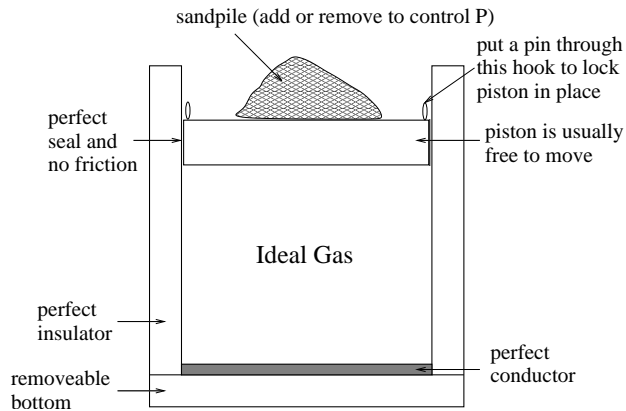


Hopefully, after lecture 23, you now have a better understanding of the meanings of INTERNAL ENERGY and of TEMPERATURE. Now we will begin to study in detail the relationship between heat, work, and the internal energy of a system; our goal is to discover how an inflow of heat can be used to do work, and how work can be used to transfer thermal energy from place to place (so as to heat or cool our houses) -- this is the study of THERMODYNAMICS.

We already know that there are two ways to change the internal energy U of a system of particles. First, we learned in Chapter 6 that you can change the U of a system by doing work on the system. For example, a block sliding across a rough floor gets hotter because of the work done by kinetic friction; the system of particles is the atoms of the block, work is done ON the system by the force of kinetic friction on the block by the floor (external to the system), and the thermal energy (and thus U) of the system increases. Second, we learned in Chapter 12 that you can increase the U of a system by putting it in contact with a second system which is at a higher temperature; heat will flow from the hotter system to the cooler system until the temperatures are the same, and the U of the first system will have been increased.

The system that we will study most often here in Chapter 15 consists of an ideal gas enclosed in a "thermodynamic gas pot", as drawn below. The piston is free to move up and down without friction, and has a perfect seal, so that none of the gas can escape. The piston and the walls of the pot are perfect insulators; no heat at all can flow through. At the bottom of the pot is a perfect heat conductor; there is a perfect insulating piece under the conductor that can be removed if we want to put the gas in contact with a second system at a different temperature. Pressure is controlled by adding or removing sand from the top of the piston. There are little hooks on top of the piston; by sticking pins through those hooks, we can lock the piston in place.



We can make heat flow out of our system by removing the insulating piece at the bottom and putting the pot in contact with a cold reservoir, such as a big block of ice, or dry ice, or even liquid nitrogen (it must be colder than our gas). If we want an inflow of heat, then we use a hot reservoir. What will happen when we use the cold reservoir? The gas will cool, and since the pressure is constant (we haven't done anything to the sandpile) the volume of the gas will decrease, i.e. the piston will move down. When the piston moves down, it does work ON the PARTICLES of our system (please read the footnote).¹ Let's calculate the amount of work done ON the particles of our system (W_{on}) if the piston comes down by an amount $|\Delta y|$. The pressure in our gas is P .

DEF of work by a constant force

$$W_F \equiv (F_{\Delta r})\Delta r$$

\mathbf{F}_{GP} is the force on the gas by the piston

$$= F_{GP}|\Delta y|$$

A is the area of the piston

$$= PA|\Delta y|$$

ΔV is the change in volume of the gas

$$= P|\Delta V|$$

¹ Work is done ON THE PARTICLES, but the net work done ON THE SYSTEM, i.e. on the center of mass (CM) of the system, is ZERO. The CM moves down a little, due to the compression, so the force on the gas by the piston, \mathbf{F}_{GP} , does positive work on the CM. But the force on the gas by the "floor" of the pot does an equal amount of negative work on the CM, because the CM moves down while that force is pointing up. So the net work done on the SYSTEM is really zero, and the KE of the CM doesn't change.

In this case, W_{on} is positive because F_{GP} is in the same direction as the displacement Δy . So in this case we need the absolute value sign on ΔV because the volume has DECREASED and so ΔV would be negative. But what is the work done BY our gas particles ON the piston (W_{by})? Since, by the 3rd Law, $\mathbf{F}_{PG} = -\mathbf{F}_{GP}$, and since both forces move through the same displacement Δy , therefore W_{by} would be the opposite of the W_{on} ; in fact, this must always be true, $W_{by} = -W_{on}$. So, returning to our example, in this case W_{by} is negative, so we could write $W_{by} = P\Delta V$ without any absolute value signs. This equation will appear on your equation sheet; it can be used to calculate the W_{by} whenever there is a change in volume of the system, as long as the pressure is constant (remember that we used the DEF of work done by a CONSTANT FORCE in calculating W_{on}). If the pressure is not constant, then we can use the same trick that we used to calculate the work done in stretching a spring. In that case, we made a graph of the changing force versus displacement and found the area under that curve (see page 3 of reading 17); for changing pressure we would make a graph of the changing pressure versus volume and again find the area under the curve. See Figures 15.6, 15.8, 15.9, and 15.10 in your text for four examples of calculating W_{by} when there is a change in volume; Figure 15.6 shows the case in which the pressure is constant. We already know that it is possible to do work on a system without any change in volume, for example we could use Joule's system of falling weights and turning vanes that you learned about in lecture 22; but with our "thermodynamic gas pot" design as drawn above, the only way to have nonzero W_{on} or W_{by} is to move the piston, and $W_{by} > 0$ whenever the piston moves up (i.e. whenever $\Delta V > 0$).

Now we are ready for the 1st Law of Thermodynamics, which is a statement of conservation of energy for a system of particles with the following properties: (1) the net force on the system is zero (\mathbf{v}_{CM} is constant);

(2) the net torque on the system is zero about any axis; (3) PE_G for the system is not changing (gravity isn't doing any work on the system -- this is easy for our "massless" gases); and (4) the system is not interacting with any external springs.² The 1st Law is simply a formal statement of what we already know, that the internal energy of a system of particles can be changed in two ways: (1) by an inflow or outflow of heat, and (2) by doing work on the system of particles. Here is one way to write the 1st Law:

$$\Delta U = Q + W_{on}$$

All three of these symbols are signed scalars. The sign of Q is given by the DEF of specific heat ($c \equiv (Q/m\Delta T)$); in that DEF, Q is positive when ΔT is positive, so positive Q means an inflow of heat. As we as seen above, W_{on} is less convenient to work with than is W_{by} , particularly for our thermodynamic pot, so the conventional way of writing the 1st Law is

$$\Delta U = Q - W_{by}$$

When doing problems with the gas pot, W_{by} is always positive when ΔV is positive, and W_{by} is always negative when ΔV is negative. For example, compression means ΔV is negative, so W_{by} is negative, so the effect of the work term is to increase U . We will do at least two examples with the 1st Law during lecture, one of which will deal with the specific heat of ideal gases.

You may have noticed that the table of specific heats on page 350 in Chapter 12 only gave the specific heats of liquids and solids; because gases are compressible, the conditions under which heat inflow occurs determine how much heat is required to increase the temperature by one degree. The 1st Law

² The 1st Law is actually more general than this, but it is easiest to grasp if we put on these restrictions.

and our thermodynamic gas pot allow us to deal with this difficulty. I will do this in lecture, but we need one new DEF which I give here:

DEF The MOLAR SPECIFIC HEAT CAPACITY (C) for a given material of n moles is

$$C \equiv \frac{Q}{n\Delta T} \quad \text{units are } \frac{\text{J}}{\text{mol} \cdot \text{K}} \text{ or } \frac{\text{J}}{\text{mol} \cdot (^\circ\text{C})}$$

where Q is the necessary heat inflow (outflow) to create an increase (decrease) in temperature ΔT . C is a positive-only scalar.

Also, please learn the following DEFs relating to the various kinds of processes that we can carry out with our thermodynamic gas pot:

DEFs An ISOBARIC PROCESS is a process that takes place at constant pressure. An ISOCHORIC PROCESS is a process that takes place at constant volume.

We have already done an example of an isobaric process; as long as the amount of sand is not changed and the piston is free to move, the process will be isobaric. For an isochoric process, the piston must not move; this is usually done by locking the piston in place -- now the amount of sand is no longer controlling the pressure in the gas. W_{by} is always zero in a isochoric process with our thermodynamic gas pot.

DEFs An ISOTHERMAL PROCESS is one that takes place at constant temperature.

To get an isothermal process, we remove the insulating piece at the bottom and put the now-exposed conducting piece in contact with a heat reservoir that is at the same temperature as our gas. Now we can add sand or remove sand; the piston will move down or up, respectively, but the temperature of our gas cannot change. Of course, what happens when we add a little sand is that the piston moves down, doing work on the gas and raising its temperature a little;

but then the gas is hotter than the reservoir so that heat outflow occurs until the gas is back at its original temperature. The idea is to add or remove the sand a grain at a time (i.e. very slowly) so that the little temporary changes in the temperature are not measureable.

DEFs An ADIABATIC PROCESS is one that takes place without any heat inflow or outflow.

To get an adiabatic process, we simply keep the insulating piece at the bottom in place. Now we can do work on the gas by adding or removing sand, but no heat flow can occur because the gas is perfectly insulated. The word "adiabatic" means "does not go through"; the "dia" has the same meaning as in "diameter" which means to measure across or through. Unlike the isothermal process, it is possible to have an adiabatic process without moving one grain of sand at a time; we can just drop a big rock on top of the piston -- the process is adiabatic as long as the insulating piece at the bottom is in place. However, we will be more interested in the very slow adiabatic process, in which we add or remove a grain at a time; these very slow processes are REVERSIBLE:

DEF A REVERSIBLE PROCESS is one which, if run backward in time, would be physically reasonable.

In other words, make a movie or video of the process and run it backwards; then ask yourself, "Is what I am watching physically possible?" If the answer is yes, then the process is reversible. For example, recall the isobaric process I described at the beginning of this reading. We removed the insulating piece at the bottom of the pot, and put it in contact with a cold reservoir; as a result the gas got colder and the piston came down. A movie of this process run backwards would show the gas getting hotter while in contact with a cold

reservoir, which is impossible; so the isobaric process is IRREVERSIBLE. On the other hand, to create a slow adiabatic process we could add one grain of sand at a time to make the piston move down; the gas would slowly be compressed and would warm up. A movie of this process run backwards would show one grain of sand being removed at a time; the piston would slowly be moving up and the gas would be cooling -- a perfectly reasonable physical process. Therefore, the slow adiabatic process is reversible. We will see in lecture 25 that reversible processes have a very special place in thermodynamics.

In lecture 24, we will look at graphs of pressure versus temperature for each of the above processes. We will finish lecture 24 by defining HEAT ENGINE and REFRIGERATOR (or HEAT PUMP). We will also define the EFFICIENCY (e) of a heat engine (it is exactly what you would expect) and the COEFFICIENT OF PERFORMANCE (COP) of a heat pump. Then, in lecture 25, we will use our thermodynamic gas pot model to calculate the maximum efficiency of any heat engine (or the maximum COP of any heat pump)!

Finally, here is a statement of the 2nd Law of Thermodynamics in its most obvious form:

SECOND LAW OF THERMODYNAMICS

Spontaneous heat flow is only from hot to cold.

We have been using this law without stating it formally since the beginning of our study of thermal physics. On this page, we used it in the argument that the isobaric process is irreversible -- we said that the gas could not get hotter spontaneously while in contact with a colder reservoir (although we didn't use the word "spontaneous", it was implied in our argument -- it only means that no work is being done to transfer the heat). Stated this way, it is rather obvious; but in lecture 25, you will see that it has remarkable consequences.