_i^2}{2m} + \sum_{i < j} \hat{u}(r_i, r_j)$$

- The kinetic part (which ignores intermolecular interactions) will lead to the ideal gas partition function. The nonideal part of the partition function will come from the interaction potentials. Mathematically,

$$\begin{aligned} Q &= \frac{1}{N!} \int e^{-\beta E(p_i, r_i)} d^3p_i d^3r_i \\ &= \frac{1}{N!} \int \exp \left\{ -\beta \left[\sum_i \frac{p_i^2}{2m} + \sum_{i < j} u(r_i, r_j) \right] \right\} d^3p_i d^3r_i \\ &= \frac{1}{N!} \left(\int \exp \left[-\beta \sum_i \frac{p_i^2}{2m} \right] d^3p_i \right) \left(\int \exp \left[-\beta \sum_{i < j} u(r_i, r_j) \right] d^3r_i \right) \\ &= \underbrace{\frac{V^N}{N!} \left(\int \exp \left[-\beta \sum_i \frac{p_i^2}{2m} \right] d^3p_i \right)}_{Q_{\text{ideal}}} \cdot \underbrace{\frac{1}{V^N} \left(\int \exp \left[-\beta \sum_{i < j} u(r_i, r_j) \right] d^3r_i \right)}_{Q_u} \end{aligned}$$

- Define

$$f_{ij} = e^{-u(r_i, r_j)/k_B T} - 1$$

- Now note that the interaction between molecules is pretty small, and in fact $f_{ij} \rightarrow 0$ as $|r_i - r_j| \rightarrow \infty$.
- Thus,

$$\begin{aligned} Q_u &= \frac{1}{V^N} \int \exp \left[-\beta \sum_{i < j} u(r_i, r_j) \right] d^3r_i \\ &= \frac{1}{V^N} \int \prod_{i < j} (f_{ij} + 1) d^3r_i \end{aligned}$$

- We can do a **cluster expansion** on this small f_{ij} :

$$\prod_{i < j} (f_{ij} + 1) = 1 + \sum_{i < j} f_{ij} + \sum_{i < j} \sum_{k < \ell} f_{ij} f_{k\ell}$$

- In particular, $\sum_{i < j}$ is the sum of pairwise interactions while $f_{ij} \cdot f_{k\ell}$ are binary interactions, $f_{ij} f_{k\ell} f_{mn}$ are tertiary interactions, and so on and so forth.

- But at low density, the dominant term is the pairwise interaction so we have

$$\begin{aligned} Q_u &= \frac{1}{V^N} \int \left(1 + \sum_{i < j} f_{ij} \right) d^3 r_i \\ &= \frac{1}{V^N} \left(V^N + \frac{N(N-1)}{2} V^{N-2} \int f_{12} d^3 r_1 d^3 r_2 \right) \\ &= 1 + \frac{N(N-1)}{2V} \int (e^{-\beta u(r)} - 1) d^3 r \end{aligned}$$

- It follows that

$$Q = Q_{\text{id}} \left[1 + \frac{N(N-1)}{2V} \int (e^{-\beta u(r)} - 1) d^3 r \right]$$

- But now we need to extract an equation of state from the partition function. To do so, we use

$$\begin{aligned} P &= k_B T \left(\frac{\partial \ln Q}{\partial V} \right)_{N,T} \\ &= k_B T \left(\frac{\partial \ln Q_{\text{id}}}{\partial V} \right)_{N,T} + k_B T \left(\frac{\partial \ln Q_u}{\partial V} \right)_{N,T} \end{aligned}$$

- We know that the first term above is equal to $Nk_B T/V$, but it takes a bit more work for the second one.
- We have that

$$\begin{aligned} \ln Q_u &= \ln \left(1 + \underbrace{\frac{N(N-1)}{2V}}_{\substack{\text{Approximately} \\ \text{the intermolecular} \\ \text{distance } 1/\rho^3}} \underbrace{\int (e^{-\beta u(r)} - 1) d^3 r}_{\substack{\text{Approximately} \\ \text{the molecular} \\ \text{volume } a^3}} \right) \\ &= \frac{N(N-1)}{2V} \int (e^{-\beta u(r)} - 1) d^3 r \end{aligned}$$

since the second term is a small number and the natural log of a small number plus 1 is approximately that small number.

- Thus,

$$\left(\frac{\partial \ln Q_u}{\partial V} \right)_{N,T} = -\frac{N(N-1)}{2V^2} \int (e^{-\beta u(r)} - 1) d^3 r$$

so

$$\begin{aligned} P &= \frac{Nk_B T}{V} - \frac{Nk_B T}{V} \frac{N-1}{2V} \int (e^{-\beta u(r)} - 1) d^3 r \\ &= \frac{RT}{\bar{V}} - \frac{RT}{\bar{V}} \frac{N-1}{2V} \int (e^{-\beta u(r)} - 1) d^3 r \end{aligned}$$

- Consequently,

$$Z = \frac{P\bar{V}}{RT} = 1 - \frac{N_A}{\bar{V}} \cdot \frac{1}{2} \int (e^{-\beta u(r)} - 1) d^3 r$$

- Therefore, by comparison with the virial expansion,

$$\begin{aligned} B_{2V}(T) &= -\frac{N_A}{2} \int (e^{-\beta u(r)} - 1) d^3 r \\ &= -\frac{N_A}{2} \int_0^\infty (e^{-\beta u(r)} - 1) 4\pi r^2 dr \\ &= -2\pi N_A \int_0^\infty (e^{-u(r)/k_B T} - 1) r^2 dr \end{aligned}$$

- Free energy as a function of (T, P) for a real gas. Definition of fugacity and fugacity coefficients.
 - We want to find $\Delta G(T, P)$.
 - We have that $d\bar{G} = -\bar{S}dT + \bar{V}dP$. It follows that

$$\left(\frac{\partial \bar{G}}{\partial P}\right)_T = \bar{V}$$

- Thus,

$$\bar{G}(T, P) = \bar{G}(T, P_0) + \int_{P_0}^P \bar{V} dP$$

- In the ideal case,

$$\begin{aligned}\bar{G}_{\text{ideal}}(T, P) &= \bar{G}_{\text{ideal}}(T, P_0) + \int_{P_0}^P \frac{RT}{P} dP \\ &= \bar{G}_{\text{ideal}}(T, P_0) + RT \ln \frac{P}{P_0}\end{aligned}$$

- In the nonideal case, we define a fugacity f by

$$\bar{G}_{\text{ideal}}(T, P) = \bar{G}_{\text{ideal}}(T, P_0) + RT \ln \frac{f}{P_0}$$

- The second term in the above equation refers to the Gibbs free energy of an ideal gas at $P_0 = 1$ bar or $P_0 = 1$ atm. Note that even at $P_0 = 1$ atm, however, there is too much pressure for truly ideal behavior, so $f \neq P_0$.
- Imagine that $\Delta \bar{G}_1$ takes us from a real gas at (T, P) to an ideal gas at (T, P) . Then

$$\begin{aligned}\Delta \bar{G}_1 &= \bar{G}_{\text{ideal}}(T, P) - \bar{G}_{\text{real}}(T, P) \\ &= \left[\bar{G}_{\text{ideal}}(T, P_0) + RT \ln \frac{P}{P_0} \right] - \left[\bar{G}_{\text{ideal}}(T, P_0) + RT \ln \frac{f}{P_0} \right] \\ &= -RT \ln \frac{f}{P}\end{aligned}$$

- Now let $\Delta \bar{G}_2$ take us from a real gas at (T, P) to a real gas at T and $P \rightarrow 0$, which will be the same as an ideal gas at T and $P \rightarrow 0$. Then let $\Delta \bar{G}_3$ take us from this ideal gas at T and $P \rightarrow 0$ to an ideal gas at (T, P) . It follows that

$$\Delta \bar{G}_2 = - \int_{P \rightarrow 0}^P \bar{V} dP' \qquad \Delta \bar{G}_3 = \int_{P \rightarrow 0}^P \frac{RT}{P'} dP'$$

- Thus, since $\Delta \bar{G}_1 = \Delta \bar{G}_2 + \Delta \bar{G}_3$ (G is a state function),

$$-RT \ln \frac{f}{P} = \int_{P \rightarrow 0}^P \left(-\bar{V} + \frac{RT}{P'} \right) dP'$$

- We then define γ to be the **fugacity coefficient** by $\gamma = f/P$. It follows that

$$\ln \gamma = \int_0^P \frac{z - 1}{P'} dP'$$

- Fugacity coefficient expressed in terms of the compressibility deviation from unity.
 - At low temperature, $z < 1$, so $\gamma < 1$ and hence $f < p$.
 - At high pressure, $z > 1$ (excluded volume), so $\gamma > 1$ and hence $f > P$.
- Introduces phase diagrams and their notable properties.

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