

Thermodynamics

Thermodynamics and the First Law

Now that we have a good idea about the concept of energy, we can truly begin our study of thermodynamics. Thermodynamics is the study of heat processes and matter. In this chapter, we will be studying the classical approach to thermodynamics, by studying equations and variables of state. What this means is that this chapter is focused on understanding heat and matter processes (primarily gas) by looking at them macroscopically. In other words, we will not actually discuss what each individual particle is doing, we will be concerned with the gas as a whole. We will want to know what properties define the state of the system.

State refers not only to phase (solid, liquid, gas), but also to internal energy in the system and all the variables that can be measured and are of consequence to us in our study of the system. In other words, the state variables refer to the variables that define the gas at that particular instant. The state variables are: pressure, volume, and temperature. If you give a value to each of these three variables, you have defined a state of the system. These three variables are what we really need to know about a system, they define it. What this chapter is all about is how to determine these variables, and how they change during different processes. Of course, we are also interested in whatever information these processes can lead us to: such as the work done, the heat absorbed, etc. From these three variables can flow a wealth of information.

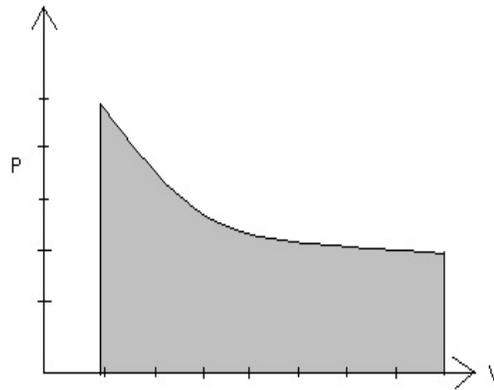
There are three laws that tell us how a system can change state, and these are called the Laws of Thermodynamics. These three laws, like Newton's Laws, have an enormous amount of consequences and can tell us all we need to know about heat processes. The initial law we will discuss (notice I do not say the first), while important in a philosophical sense, is not that important in a practical sense. In fact, it was discovered and formulated after the other three laws were already known. After people had been using these laws for a while, they began to think about them in more detail and noticed that an assumption had been made but never proven or demonstrated. This assumption was tied into all the other laws and if it had been incorrect, so would all of the other laws. Therefore, they made it into a law and put it in front of the other three laws. It is called **the Zeroth Law of Thermodynamics**. It states, "If two bodies are both in thermal equilibrium with a third body, then they are in thermal equilibrium with each other." It is a sort of transitive law for thermodynamics. What it allows is for the existence of thermometers. Thus if a thermometer is placed in a container of liquid and then is placed in another container, if it reads the same in both, we can conclude that the two are in thermal equilibrium (the same temperature).

However, as far as being able to solve problems and have practical applications, the **First Law of Thermodynamics**, is by far more important. It states: "The change in the internal energy of a system equals the heat added minus the work done by the system." Mathematically:

$$\Delta E = Q - W$$

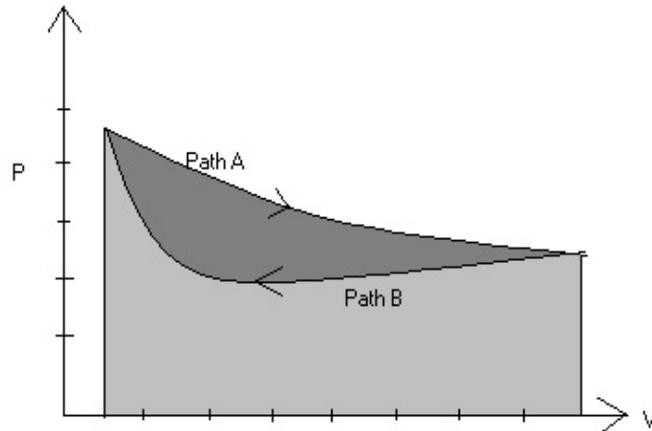
Where Q = heat **added** and W is work done **by** the system **on something else**. An astute student

should already have recognized this law, it is nothing more than the conservation of energy, as applied to a thermodynamical system. Remembering that our goal in this chapter is to find out quantities about a system using our equations of state, our next question is how to relate this first law to things like pressure, temperature and volume. The usual method of doing this is through the work done by the system. The connection is that the work done by the system is equal to the area under the Pressure versus Volume graph. For any thermodynamical system, one very useful piece of information is how the pressure is changing with the volume. Often we show this with a PV diagram, as shown below:



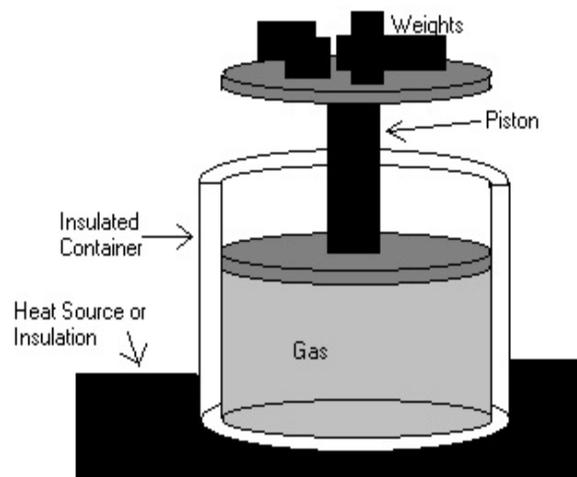
Understanding the PV diagram for a particular situation is absolutely essential for solving problems in this chapter. The main thing to remember is what was mentioned before: the work done by the system on an outside system is the area under the graph. An astute student might want to verify that pressure times volume equals joules and might want to derive an equation that demonstrates that the pressure time the volume is the work done.

It should also be noted that on PV graphs it is essential to indicate which is the initial and final points. Consider the graph above. It could be a graph of a gas dropping in pressure and expanding in volume, or it could be a graph of a gas contracting in volume and increasing in pressure. There is no way to tell simply from the graph, you must be told either by the problem, or more usually, by arrows on the graph.



On the graph above, path A goes from a high pressure, low volume gas to a lower pressure, higher volume gas, while B goes in the reverse. Areas under a graph for a system going from left to right (path A) are positive, while areas under a graph going from right to left are negative. For the sake of illustration, imagine that the light grey area was equal to 14 joules and the dark grey area was 6 joules. The work done going along path A would then be 20 J and the work done along path B was -14 J. A simple conclusion would be that the overall work for this gas was 6 J. This will be explained in more detail later, but the concept is easy enough to grasp.

As we examine systems in terms of the first law of thermodynamics, let us remember that we are interested in the pressure, volume, temperature, internal energy, heat added and work done by the system. To explain how different processes work, we will need to use a sample system. Such a system is pictured below.



In the above situation, the gas in question is contained inside a cylinder fitted with a piston that is free to move. The pressure, volume and temperature of the gas are measured in the cylinder and heat added comes from the heat source below (which can also be considered a heat sink or simply insulation, depending on the circumstances). The work done by the system is the amount of energy expended in moving the weights. For example, if the gas is heated by the source, it

would expand and lift the weights. Likewise, if the bottom is used as a heat sink (where a cold surface draws heat out of the gas) the weights would fall. We will reference this sample set up over and over again the next few sections. There is one other thing about our system that needs to be stated: it is assumed that any changes that place in the system do not happen helter-skelter, but instead happen uniformly. For example, if the system changes pressure, it should change pressure at all points at roughly the same time. If it does not do so, but instead has areas of high and low pressure, we could not define the entire gas as one system with one set of state variables.

Before we go into doing sample problems further, we need to take a short side-track and discuss molar specific heats, since we will be using these to solve some problems. We have learned about specific heat previously, but where gases are concerned, there are a few differences that need to be enumerated. First, when dealing with gases, it appears that mass is not as important (for determining energy considerations) as number of moles is. Thus, for gases, we do not use regular specific heat, but instead we use molar specific heat, abbreviated with a capital C , and having units of $\text{J/mol}\cdot\text{K}$ instead of J/gK . Secondly, unlike solids and liquids, specific heat for gases is very dependent on pressure and volume. In fact, we cannot use one single specific heat for any given gas. We must have one specific heat for cases where the pressure changes and volume is constant (C_v = specific heat at constant volume) and another for the cases where the volume changes and the pressure is constant (C_p = specific heat at constant pressure). In the cases where both the pressure and volume change, we will not be able to calculate energy using specific heat, since the calculations are too advanced for a student at this level.

Thirdly, and perhaps most interestingly, specific heats for gases seem not to depend on the gas itself, but only on its structure. It seems that all gases that are monatomic (such as the noble gases) have similar specific heats ($C_v = 12.5$ and $C_p = 20.8$) while all diatomic gases (H_2 , CO) seem to have $C_v = 20.8$ and $C_p = 29.1$. All polyatomics seem to have $C_v = 24.9$ and $C_p = 33.2$. However, the pattern does begin to deteriorate when we reach the polyatomics. Not all gases meet these criteria. However, especially for the monatomic and diatomic gases, this principle seems to be followed.

To recap these important points, the student should remember that:

- 1.) **When dealing with specific heats and gases, moles are used instead of mass.**
- 2.) **All gases have different specific heats depending on whether the process changes their volume or their pressure. The correct specific heat must be used in each case.**
- 3.) **No matter what the gases are, their specific heats are determined by their structure, meaning whether the gas is monatomic, diatomic or polyatomic.**

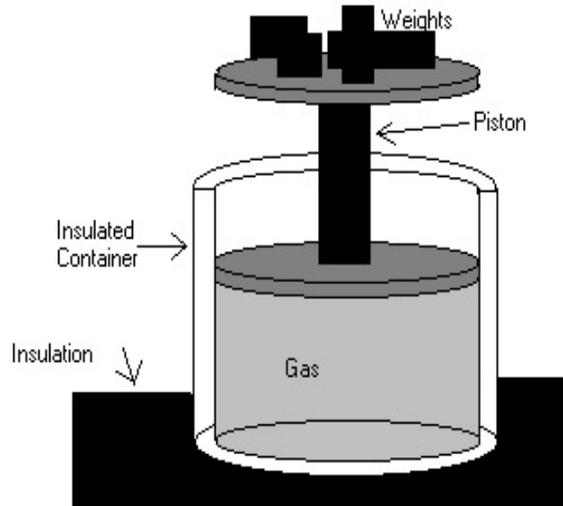
We shall examine these principles again in the next chapter, discussing more about why these results are the way they are. For now we must satisfy ourselves with the conclusions.

Our task now is to look at special situations that can occur in our set up, and use our knowledge of gases, along with the first law to draw up some conclusion. We will look at six separate situations, each with its own name. While they are all important, there are a few that are more important than others. These are astericked. The student should memorize the names

of these processes and have a good solid knowledge of them.

Adiabatic Processes*

An adiabatic process is one where there is not heat added or taken from the system (from the Greek word for “impassable”). In this case, we expect the bottom surface to be insulated.



From the first law, we see that since $Q=0$, $\Delta E = -W$. If no heat is added, then any work that is accomplished is done so at the expense of the internal energy of the gas. Since the system seems to be in equilibrium at the moment, the only way that work could be done would be to change the amount of weight on the top, so that the system would change. If weight were added, it would compress the gas, and since the weights would go down, it would be negative work done by the system. The second negative would make the change in internal energy of the system positive. If weights were taken away, the gas would expand and lose energy, as the weights gained gravitational potential energy.

Consider the second situation from the standpoint of the ideal gas law, which you remember from chemistry. The law said:

$$PV = nRT \quad \text{Or} \quad PV/T = nR$$

In the second form, nR must be constant for any gas, thus if some gas went from state 1 to state 2 by some thermodynamical process, we could write:

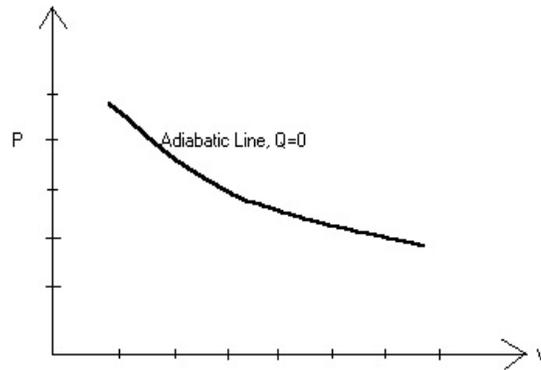
$$P_1 V_1 / T_1 = P_2 V_2 / T_2$$

The pressure of the gas is directly related to the weights on the piston, and the volume of the gas to the height of the piston. The temperature relates to the gases internal energy. If the pressure goes down, the volume goes up and the temperature goes down, the equation could be satisfied. Notice, however, that we cannot say for sure that the temperature must go down. Depending on

the other two variables, it could go either up or down. However, we can say that the temperature going down is certainly a possibility.

In real life, adiabatic processes occur very often. However, most of the time that they occur, they occur because the process happens so fast that heat cannot be transferred, instead of because the system is insulated. In an automobile engine, for example, when the fuel is ignited, it expands so quickly (pushing down on the piston) that very little energy has time to escape. Another example is a hand operated bicycle pump. In that case the pumping is often so fast that little heat can escape from the system between pumps and the pump heats up. Notice that in this case it is not that the system does not heat up, but that heat does not leave.

On a PV diagram, an adiabatic process is represented by what is called an adiabat, which is a line similar to the one shown below:



Adiabatic expansion follows an adiabat whose equation is:

$$pV^\gamma = \text{constant.}$$

We will discuss this equation in more detail in a later chapter. Before we move on, however, it should be noted that many different processes produce PV lines that resemble adiabats. Without being told, there is no way for a student to recognize when a line is an adiabat and when it is not.

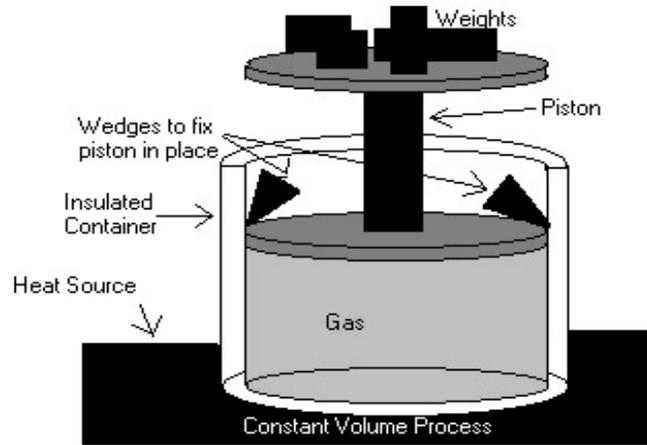
EX HYN.) Suppose a gas lifts a mass of 2 kg up 15 cm in the air and does so adiabatically. By how much has the internal energy of the gas changed?

Constant Volume (Isochoric) Processes

The next process of interest is constant volume processes. In this case, the base of our example container will be used to either absorb heat from the gas or to add heat into the gas. The only way for our system below to remain at a constant volume is if the piston is fixed in place. In such a situation, the gas could never raise or lower the weights and thus:

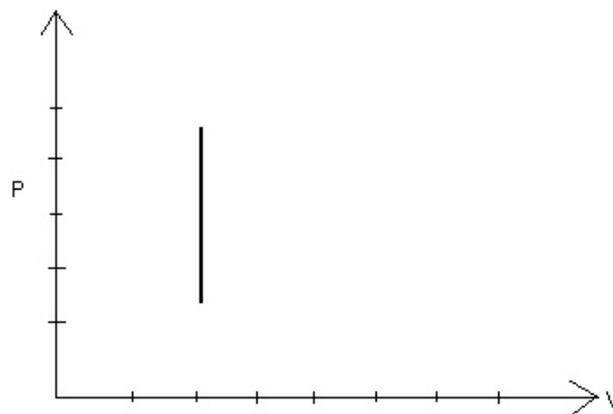
$$W = 0 \quad \text{and thus} \quad \Delta E = Q.$$

Which makes sense, since whatever heat is added to the system must go directly to changing its internal energy.



The PV diagram simple to draw:

for this process is very



Notice how the for no area to be

vertical line allows “under the graph”,

which is what we would expect since no work can be accomplished. In this case, looking at our old equation of:

$$P_1V_1/T_1 = P_2V_2/T_2$$

Since the volume doesn't change, we get:

$$P_1/T_1 = P_2/T_2$$

Since our volume is constant, this is one of those rare cases where the molar specific heat is constant and easy to work with. Thus, since the heat that goes into the gas goes directly into raising its internal energy, we get

$$Q = \Delta E = nC_v\Delta T$$

EX TRCV.) If a 6 moles of gas with a molar specific heat of 12.6 J/mol.K is taken from a pressure of 500 Pa and temperature of 30° C to a higher temperature and pressure isochorically by the addition of 1000 J of heat energy, what is the new pressure and temperature?

Isobaric Processes

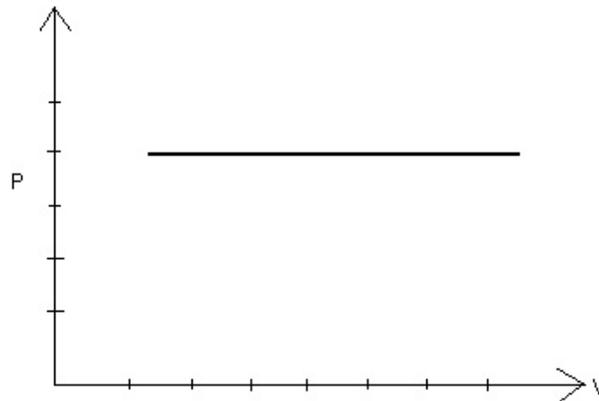
Another type of process that is worth examining is an isobaric process. Isobaric stands for same pressure, thus this is a process carried out by a gas at a constant pressure. An isobaric process is not one that can eliminate any our variable in the first law, thus it leaves it as:

$$\Delta E = Q - W$$

however, it can give us one equation, since the pressure is constant:

$$\Delta E = nC_p\Delta T$$

This could help us solve quite a few problems. The PV diagram for an isobaric process is as follows:



So we can see that the work done is easy to calculate (the area under the graph, which is simply $p\Delta V$). And we can also see that our ideal gas law becomes:

$$V_1/T_1 = V_2/T_2.$$

EX PFN.) Suppose 2 moles of a gas, held at 500 Pa expands from 0.8 m^3 to 1.4 m^3 . If it has a specific heat of 20.8 J/mol.K (at constant pressure), what was its change in internal energy?

Isothermal

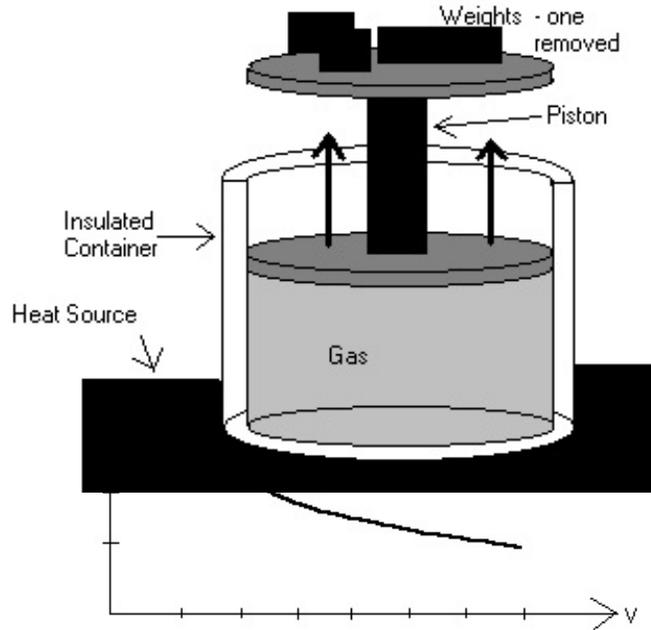
As the student can probably guess, isothermal processes are those where the temperature is held constant. In such a case, since the temperature does not change, the internal energy cannot change. Therefore we have:

$$\Delta E = 0 \quad \text{and} \quad Q = W$$

from the first law of thermodynamics. Our ideal gas law then gives us:

$$P_1 V_1 = P_2 V_2$$

and our PV diagram



looks like this:

Notice this diagrams
adiabatic (no heat

There is no real way to tell these two lines apart simply by glancing at them. You must be told which is which. There is one other formula that goes along with isothermal expansion, but it requires calculus for its derivation. We will not derive it, but we will state it for the student to use.

similarity to the
added) drawing.

$$W = (nRT)\ln(V_2/V_1)$$

The above equation is only good for ideal gases at constant temperatures.

An example of an isothermal process would be the following: a weight is removed from the top of our piston, while at the same time heat is supplied from the bottom. This causes the volume to increase, the pressure to decrease and the heat being supplied causes the temperature to remain the same in the process.

EX UNH.) 2 moles of a gas goes isothermally from 2000 Pa to 3500 Pa beginning at a volume of 0.2 m^3 . What is its final volume? How much heat was lost?

EX HU.) Using the ideal gas law, derive an equation for the pressure of a gas as a function of its volume at constant temperature. Graph this for the following situations:

- A.) $n = 2$ moles, $T = 270 \text{ K}$
- B.) $n = 2$ moles, $T = 300 \text{ K}$
- C.) $n = 2$ moles, $T = 330 \text{ K}$

(These lines are called isotherms.)

The next two examples of thermodynamic processes are variations from the kinds of processes we have outlined. The first is a combination and the second is one that we cannot graph.

Cyclical Processes

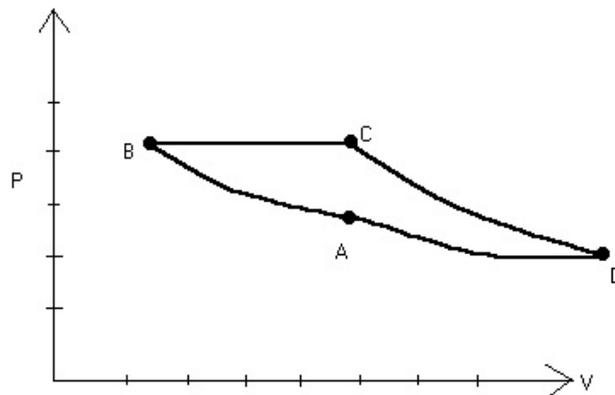
One of the very important processes that can happen in a thermodynamical system is a cyclical process. In this case, the gas will carry out a number of processes, and eventually arrive back at the same state (P,V,T) that it originally began. The cycle consists of a number of processes (from 2 to 4) and each one is usually one of the processes described above. Such a cycle is often carried out in engines or refrigerators. In these cases, the cycle is used to cause the gas to do work on the outside environment. Since the gas begins and ends at the same state, its internal energy cannot have changed (for the entire cycle). Thus, applying the first law of thermodynamics gives:

$$\Delta E = 0 \quad \text{and} \quad W = Q$$

The work that is accomplished is equal to the heat that is supplied. Makes sense. An example of this would be using our set up in the following manner:

- 1.) An extra weight is placed on the piston, causing it to contract (A to B)
- 2.) Heat is added and the piston raises the weights up to the original height (B to C).
- 3.) The weight is removed and the piston raises (C to D).
- 4.) The gas is slowly cooled down until the piston is back to the original position (D to E).

Now the gas is back in its original form, and the process could be repeated. A PV diagram for this situation would be:



The student should look at this diagram understand it. Of

work done in each step. Notice how some of the work is positive and some is negative. Overall, the work done is positive. What exactly was that work? Where did the energy come from?

Understanding cyclical processes is very important, thus we will take a look at three examples.

take a few minutes to and be sure they special note is the

EX UHN.) Consider a hypothetical engine that operates on four strokes in a cycle. The strokes are alternating isobaric and isochoric processes. If the engine operates between pressures of 200 Pa and 350 Pa and volumes of 0.05 m^3 and 0.08 m^3 , how much work does the engine do? Assume the engine does positive work on the environment. Draw a PV diagram for such an engine.

Before we leave this section, we should take a few minutes to discuss how this relates to our own automobile engines. These engines are called four-stroke, since they have four phases. However, the first and last phases of the cycle do not count since they introduce new gas and take gas out of the cylinder. When this happens, our previous discussion is not longer valid for these things.

The first cycle is that gasoline and air are sprayed into the cylinder as the piston moves down. Secondly, the piston moves up and compresses the gas adiabatically (it is adiabatic since it happens so fast that no heat has any time to transfer). Thirdly, the fuel ignites and pushes the piston down, once again adiabatically. Lastly, the burned up fuel is ejected from the cylinder, leaving it empty and ready to repeat the cycle.

EX UNH.) Draw a PV diagram for the middle two strokes of an automobile engine.

Free Expansion

The last type of thermodynamic process to discuss is free expansion. A free expansion is an adiabatic process where no work is accomplished. An example of this would be if we insulated our cylinder and then rapidly pulled the piston up to a higher level. This would cause the gas to expand into the new volume, not doing any work and not absorbing any heat. The reason no work is done in this case is that the gas did not push the piston, we did (remember in our other cases we removed a weight and let the gas push the piston up). Since no work is done and no heat absorbed:

$$Q = W = \Delta E = 0$$

The only equation we can then use is the idea gas law, when the temperature does not change (since $\Delta E = 0$).

Entropy

We have discussed the concept of state variables, or variables that tell us something about what state or condition the system is in at a particular time. Besides the pressure, temperature, volume, and internal energy of a system, there is one other very important variable. Its introduction has been put off until now because of its difficulty and its complexity. Very often it is misunderstood or not fully understood by students and teachers alike. It is called entropy. Entropy has many different definitions, depending on the author of the book being read. Although all of these definitions are correct, each has a different strength and weakness. We will begin this section by defining entropy a number of different ways, first conceptually, and then mathematically. The best and most concise definitions of entropy are the mathematical ones, one of which is given in this chapter and another of which is given in the next chapter.

Entropy is most commonly defined as “a measure of the degree of disorder in a system”. It is a measure of organization. Highly ordered things (like a book shelf neatly organized in a library) have low entropy and messy things (like all the same books laying pell mell on the ground) have high degrees of entropy. **The Second Law of Thermodynamics** states that the entropy of a system must always either stay the same (if the process is reversible) or increase. In a sense, it says that systems must always tend towards disorder. You cannot create greater order in the universe, without causing more disorder somewhere else. For example, you may be able to reorganize those previously mentioned books, but in doing so your body creates more disorder and entropy than was there previously (by using food, breathing, etc.).

Notice that the second law differentiates between what are called reversible and irreversible processes. A reversible process is one that can be reversed, leaving the system and the surroundings, the same as before the process occurred. For example, knocking the books to the floor is not a reversible process, since it takes more energy for you to lift the books back up than it did to knock them down. The system (the books) are the same after the process, but the environment (you) are not. However, imagine that you used 10 J of heat to increase the energy and volume of a gas. If you could then cool that gas back to the original volume and extract 10 J back out, it would be a reversible process. The gas would be back to its original state and the

environment would have given up 10 J and then received it back.

Another definition of entropy is “a measure of the unusable heat produced in a process”. In all reversible processes, there is some useless heat produced. We have discussed this idea previously in our discussion of energy, when we said “there is no such thing as a free lunch”. Every time energy is exchanged (in an irreversible process) there is some useless heat produced. Entropy, in a sense, measures this. Entropy is useless heat. It is heat we cannot get back. In irreversible processes, entropy is produced (increases). In reversible processes, entropy stays the same and no useless heat is produced. By thinking about this, we can see that reversible processes are an ideal. In reality, we cannot create a reversible process, since that process would be 100% efficient. NCD.

Another definition of entropy is that entropy is “the arrow of time”. That rather vague and imposing definition requires a bit of explanation. Imagine dropping a glass of water. If it hit the ground and broke, it would shatter and spill. Now imagine that the glass miraculously reordered itself and the liquid went back into the cup. Both processes are the same, only they are the reverses of each other. So why does one happen and the other does not? Or imagine this: you drop a ball (in a vacuum, but with gravity) and it falls to the ground. Now imagine that the ball, of its own accord, magically goes back up to the same height (not on the bounce, simply on its own). Both situations completely satisfy the conservation of energy (think about it). So why does one happen and the other does not? There are many such situations in nature, where no laws would be violated, yet the phenomena does not occur. Sand castles are built and the waves wash them away. Why don't waves ever wash up complete sand castles on the shore? The answer appears to be entropy. The glass will not rearrange itself simply because to do so would mean an decrease in entropy. No other laws would be violated. The ball will not simply rise up to some height, to do so would mean an decrease in entropy. A sand castle will not wash up on the shore, to do so would mean a decrease in entropy. In this way, entropy tells up which way time will point and events will occur. Entropy, in this sense, also provides another important piece of information: it tells us that some events simply will not happen. These, like the examples given above, are events for which there is no other possible explanation for their non-occurrence. Entropy is the only thing standing between us and a very different and strange world.

We now turn to the mathematical definition of entropy, which is more useful in solving problems, but unfortunately does not paint a complete conceptual picture of entropy by itself. It is:

$$\Delta S = Q/T$$

The above equation gives the change in entropy for any process, by it is not easy to use unless the process occurs at a constant temperature. If the temperature is not constant, then we must break the situation up into many, tiny segments over which the temperature is roughly constant and add them together. The best way to do this is to break it up into infinitesimally small segments and add them all up (humph, sounds like calculus!). A few notes should be made before we progress.

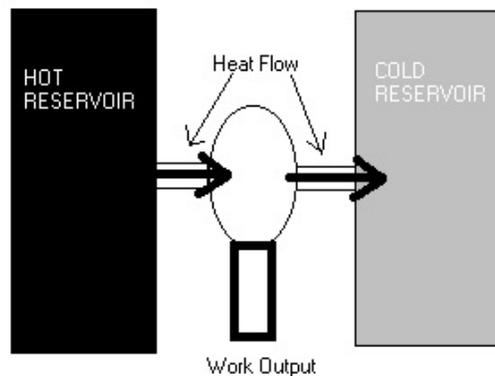
- 1.) Entropy is a state variable, thus it only depends on the initial and final states, not on the process that takes the system from one to the other. Therefore, reversible and irreversible processes that begin and end at the same states have the same entropy changes. The equation given above actually only applies to reversible processes, but since the entropy changes are the same, we will substitute a reversible process for the more difficult irreversible processes.
- 2.) Entropy applies to any changes in the universe. However, we will only consider problems involving gases and heat exchanges. These are the most obvious and most direct uses of entropy and they are the ones that demonstrate the principle the best.
- 3.) The total entropy of a system is the sum of all the entropy changes in the system. Thus if more than one thing is involved, or if the process requires a number of steps, we can calculate them separately and add them together.
- 4.) Remember that the above equation only works at constant temperature processes (isothermal).

EX JNH.) If 10,000 J of heat energy moves through the wall of your house to the outside, what is the increase in entropy of the universe? Take the temperature inside the house to be a constant 27° C and outside to be 15° C.

EX UBNB.) If a hot (80° C) piece of copper (specific heat = 0.386 J/gC°, mass 200 g) is thrown into a cold (10° C) lake, what is the change in entropy for the lake?

The Second Law and Heat Engines

We have seen in a previous example, how a engine can be constructed between a hot and cold reservoir to produce work. Such a device is called a heat engine. In this section we will look at how these work and how the second law applies to these devices. A basic drawing of such an engine is shown below:



The basic idea is that heat moves from the hot to the cold reservoir, (something it does by itself) and it can be intercepted and used to produce work. Like many other devices, a heat engine could be operated in reverse. In this case, work would be put into the device (electricity, for example) and heat would be moved from the cold reservoir to the hot side. To aid our discussion, let us introduce the following variables:

T_H, T_C = temperature of the reservoirs

Q_H, Q_C = Heat removed or added to hot and cold reservoirs

W = Work done by engine

ξ = efficiency of engine

Since heat is coming into the engine according to Q_H and leaving the system according to Q_C , there is a difference in energy equal to $\Delta Q = Q_H - Q_C$. This is available to produce the work done:

$$W = Q_H - Q_C$$

(This can also be arrived at by the first law of thermodynamics)

Efficiency is always what you get out divided by what you put in, thus

$$\xi = \text{out/in} = W/Q_H = (Q_H - Q_C)/Q_H = 1 - Q_C/Q_H$$

At first glance, this equation seems to say that the efficiency of a heat engine must be 1 or less. It would seem that the perfect engine would be one that take out a quantity of heat (Q_H) and turns it all into work, putting no heat out on the cold side (Q_C). However, we have forgotten the second law. The total entropy must (at best) remain constant or must increase. This system has two entropy changes:

$$\Delta S_1 = \text{entropy loss of } Q_H \text{ leaving hot reservoir}$$

and

$$\Delta S_2 = \text{entropy gain of } Q_C \text{ entering cold side}$$

This tells us immediately that to have no heat leaving the cold side would violate the second law, since out entropy would go down. The **absolute best we could ever hope for would be:**

$$\Delta S_1 = -\Delta S_2$$

This ends up giving us:

$$Q_H/T_H = -Q_C/T_C$$

Dropping the negative and rearranging gives:

$$Q_C/Q_H = T_C/T_H$$

and combining with the above result:

$$\xi = 1 - T_C/T_H$$

A simple, but interesting result. If we examine this for a while and ponder it, we will reach the following conclusion: Nature itself imposes a restriction on the efficiency of an engine. No matter how much friction you eliminate, no matter how perfect the engine, you will never get 100% efficiency. There is no such thing as a free lunch. If you want work out of a device, you must sacrifice some of the input energy. The only way to get a perfect heat engine would be to make $T_c = 0$ or $T_H = \text{infinity}$, neither of which are practical suggestions.

Thus, any heat engine, operating between two temperatures has a built in limit to its efficiency. An automobile engine, operates between two heat reservoirs, the burning gas (at about 600 K) and the outdoors (about 300 K), giving it a maximum efficiency of about 50%. Thus if all the friction were eliminated in your car, the best your engine could be would be about 50%.

All of the above information and equations applies precisely to refrigerators and air conditioners as well as heat engines. In a refrigerator, the heat is removed from the “cold” side and pumped out to the “hot” side. For a common refrigerator, the cold side is inside and hot side

is out into the room. For an air conditioner, the cold side is in the house and hot side is outside the house.

Before we begin our example problems, we need to once again define the second law of thermodynamics. Very often the second law of thermodynamics is stated as “There is no such thing as a perfect engine.” This is called the Carnot Principle named after a man who did much of the pioneering work in thermodynamics. The idea of no perfect engine existing is important enough to be used as a statement of the second law.

Let us begin our example problems by trying a few conceptual questions:

EX JKU.) During what season should you expect to get the best gas mileage from your car?

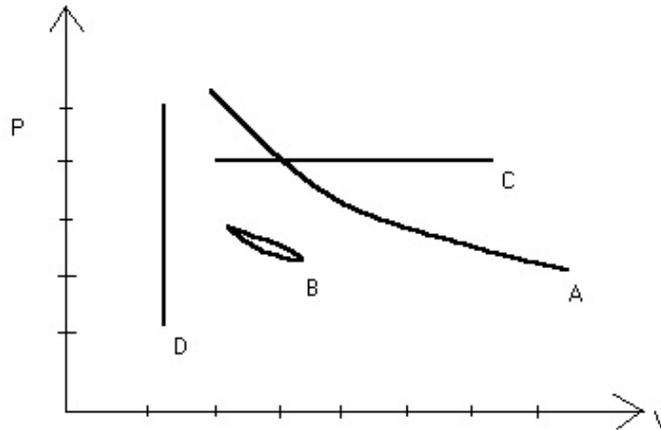
EX HU.) If you wanted to cool off your room, would it work to open the refrigerator door?

EX UJH.) During what days is your air conditioner most efficient? Does that mean it costs less to keep the house at a steady temperature during those days?

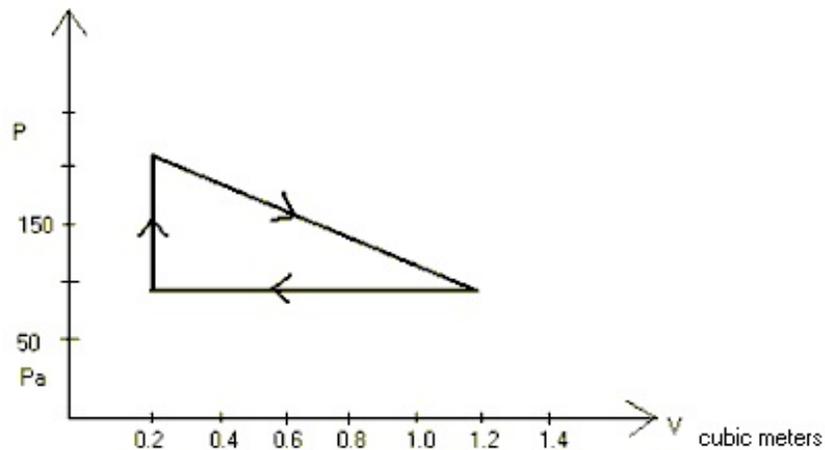
EX JUH.) What is the maximum efficiency of a refrigerator operating between the interior (5°C) and the room (28°C)? If a 2 L bottle of soda is placed in the refrigerator, how much energy (electrical) is required to cool it down? Assume the soda to be mostly water (4.19 J/gC°) and beginning at room temperature.

Assignment #37

- 1.) The graph below shows four different thermodynamic processes. Identify which line shows a(n):
- Isothermal process
 - Cyclical process
 - Isochoric
 - Isobaric



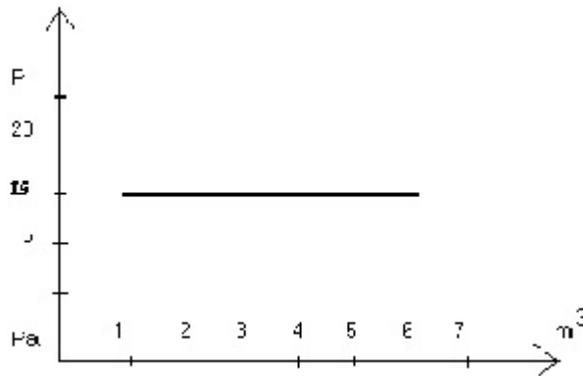
- 2.) How much work is done in one cycle by the process shown below?



- 3.) (Skip this problem) Imagine that a balloon at room temperature has a radius of 20 cm.

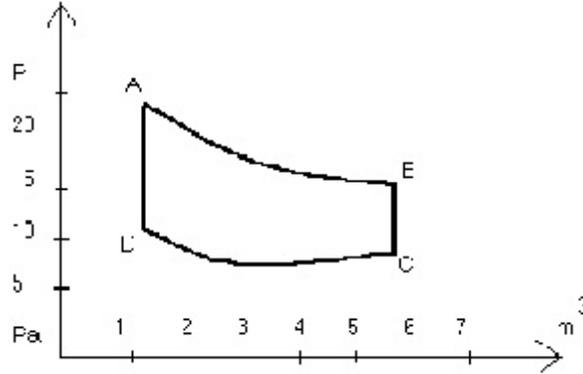
and is at a pressure of 2 atmospheres (1 from the air, 1 from the elastic in the balloon). It is quickly submerged into a pool, to a depth where the water pressure is 1.2 atmospheres, what is its new volume? What is its change in internal energy? What is its new temperature? Throughout this problem, assume the elastic pressure does not change significantly and that the balloon contains 2 moles of gas and that 2 J of work are done.

- 5.) A gas is trapped in a bottle (1.5 L) and placed in a freezer, where its temperature goes from room temperature to 5° C. If it is 3 moles of a diatomic gas, how much heat was exchanged between the gas and the freezer?
- 6.) Draw a PV diagram for the above process (include direction).
- 7.) Helium gas is trapped in a 0.07 m³ balloon at a pressure of 2 atmospheres. The gas is taken down to a volume of 0.06 m³ as its temperature drops. If the volume change was not enough to appreciably change the pressure from the elastic walls of the balloon, how much did its temperature change if it began at room temperature? How much heat did the balloon exchange with its environment?
- 8.) If 10 J of heat was used to cause the gas below to undergo an isobaric process, by how much did its internal energy change?



- 10.) A cyclical process operates on the following processes: 1.) Isobarically at 50 Pa from 0.3 m³ to 0.8 m³, 2.) Isochorically down to 20 Pa, 3.) Isobarically to 0.3 m³, and 4.) Isochorically back to its starting point. Determine the work accomplished and draw a PV diagram.

- 11.) Determine what type of thermodynamic process is represented by each leg of the cyclical process shown below. Also determine the work done (approximately) by the process. Hint: the temperature changes during each process.



- 12.) Suppose a 100 g ice cube melts at room temperature and then the water heats up to equilibrium with the room. What is the change in entropy for the room? (Heat of fusion = 334 J/gC°, c = 4.18 J/gC°)
- 13.) If a house is maintained at 75° F by an air conditioner on a day when the outside air is 95° F, what is the total change in entropy of the house/air system if 10,000 J of heat escapes by conduction through the walls and windows?
- 14.) During an isothermal (45° C) process, 6 moles of an ideal monatomic gas is taken from one state to another, using 65 J of energy. What is the change in entropy of the gas?
- 15.) During an isochoric process, 4 moles of an ideal monatomic gas is taken from 2 atm at 60° C to 30° C. If the heat is given off to a reservoir held at 30° C, what is the change in entropy of the reservoir?
- 16.) What is the maximum efficiency of an engine operating between reservoirs of 800 K and 300 K?
- 17.) A freezer operating between room temperature and -5° C is used to freeze 1 kg of water into ice (at the same temperature). The water is placed in the freezer at room temperature. Assuming maximum efficiency, how much energy is required by the freezer during that time? (c of ice = 2.22 J/g°C)

- 18.) An engine uses 200 J of energy to provide 60 J of work and is 60% as efficient as an ideal engine operating between the same reservoirs. If the cold reservoir is 30°C , what is the temperature of the hot reservoir?
- 19.) An ideal engine extracts 30,000 J of energy from a reservoir at 400 K and expels energy into a reservoir at 300 K, how much heat is expelled? What is the change in entropy at each side of the engine? How much energy is required to do 500 J of work?