

# Thermodynamics II

For an introduction to classical thermodynamics, conduction, and radiation at an appropriate level, see chapters 3 and 4 of Wang and Ricardo, volume 2. For further detail on all the topics covered here, see chapters 10–16, 18, 21, and 23 of Blundell and Blundell. For interesting discussion, see chapters I-44 through I-46 of the Feynman lectures. There is a total of **82** points.

## 1 The First Law

In this section we'll cover the first law of thermodynamics, introducing a little multivariable calculus. This is intended to give you a little practice thinking of thermodynamic systems more generally; however, the material here is not required to do the rest of the problem set.

### Idea 1

The first law states that

$$\Delta U = \Delta Q + \Delta W$$

where  $\Delta U$  is the change in internal energy,  $\Delta Q$  is the heat given to the system, and  $\Delta W$  is the work done on the system.

Since only  $U$  is an intrinsic property of the system itself, i.e. a state function, some resources make this distinction explicit by writing the differential version of the first law as

$$dU = dQ + dW.$$

The  $d$ 's indicate that  $Q$  and  $W$  aren't properties of the system like  $U$  is.

### Idea 2

For a gas, the state of the system is specified by the pressure  $P$ , volume  $V$ , and temperature  $T$ , and the three are related by the ideal gas law. The infinitesimal work done is

$$dW = -P dV.$$

Many thermodynamic systems will be similar: there will be a temperature, and then a pair of so-called “conjugate variables”. For example, a bubble is specified by the surface tension  $\gamma$ , area  $A$ , and temperature  $T$ , with

$$dW = \gamma dA$$

along with some relation between  $\gamma$ ,  $A$ , and  $T$ . In general, this relation is called the system's “equation of state”. Another example would be a rubber band, which is described by its tension  $F$ , length  $L$ , and temperature  $T$ , with

$$dW = F dL$$

and a appropriate equation of state.

Therefore, for simple thermodynamic systems, you need two parameters to describe the state, since you start with three and can solve for one using the equation of state. This means that changes in the state should be described by partial derivatives, generalizations of derivatives which specify which parameter is being held constant.

This is not fundamentally different than what you did in introductory thermodynamics. For example, if you talked about the work done during isothermal expansion, you meant  $(\partial W/\partial V)|_T$  where the subscript  $T$  indicates the temperature is held constant.

### Example 1: Partial Derivatives

Consider describing a plane with Cartesian coordinates  $(x, y)$  and polar coordinates  $(r, \theta)$ . Calculate  $(\partial x/\partial r)|_\theta$  and  $(\partial x/\partial r)|_y$ .

### Solution

To evaluate the first partial derivative, we need to write  $x$  as a function of  $r$  and  $\theta$ . This is done by  $x = r \cos \theta$ . Differentiating with respect to  $r$  and treating  $\theta$  as a constant, we have

$$\left. \frac{\partial x}{\partial r} \right|_\theta = \cos \theta.$$

To evaluate the second, we should write  $x$  in terms of  $r$  and  $y$ . This is done starting with

$$r^2 = x^2 + y^2.$$

We could solve for  $x$  here. However, it's quicker to use the usual trick of implicit differentiation: note that differentiating both sides with respect to  $r$ , at fixed  $y$ , gives

$$2r = 2x \left. \frac{\partial x}{\partial r} \right|_y + 0$$

which yields the answer,

$$\left. \frac{\partial x}{\partial r} \right|_y = \frac{r}{x} = \frac{1}{\cos \theta}.$$

Note that these two results are different: changing what you keep constant changes the derivative. You can see this geometrically by evaluating the two versions of  $\partial x/\partial r$  at a few well-chosen points.

### Idea 3: Multivariable Differential

If the variables  $x$  and  $y$  may be varied independently, we may write

$$df = \left. \frac{\partial f}{\partial x} \right|_y dx + \left. \frac{\partial f}{\partial y} \right|_x dy$$

for a function  $f(x, y)$ . Conceptually, the change in  $f$  is built out of the changes resulting from changing  $x$  and  $y$  alone.

This formulation is useful in thermodynamics, because many things are given as differentials.

When  $f$  isn't a state function, we write the left-hand side as  $\bar{d}f$ .

### Example 2: Heat Capacities

The heat capacities at constant volume and constant pressure are defined as

$$C_V = \left. \frac{\partial Q}{\partial T} \right|_V, \quad C_P = \left. \frac{\partial Q}{\partial T} \right|_P.$$

Find a general formula for  $C_P - C_V$ , in terms of the variables  $P$ ,  $V$ , and  $T$ , and the internal energy  $U(P, V, T)$ .

### Solution

First off, note that by definition, we have

$$C_V = \left. \frac{\partial Q}{\partial T} \right|_V, \quad C_P = \left. \frac{\partial Q}{\partial T} \right|_P$$

as you can see by translating these partial derivatives into words. To compute  $C_P - C_V$ , we need another expression for  $\bar{d}Q$  to compare this to.

Using the equation of state, we can solve for  $P$  and write the internal energy as a function  $U(T, V)$ , where  $T$  and  $V$  can be varied independently. Then

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

Using the first law of thermodynamics, the fact that  $\bar{d}W = -P dV$ , and the definition of  $C_V$ , we have the expression

$$\bar{d}Q = C_V dT + \left( \left. \frac{\partial U}{\partial V} \right|_T + P \right) dV.$$

Here,  $\bar{d}Q$  is in terms of changes of  $T$  and  $V$ . However, in order to say anything useful about  $C_P$ , we need a differential in terms of changes of  $T$  and  $P$ .

This can be done by eliminating the volume  $V$ . Specifically, the equation of state lets us solve for  $V$  in terms of the other variables,  $V(T, P)$ . Then changes in volume are given by

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

which implies

$$\bar{d}Q = C_V dT + \left( \left. \frac{\partial U}{\partial V} \right|_T + P \right) \left( \left. \frac{\partial V}{\partial T} \right|_P dT + \left. \frac{\partial V}{\partial P} \right|_T dP \right).$$

Now  $C_P$  is the coefficient of  $dT$  in this expression, so reading it off gives the answer,

$$C_P - C_V = \left( \left. \frac{\partial U}{\partial V} \right|_T + P \right) \left. \frac{\partial V}{\partial T} \right|_P.$$

Intuitively, the extra heat we have to put in at constant pressure, when we allow the volume to change, comes from two sources. First, we have to make up for the energy lost to  $PdV$  work. Second, we have to make up for the energy intrinsically needed to expand the gas, by doing work against the attractive intermolecular forces between the gas molecules. This term is zero for an ideal gas by definition.

- [2] **Problem 1.** For an monatomic ideal gas, we have  $U = (3/2)nRT$  and the equation of state  $PV = nRT$ . Evaluate  $C_V$  and  $C_P$  using the results above.
- [3] **Problem 2.** In this problem we modify the ideal gas to account for a short-ranged attractive intermolecular force. This attraction yields a negative contribution to the energy, so that

$$U = \frac{3}{2}nRT - \frac{an^2}{V}.$$

- (a) Explain why the change in energy is proportional to  $n^2/V$ .
- (b) Argue that because of this extra term, the pressure is  $an^2/V^2$  less than it would be for an ideal gas. Hence conclude the equation of state is

$$\left(P + \frac{an^2}{V^2}\right)V = nRT.$$

- (c) Show that the heat capacities are

$$C_V = \frac{3}{2}nR, \quad C_P - C_V = nR + \frac{2an^2}{TV} + O(n^3)$$

in the limit where the gas is dilute.

In **T3**, we will combine this pressure correction with a “volume exclusion factor” to yield the van der Waals gas, which is useful for modeling the liquid-gas phase transition.

### Remark: Heat vs. Work

As long as we’re discussing fundamentals, we might as well ask: what is the real difference between work and heat? If we have a gas in a cylinder with a piston, moving the piston counts as work. Clearly, if we split the piston into two halves and moved them independently, this would still count as work.

We can keep splitting the piston until we are down to individual atoms of the piston. These can transfer energy to gas particles because they can slam into them as the gas particles are bouncing off, like a batter hitting a baseball. But as mentioned in **T1**, this is precisely how a hot wall transfers energy to a gas! So when did we transition from doing work to supplying heat?

The fact is that there is no absolute distinction. Heat and work are the two ways of transferring energy to a system. Work is just the subset of energy transfers that we bother to keep track of

in detail. As explained below, entropy quantifies our ignorance of the precise state a system is in, which is why heat (i.e. energy transfer we aren't keeping track of) can change it. That's also why free expansion can change the entropy, without either heat or work.

## 2 The Second Law

### Idea 4: Entropy

The increase in entropy of a body at temperature  $T$  that receives heat  $dQ$  is

$$dS = \frac{dQ}{T}$$

if the process is reversible; moreover, entropy is a state function.

### Idea 5: Second Law

The second law of thermodynamics takes several equivalent forms, which are all useful.

- For an isolated system,  $dS > 0$  for any spontaneous process.
- Clausius: no process is possible whose sole result is the transfer of heat from a colder to a hotter body.
- Kelvin: no process is possible whose sole result is the complete conversion of heat into work.
- Carnot: of all heat engines working between two given temperatures, none is more efficient than a Carnot engine.

### Example 3: Mixing Entropy

A container of volume  $V$  is divided in half by a partition. The two halves contain pure nitrogen and oxygen gas, respectively, both at pressure  $P$  and temperature  $T$ . The partition is removed and the gases are allowed to mix. How much does the entropy increase?

### Solution

We cannot apply the formula  $dS = dQ/T$ , because the process is not in equilibrium. However, because entropy is a state function, we can calculate its change by considering a path that only goes through equilibrium states.

Suppose we introduce a piston at one end of the container, which is magically transparent to oxygen atoms, and slowly and isothermally move it to the center of the container. This

puts the nitrogen back where it started. The change in entropy of the nitrogen is

$$\Delta S_N = \frac{Q}{T} = -\frac{W}{T} = nR \log \frac{V_i}{V_f} = -nR \log 2 = -Nk_B \log 2$$

where  $N$  is the number of nitrogen molecules, which is also the number of oxygen molecules. If we do the same for the oxygen with another piston, introduced from the other side, we get another, equal decrease in the entropy. This takes us back to the original state, so since entropy is a state function, the increase in entropy when we allowed the gases to mix was

$$\Delta S = -2\Delta S_N = 2Nk_B \log 2.$$

**[3] Problem 3.** Some basic conceptual questions about the second law.

- (a) Verify explicitly that the second law of thermodynamics allows heat to flow from a hot body to a cold one, but not vice versa.
- (b) Derive the efficiency of a Carnot engine operating between reservoirs of temperatures  $T_H > T_C$  by using the fact that the Carnot cycle is reversible, i.e. that after a complete Carnot cycle the entropy of the universe is unchanged.
- (c) Let's suppose that in addition to reservoirs at temperatures  $T_H$  and  $T_C$ , you had a reservoir at temperature  $(T_H + T_C)/2$ . Is the Carnot cycle between the extreme reservoirs still the most efficient option? If not, what's the maximum possible efficiency?
- (d) Consider the free expansion of an ideal gas. No work is done, and the internal energy  $U$  does not change. However, the entropy increases since the process is irreversible. On the other hand, the first law states

$$dU = TdS + \bar{d}W$$

so there appears to be a contradiction. What is wrong with this argument?

**[3] Problem 4.** Most thermodynamic systems have positive heat capacities, but a gravitationally bound system of particles has a negative heat capacity by the virial theorem, as does a black hole. Suppose a system with negative heat capacity  $C_1$  is placed in thermal contact with a slightly cooler system with positive heat capacity  $C_2$ .

- (a) What happens in the long run if  $|C_1| > |C_2|$ ?
- (b) What happens in the long run if  $|C_1| < |C_2|$ ?
- (c) Some systems, such as lasers, have a negative temperature and a positive heat capacity. (You'll see another example of negative temperature in problem 11.) What happens if such a system is placed in thermal contact with a system with positive temperature and heat capacity?

**[3] Problem 5** (Kardar). The solar system originated from a dilute gas of particles, sufficiently separated from other such clouds to be regarded as an isolated system. Under the action of gravity the particles coalesced to form the sun and planets.

- (a) The motion and organization of planets is much more ordered than the original dust cloud. Why does this not violate the second law?

- (b) The nuclear processes of the sun convert protons to heavier elements such as carbon. Does this further organization lead to a reduction in entropy?
- (c) The evolution of life and intelligence requires even further levels of organization. Why does this not violate the second law?

For a nice discussion of these issues, see section 36.7 of Blundell and Blundell.

- [4] **Problem 6.** We can find an expression for entropy of an ideal gas using pure thermodynamics. Consider  $n$  moles of an ideal gas with adiabatic index  $\gamma$ , initially at pressure  $P$  and volume  $V$ .

- (a) Suppose the gas is heated at constant volume to pressure  $P'$ . Compute the change in entropy.
- (b) Suppose the gas is instead expanded at constant pressure to volume  $V'$ . Compute the change in entropy.
- (c) By combining your two results, show that the entropy  $S(P, V)$  of the ideal gas is

$$S = nC_V \log(PV^\gamma) + \text{constant}$$

when  $n$  is constant. The constant can't be found using classical thermodynamics alone, since it only can determine changes in entropy.

- (d) Now consider the case where the ideal gas is allowed to freely expand into a vacuum, reaching a final volume  $V'$ . This is an example of a nonequilibrium process that changes the entropy, despite no work or heat transfer. Find the resulting change in entropy.
- (e) So far, we don't know anything about how  $S$  depends on  $n$ , because we only considered processes that kept  $n$  constant. However, the entropy should be extensive, meaning that if the pressure and temperature are kept the same, then  $S$  is proportional to  $n$ . Thus, if in general

$$S/n = C_V \log(PV^\gamma f(n)) + \text{constant}$$

then the right-hand side must be independent of  $n$ . Using this, find  $f(n)$ .

- [3] **Problem 7.** ⌚ USAPhO 2014, problem A2.
- [3] **Problem 8.** ⌚ USAPhO 2017, problem A2. A simple but tricky problem. (For a more involved problem based on the same fundamental idea, see [Physics Cup 2019, problem 2.](#))
- [3] **Problem 9.** ⌚ USAPhO 2020, problem B3. An advanced problem which shows you how to compute corrections for a slightly nonideal gas. (Note that you can use the quantity  $T$  in any of the parts, not just (e) and (f).)

### 3 Quantum Statistical Mechanics and the Third Law

#### Idea 6: Third Law

The entropy of any thermodynamic system approaches zero as the temperature goes to zero.

The third law of thermodynamics is a bit of a black sheep. Unlike the other laws, it is inherently quantum mechanical, and comes from the statistical definition of entropy, as described below.

- [2] **Problem 10.** Show that the third law requires the specific heat  $C$  to approach zero as the temperature approaches zero. (Since this isn't true for a classical ideal gas, the ideal gas law *must* break down in quantum mechanics, at low temperatures. You can also see this from your result for the entropy in problem 6, which diverges at low temperatures; there's no way to add a constant to it to make it go to zero at zero temperature.)

### Idea 7: Microcanonical Ensemble

If a large isolated system in thermodynamic equilibrium is in one of  $\Omega$  quantum states, then we assume all of the states are equally likely, and the entropy is

$$S = k_B \log \Omega.$$

This is the fundamental definition of entropy; in fact the equation  $dS = dQ/T$  does not define entropy, but rather defines temperature! The third law follows from this result because at temperature  $T = 0$ , all systems settle into their ground state, so  $\Omega = 1$  and  $S = 0$ .

### Remark

The assumption that all states are equally likely may sound a bit confusing, because in **T1** we used the Boltzmann distribution, which stated that higher-energy states were less likely to be occupied. The difference is that for the Boltzmann distribution, we assumed the system was inside an environment at temperature  $T$ . In fact, all quantum states of the system and environment together remain equally likely in this context, but lower-energy states of the system correspond to more energy in the environment, and hence (usually) more possible environment states. Thus, accounting for the number of possible environment states, lower-energy states of the system are more likely. You will make all this precise in problem 11, where you will use it to derive the Boltzmann distribution.

### Example 4: Mixing Entropy

Repeat example 3 using statistical mechanics.

### Solution

In statistical mechanics, we can compute  $S$  directly using  $S = k_B \log \Omega$ . Focusing on a single oxygen or nitrogen molecule, let  $\Omega_i$  be the initial number of possible states. When the partition is removed, the possible position space for the particle doubles, and since quantum states all have the same volume in phase space, the number of states available doubles. So for one particle,

$$\Delta S_{\text{part}} = k_B \log(2\Omega_i) - k_B \log \Omega_i = k_B \log 2.$$

Since the  $2N$  particles are independent,

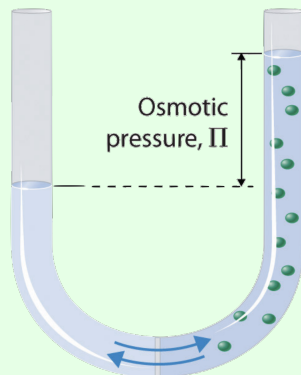
$$\Delta S = 2N \Delta S_{\text{part}} = 2N k_B \log 2$$

in agreement with the result from thermodynamics.



### Example 5: Osmotic Pressure

A water-filled tube is divided by a membrane which only allows water molecules through. When a substance is dissolved in one half of the tube, forming  $N$  ions, the water level on that side is observed to go up, as shown.



Explain why, and calculate the height difference, neglecting interactions between different ions, and between the ions and the water.

### Solution

Osmotic pressure is an example of an entropic force. Suppose that the water level on one of the tube goes up. The water is otherwise unchanged, since it's just translated along the tube, so its entropy has not changed. But its energy is higher, and this energy must have come via heat transfer from the environment. Hence the entropy of the environment has decreased.

If there were no dissolved ions, this would show why water levels don't spontaneously go up: it is forbidden by the second law. But here, increasing the water in the section with the ions increases the ions' entropy, since they now have more positions they could be in. Since the number of available quantum states is proportional to the available volume, we have

$$S_{\text{ion}} = Nk_B \log V$$

up to an additive constant, which means that

$$dS_{\text{ion}} = \frac{Nk_B}{V} dV.$$

In equilibrium, let there be an osmotic pressure difference  $P$  across the membrane. Then moving a volume  $dV$  of water across it costs energy  $P dV$ , so

$$dS_{\text{env}} = -\frac{dQ}{T} = -\frac{P}{T} dV.$$

In thermodynamic equilibrium the total entropy is maximized, giving

$$dS_{\text{ion}} + dS_{\text{env}} = 0.$$

Solving for the osmotic pressure gives

$$PV = Nk_B T.$$

The height difference is simply found using hydrostatic pressure,

$$h = \frac{P}{\rho g} = \frac{Nk_B T}{\rho V g}.$$

The appearance of a pressure here in response to an opportunity for decreasing entropy is generally called an “entropic force”. In fact, you might have already seen this result in chemistry class, but in rather different notation. In chemistry textbooks, you’ll typically see

$$\Pi = iMRT$$

where  $\Pi$  is the osmotic pressure,  $M$  is the molarity of solute, and the van ’t Hoff factor  $i$  is the number of moles of ions per mole of solute. This is perfectly equivalent to what we found.

### Remark

Why does the expression for osmotic pressure bear a suspicious resemblance to the ideal gas law? The reason is that, by completely neglecting interactions between the solute and solvent, we have effectively treated the solute ions like an ideal gas, from the standpoint of entropy. This gives an additional contribution to the pressure, which can be derived just like the pressure of an ideal gas is in **T1**. (Using this reasoning backwards, one can conclude that the pressure of an ideal gas can also be described as an entropic force, using the same reasoning as above.)

The reason that this kinetic theory connection to the ideal gas law is not emphasized in chemistry classes is that the solute particles don’t actually behave like an ideal gas at all. They are part of a liquid, and hence are constantly bumping into the solvent particles. The thermodynamic reasoning we used above is more general, because it only requires that the interaction between the solvent and solute doesn’t significantly change the solute’s entropy.

By the way, if you’re wondering precisely when an entropic force *is*, rest assured that nothing magical is going on here. In the end, there are no forces at play except for the ordinary forces you’re familiar with. It’s just that often, calculating what those forces do in aggregate in thermodynamically large systems is very hard. An entropic force simply means a force whose value we can calculate most easily using the tool of entropy.

- [5] **Problem 11.** This is a long problem, but a really useful one that ties a lot of things together. Consider a set of  $N$  atoms, each of which may be in the ground state, with zero energy, or in an excited state with energy  $E_0$ . Suppose it is only known that the total energy of the system is  $\alpha N E_0$ .

(a) For large  $N$ , Stirling’s approximation states that

$$\log N! \approx N \log N - N.$$

Derive this expression by writing  $\log N!$  as a summation. Is the exact answer larger or smaller?

(b) Using Stirling’s approximation, show that the entropy of the system is

$$S = Nk_B \left( \alpha \log \frac{1}{\alpha} + (1 - \alpha) \log \frac{1}{1 - \alpha} \right).$$

Sketch the entropy as a function of  $\alpha$ .

- (c) Using the definition of temperature,  $dS = dQ/T$ , show that the system has a temperature of

$$T = \frac{E_0}{k_B} \frac{1}{\log(1 - \alpha) - \log \alpha}.$$

Sketch the temperature as a function of  $\alpha$ . In particular, what temperature do you need to get  $\alpha = 1/2$ ? How about  $\alpha = 1$ ?

- (d) Show that the third law is satisfied.
- (e) Now consider just a single one of the  $N$  atoms, where the total energy of the system is  $\alpha N E_0$  as before. Show that the probability it is excited obeys the Boltzmann distribution.

The following question illustrates some important features of free energy.

- [5] **Problem 12.** A rubber molecule can be modeled in one dimension as a chain consisting of a series of  $N = N_+ + N_- \gg 1$  links of length  $a$ , where  $N_+$  point in the  $+z$  direction and  $N_-$  point in the  $-z$  direction. The orientations of each of the links are independent, and it costs no energy to rotate a link. Thus all configurations are equally likely at any temperature.

- (a) Find the length  $L$  of the chain in terms of  $N_+$  and  $N_-$ .
- (b) Using Stirling's approximation,

$$\log N! \approx N \log N - N$$

and assuming that  $L \ll Na$ , show that

$$S = Nk_B \left( \log 2 - \frac{L^2}{2N^2 a^2} \right)$$

and hence show that the entropy decreases as  $L$  increases. This model is very basic; for instance, it tells us nothing about the tension in the chain, but the sign here is correct. (Hint: notice that the changes in entropy are of order  $L^2/N$ , and hence are very small. You will have to be careful with your approximations to avoid dropping an important contribution.)

- (c) Note that the entropy doesn't depend on temperature, because all configurations have the same energy. In a more realistic model, the total entropy of the molecule would be the sum of the entropy computed above, plus the entropy due to additional thermal motions, such as vibrations of the links. Given this, if a rubber band is quickly stretched, does it feel warm or cold to the touch? (Hint: do the experiment!)
- (d) Now suppose a mass  $m$  is hung from the bottom of the chain. Find the average length of the chain at temperature  $T$ , using the Boltzmann distribution. Do not assume  $L \ll Na$ . (The force that raises this mass is another example of an entropic force.)
- (e) It turns out that you can also arrive at this answer by maximizing the entropy of the entire universe  $S_{\text{univ}}$ . Argue in general that  $S_{\text{univ}}$  is maximized when the Helmholtz free energy

$$F = U - TS$$

is minimized, where  $U$  is the total energy of the rubber band and mass, and  $T$  is the temperature of the system and its surroundings. (Hint: use the expression for  $dS_{\text{univ}}$ , along with conservation of energy.)

- (f) Show that in the limit  $L \ll Na$ ,  $F$  is indeed minimized at the equilibrium length.

### Remark: Is Entropy Subjective?

Our definition of entropy talked about the number of states the system could be in, but of course in reality the system is actually in just one state. We really mean the number of states the system could be in given the information you know. For example, for an ideal gas we usually know the temperature, pressure, and volume.

You might find this disturbing because it implies the entropy is subjective, depending on what each person personally knows. But the entropy determines things like  $F$ , and hence the amount of work that can be extracted from the system, so how could it differ?

In fact, there is no contradiction between these two things. For example, suppose that besides the usual properties, particles in an ideal gas can be slightly bigger or smaller. If you happen to somehow know that all of the big particles are on the left side of a box and all the small particles are on the right side, then the entropy you assign to the box is lower than somebody who doesn't know about this, and accordingly you can extract more work from the box. For instance, you can put a filter that only allows small particles through in the middle, then collect energy as small particles move from right to left.

On the other hand, if you are hit on the head and forget about the particle size, your entropy for the gas goes up. And accordingly, you can now extract less work, because you wouldn't know it made sense to put in such a filter in the first place!

While this example may seem contrived, it's actually exactly the same as our earlier mixing entropy example! If you don't have any equipment that can distinguish between oxygen and nitrogen molecules, then the "unmixed" state already has just as much entropy as the mixed state. In this case, the entropy doesn't change at all when you allow the gases to mix.

What this example shows is that you don't have to know everything about a system to use thermodynamics. Even if you are wrong about how the system behaves microscopically, you can still use the theory to describe what *you* can do with the system, using *your* knowledge. For more about this, see the insightful articles by E.T. Jaynes, [\*The Gibbs Paradox\*](#) and [\*Information Theory and Statistical Mechanics\*](#).

## 4 Blackbody Radiation

### Idea 8: Planck's Law

The radiation power per unit area from an object at temperature  $T$  at frequency  $f$ , per unit frequency, is

$$I(f) = \frac{2\pi h f^3}{c^2} \frac{1}{e^{hf/k_B T} - 1} e(f)$$

where  $e(f)$  is called the emissivity. For an ideal blackbody,  $e(f) = 1$ . By integrating over all

frequencies, one finds that the total power per unit area from an ideal blackbody is

$$J = \int_0^\infty I(f) df = \sigma T^4, \quad \sigma = \frac{2\pi^5 k_B^4}{15c^2 h^3}.$$

This is the Stefan–Boltzmann law. We derived it in the preliminary problem set with dimensional analysis.

[2] **Problem 13.** Some basic manipulations with Planck’s law.

- (a) Show that  $J \propto T^4$  by nondimensionalizing the integral, as explained in **P1**.
- (b) Prove Wien’s displacement law, which states that the maximum of  $I(f)$  for an ideal blackbody occurs at a frequency which is proportional to  $T$ .

We will derive Planck’s law in **X1**.

The Stefan–Boltzmann law comes up fairly often in an Olympiad context.

### Example 6

At high temperatures, the resistivities of many temperatures scale approximately linearly with temperature. Suppose that a light bulb supplies total radiation power  $P$  when connected across an AC voltage  $V$ . How does  $P$  scale with  $V$ ?

### Solution

The total power dissipated in the resistor scales as  $V^2/R$ , and  $R \propto \rho \propto T$ . This power must equal the blackbody radiation power emitted, so by the Stefan–Boltzmann law,

$$P \propto T^4 \propto \frac{V^2}{T}$$

from which we conclude that

$$T \propto V^{2/5}, \quad P \propto V^{8/5}.$$

However, note that this doesn’t correspond directly to what you would see. At low temperatures, most of the energy is below the visible range, so you wouldn’t see anything. At very high temperatures, a significant fraction would go into UV as well.

### Example 7

A nonideal blackbody absorbs a fraction  $a(f)$  of incident radiation of frequency  $f$ . (For an ideal blackbody,  $a(f) = 1$ .) Show that the second law implies Kirchoff’s law,  $e(f) = a(f)$ .

**Solution**

First let's consider the case where  $a(f)$  and  $e(f)$  are constants, which don't depend on  $f$ . If we place two blackbodies of temperature  $T$  right next to each other, then the rate at which the first heats the second is proportional to  $e_1 a_2$ , while the rate the second heats the first is proportional to  $e_2 a_1$ . By the second law, these must be equal, which implies  $e_1/a_1 = e_2/a_2$ . Hence the ratio is constant. But for a perfect blackbody,  $e = a$ , so the two are equal.

More generally, suppose we place in between the two blackbodies a filter that only lets light of frequency  $f$  through, and blocks everything else. Then the same reasoning gives  $e_1(f)/a_1(f) = e_2(f)/a_2(f)$ , so  $e(f) = a(f)$ . Thermodynamics forces good absorbers to also be good emitters.

You might complain that this argument assumes filters that let light through in one direction have to also let it through in the other direction. In fact, this is also required by the second law of thermodynamics. If it weren't true, then the filters would act like Maxwell's demons for photons. If you used one to divide a room full of photons, it would start directing photons of certain frequencies to one side or the other, decreasing the entropy.

**Example 8**

A satellite contains a nuclear reactor at its core, fixed at temperature  $T_H$ . Suppose the exterior of the satellite is somehow fixed at temperature  $T_C$ , and the satellite runs an ideal heat engine, using the core and exterior as hot and cold reservoirs. Treat the exterior as an ideal blackbody, and neglect the temperature of outer space. What value of  $T_C$  maximizes the rate of work extracted?

**Solution**

One might think that  $T_C$  should be as low as possible, to increase the efficiency of the cycle. However, a lower  $T_C$  means a lower rate of heat emission  $\dot{Q}_{\text{out}}$ , since we are limited by emission of blackbody radiation, which lowers the rate of work done  $\dot{W}$ .

The rate of heat emitted from the satellite is

$$\dot{Q}_{\text{out}} = \sigma A T_C^4 \propto T_C^4.$$

Therefore, the rate of work done is

$$\dot{W} = \eta \dot{Q}_{\text{in}} = \eta \frac{T_H}{T_C} \dot{Q}_{\text{out}} \propto \left(1 - \frac{T_C}{T_H}\right) \frac{T_H}{T_C} T_C^4 = \left(\frac{T_C}{T_H}\right)^3 - \left(\frac{T_C}{T_H}\right)^4.$$

This is maximized when  $T_C = (3/4)T_H$ , at which point the efficiency is  $1/4$ .

- [2] **Problem 14.** Consider two adjacent blackbodies with similar temperatures, exchanging energy by radiation. Show that the rate of change of the temperature difference is approximately proportional to the temperature difference itself. This is an example of Newton's law of cooling, which implies

the temperature difference decays exponentially. How does the time scale for cooling depend on the average temperature of the blackbodies?

- [3] **Problem 15** (IPhO 1992). A satellite is a sphere of diameter  $D$  orbiting about the Earth.
- First suppose the satellite is perfectly black. Ignoring the effect of the Earth, find its temperature  $T$  in terms of the temperature  $T_{\odot}$  of the Sun, the radius  $R_{\odot}$  of the Sun, and the radius  $R$  of Earth's orbit.
  - How does the answer change if the satellite has a uniform emissivity  $e(f) = e_0 < 1$ ?
  - If an engineer wishes to make  $T$  as small as possible using a special paint, which can have an arbitrary  $e(f)$ , sketch how the function  $e(f)$  should look.
- [3] **Problem 16** (IPhO 1996). Two perfectly black surfaces of temperatures  $T_h$  and  $T_l$  are parallel to each other in vacuum, and the net heat flux from the hotter surface to the colder one is  $P$ .



Now suppose that  $N$  parallel, thermally insulating, perfectly black plates are placed in between them. This shielding reduces the heat flux to  $P'$ . Find  $P'$  in terms of  $P$ .

- [3] **Problem 17.** ⌚ USAPhO 2011, problem A3.
- [3] **Problem 18.** ⌚ USAPhO 2006, problem B2.
- [3] **Problem 19.** ⌚ USAPhO 2019, problem A2. An elegant, tricky radiation problem.

As shown in this example and the following problem, thermodynamic reasoning can be very powerful!

### Example 9

Suppose a black sphere of radius  $r$  is placed inside a large container of arbitrary shape, whose walls are a blackbody with temperature  $T$ . In the steady state, what is the rate of power absorbed by the sphere from the walls?

### Solution

Radiation is continually emitted from all points on the walls in all directions, so it would seem that we have to do a complicated integral to find how much hits the sphere, which would be hard even if we were given the container's shape. But the problem is actually trivial due to the second law of thermodynamics!

Let's suppose the sphere itself is also at temperature  $T$ . Then it sends power  $(4\pi r^2)\sigma T^4$  to the walls, since all radiation that exits the sphere hits the walls. But if the sphere and walls are at the same temperature, a temperature difference cannot spontaneously appear. Therefore, the power going from the walls to the sphere must also be  $(4\pi r^2)\sigma T^4$ . And

since this quantity can also be expressed as a geometric integral, it can't depend on the temperature of the sphere, so it must be true for any sphere temperature.

For a tough problem which uses a variant of this idea, see [EFPPhO 2010, problem 2](#).

[3] **Problem 20.** Planck's law can be used to understand atomic physics, as Einstein showed and as you will now show. Suppose we have a collection of two-state atoms encased inside a cavity, whose walls are blackbodies with temperature  $T$ . In equilibrium, there will be  $N_1$  atoms in the ground state and  $N_2$  atoms in the excited state, bathed in a photon gas of temperature  $T$ . Three physical processes can occur:

1. Atoms in the excited state can each spontaneously decay with rate  $A$ .
2. Atoms in the ground state can absorb a photon. For each atom, the rate of this process is  $B\rho(f)$  where  $\rho(f)$  is the energy density of radiation at frequency  $f$ .
3. Atoms in the excited state can undergo stimulated emission when interacting with a photon, in the process  $e + \gamma \rightarrow g + \gamma + \gamma$ . For each atom, the rate of this process is  $B'\rho(f)$ .

These three "Einstein coefficients" depend on the frequency, but not on the temperature.

- (a) Argue that  $\rho(f)$  is proportional to  $I(f)$  in Planck's law. In fact,

$$I(f) = \frac{c}{4} \rho(f)$$

though you do not have to show this.

- (b) Write down an expression for  $dN_2/dt$ .
- (c) In the steady state, for any temperature,  $N_2/N_1$  must be given by the Boltzmann distribution. Using this, find  $A$  and  $B'$  in terms of  $B$ .

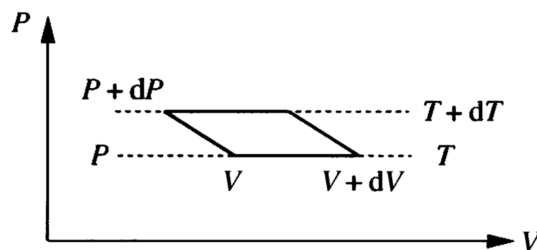
This is an extraordinary result: one might have expected a difficult calculation to find each of  $A$ ,  $B$ , and  $B'$ , but if you know one, then by thermodynamics you know them all. In particular, thermodynamics *requires* spontaneous emission to be possible,  $A \neq 0$ .

[3] **Problem 21.** In T1, we derived some properties of photon gases using basic kinetic theory. Here, we'll derive some more properties, starting from Planck's law and then sticking with pure thermodynamics. As in problem 20, we assume there is a photon gas at temperature  $T$ , with pressure  $P$  within a cavity of volume  $V$ , whose walls are perfect blackbodies. (Note that since photons can be absorbed and emitted by the walls, it doesn't make sense to talk about  $N$  as for an ideal gas. Instead,  $N$  is determined by the other parameters. This actually makes things simpler, since there's one less variable to worry about.)

- (a) It turns out that the pressure of the photon gas is  $P = AT^4$  where  $A$  is a constant. Explain why the pressure depends only on the temperature. (Harder, optional task: explain why  $P \propto T^4$  starting from Planck's law.)



- (b) Our next goal is to compute  $U(T, V)$ . Consider an infinitesimal Carnot cycle, shown below.



By equating the efficiency of this cycle to the Carnot efficiency, find  $(\partial U / \partial V)|_T$ .

- (c) By integrating this result, and using  $U(T, 0) = 0$ , find  $U(T, V)$ .
- (d) We can now use these results to find  $S(T, V)$ , just as we did for an ideal gas in problem 6, i.e. by considering the change of entropy during some infinitesimal process and then integrating the result. Do this in any way you like. Can the third law be satisfied?

[5] **Problem 22.** [Physics Cup 2018, problem 9.](#)

## 5 Heat Conduction

Now we treat heat conduction, which you investigated using kinetic theory in **T1**.

### Idea 9: Fourier's Law

The flux of heat (i.e. the power per unit area) due to thermal conduction is related to the temperature gradient by

$$J = -\kappa \frac{dT}{dx}.$$

By considering the net heat flowing in and out of a slab of width  $dx$ , we have

$$\frac{du}{dt} = -\frac{dJ}{dx} = \kappa \frac{d^2T}{dx^2}.$$

where  $u$  is the energy density. Intuitively, this shows how heat conduction works to smooth out temperature gradients. For example, if the temperature had a local minimum, then  $du/dt$  would be positive at that point, as heat flows in from all directions.

### Example 10

Find the steady state heat flow through a rectangular slab of area  $A$  and thickness  $d$ , whose left and right ends are held at temperature  $T$  and  $T + \Delta T$ .

### Solution

By the argument in the above idea, in the steady state the energy density within the slab is constant, so  $d^2T/dx^2 = 0$  and the temperature gradient is uniform. In this steady state, heat simply flows uniformly through the slab, without accumulating anywhere in it, just like current flows uniformly through a resistor. (Since this “internal” steady state is often

reached quickly, in many problems it is just assumed to hold, without being explicitly stated.)

Therefore, the temperature gradient is just  $\Delta T/d$  everywhere in the slab, so

$$P = JA = \frac{\kappa A \Delta T}{d}.$$

This is single most important equation for heat conduction problems.

### Remark: Continuity Equations

We can generalize the results above to three dimensions. In this case, the flux of heat is described by a vector,

$$\mathbf{J} = -\kappa \nabla T.$$

The conservation of energy equation becomes

$$\frac{du}{dt} = -\nabla \cdot \mathbf{J} = \kappa \nabla^2 T.$$

The first equality here is known as a continuity equation, and it is the general way that the density of a conserved quantity is related to a flow. For example, the continuity equation for fluids, which we saw in one-dimensional form in **M7**, is written in three dimensions as

$$\frac{d\rho}{dt} = -\nabla \cdot (\rho \mathbf{v})$$

and expresses conservation of mass. The continuity equation for charge, which we'll see in **E4**, is very similar,



$$\frac{d\rho}{dt} = -\nabla \cdot \mathbf{J}$$

where  $\rho$  is a charge density and  $\mathbf{J}$  is a current density.

- [2] **Problem 23** (Grad). Estimate how long it will take for a small pond of depth  $D = 0.5$  m to freeze completely in a very cold winter. Take the thermal conductivity of the ice to be  $\kappa = 2.2$  W/mK, the latent heat of fusion  $L = 3.4 \times 10^5$  J/kg, and the density  $\rho = 0.9 \times 10^3$  kg/m<sup>3</sup>. Take the outside temperature to be a constant  $T_0 \approx -10^\circ\text{C}$ , and for simplicity, suppose the entire pond starts at a uniform temperature  $0^\circ\text{C}$ .

- [3] **Problem 24**. Two problems on heat conduction through spheres.

- Consider a brick pizza oven, which we model as a sphere with inner radius  $R_1$ , outer radius  $R_2$ , and thermal conductivity  $\kappa$ . The inside of the oven is held at temperature  $T_1$  and the outside of the oven is held at temperature  $T_2$ . In equilibrium, find  $T(r)$ , and the power needed to keep the inside of the oven at temperature  $T_1$ .
- Consider a sphere of a radioactive rock, which constantly produces heat  $\sigma$  per unit volume. The outside of the sphere is held at temperature  $T_0$ , the sphere's radius is  $R$ , and its thermal conductivity is  $\kappa$ . Find the temperature at the center of the sphere.

- [2] **Problem 25.** Supposing that it takes five minutes to boil an egg. Estimate how long it takes to boil an ostrich egg, which has three times the diameter.
- [3] **Problem 26.**  USAPhO 2016, problem A4. A practical conduction question.
- [3] **Problem 27.**  USAPhO 2013, problem A1. An elegant conduction question.