
I.4. THERMAL ENERGY TRANSPORT: Useful Background Information

Some Basic Thermodynamics:

There are two concepts to keep straight about thermal energy. There is heat, and then there is heat transfer. Heat transfer occurs via radiation, conduction, or through phase changes. Radiation is discussed in class and in the chapter on radiation physics in Eagleson. Conduction is straightforward, as discussed in class. The heat transfers associated with phase changes requires a little bit more detailed understanding of heat itself, and a few basic concepts of thermodynamics. This is a brief review for those who may not have these concepts right at their fingertips.

HEAT: Heat is a form of energy, and therefore it is an entity that is either conserved or converted to some other form of energy. An object can gain heat, it can lose heat, or it can store heat in some form. You can always (in principle) specify how much heat an object contains at any moment. An analogy is wealth. You can gain wealth, you can lose wealth, or you can store it in one of several forms. You can (in principle) specify how much wealth a person has at any moment. Just as wealth occurs in several forms (cash, paintings, Ferraris, education) which are interconvertible, so heat occurs in several forms, which are interconvertible. The amount of heat is not the same as the amount of heat transfer, just as spending money is not the same thing as being wealthy.

TWO MAIN FORMS OF HEAT: Sensible Heat and Latent Heat

SENSIBLE HEAT: As the name implies, this is the heat you can feel. The sensible heat possessed by an object is evidenced by its temperature. The higher the temperature, the greater the sensible heat content. However, for a given change in sensible heat content, not all objects change temperature by the same amount. Each substance has its own characteristic relationship between heat content and temperature. The factor that is the proportionality constant between temperature rise and change in heat content is called the Specific Heat. It has units like calories per gram per degree celsius, or joules per kilogram per Kelvin. Water, for example has a specific heat of $1 \text{ cal/g/}^\circ\text{C}$. This means that if a gram of water increases in temperature by 1°C , its sensible heat content has gone up by one calorie. It takes 75,000 cal (75 kcal) to raise 1 L (= 1 kg) of water from 25°C to 100°C . (As an aside, the dietary "Calorie" [capital C] is actually equal to 1 kcal; a candy bar with 250 Cal has an energy content of 250,000 cal).

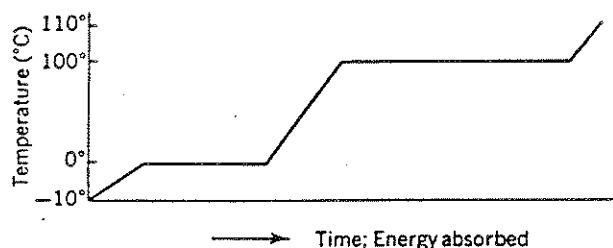
In general, the gain in heat is accompanied by either a change in volume or a change in pressure (e.g., the water in the pot swells somewhat as you heat it; if you heat gas in a fixed volume, its pressure goes up.) Thus, to be precise, thermodynamics requires that you specify which is going to be held constant when you define the specific heat: constant pressure or constant volume. If you want the volume to remain constant you use C_v , the specific heat at constant volume. If you want the pressure to remain constant, you use C_p . For water, the pressure normally remains constant while it almost impossible to keep the volume constant (as we shall see later, water is an "incompressible" fluid). Hence most of our equations for water will use C_p . (The value of $1 \text{ cal/g/}^\circ\text{C}$ given above as the "specific heat", is actually C_p). For air in the atmosphere,

both pressure and volume can change, so either C_v or C_p may appear in an equation. We will not need to worry too much about it, but remember that if air is changing both in volume and in pressure, an exact equation will have extra terms that account for both changes, and both C_p and C_v will appear.

LATENT HEAT: When a solid turns into a liquid (melts) or a liquid turns into a gas (evaporates), the loosening of attractions among the molecules requires energy. If you raise ice to from -20°C to 0°C , you put in sensible heat. If you keep adding heat to the ice, it melts but its temperature does not go up, it stays at 0°C . As long as the temperature is constant, the sensible heat of the ice/water system is not increasing, yet you keep putting heat energy in. Energy is conserved, so where is it going? The extra heat is going into tearing apart the frozen ice molecules and setting them loose as a liquid. The liquid water is therefore storing this energy in a form that you cannot sense. This energy is called latent heat.

To melt all the ice, you have to pump in quite a bit of heat, but you cannot sense any change in the heat content because the ice/water system remains at 0°C . Only after all the ice has melted does the temperature of the water rise, at which point the heat you put in is once again creating a change in sensible heat. However the liquid still has stored all that latent heat, even though you cannot feel it. The only way you will observe that latent heat again is if you try to turn the water back to ice. If you take the temperature down to 0°C , that alone will not freeze the water; you must keep pulling out heat until you have removed every joule of latent heat. Only then will all the water freeze and only then can you start removing more sensible heat and lower the temperature of the system below 0°C .

The relationship between heat input and temperature for water is summarized in the diagram below:



When water at -10°C is converted into steam at 110°C , two plateaus are noted, one for each phase change. The lengths of the plateaus are not drawn to scale.

As you can see in the figure, the same sort of temperature plateau occurs when water evaporates: it may stay at the same temperature (no sensible heat gain) but it takes up a lot of heat energy and stores it as latent heat in the vapor. When you condense out the water, you must remove that heat.

Heat energy is conserved no matter how the phase change occurs. If you put heat into water it can evaporate. If you do not add heat and it evaporates on its own, the water will cool off; i.e., some of the sensible heat is lost and converted to latent heat. Conversely, if you cool off some water vapor, it can condense into liquid. If it condenses on its own, it will give off (sensible) heat and get warmer.

The amount of latent heat per gram required to effect a phase change is called either the latent heat of fusion (melting) or else the latent heat of vaporization. This is a substantial amount of heat for water. In fact, water is very peculiar in having such enormous latent heat values. [I have a section from a book chapter I am writing that deals with the reasons for the curious properties of water; if you have a detailed interest in the structure of water, ask me for a copy].

Consider that to raise the temperature of a gram of liquid water by 1°C requires only 1 cal. In contrast, to melt one gram of ice requires 80 cal, and to evaporate 1 g of water at 100°C requires 540 cal. In our example above it took 75 kcal to raise 1 L of water from 25°C to 100°C. To evaporate that liter of water at 100°C requires an additional 540 kcal. The water at 100°C has 75 kcal more heat than the water at 25°. The steam at 100°C has 615 kcal more heat than did the water at 25°C. Let's put this into everyday experience. If you stick your hand in hot water at 100°C, you'll be exposed to ~75 kcal of excess heat; if you stick your hand into steam at the exact same temperature, you'll be exposed to ~615 kcal of excess heat. (Hence the well founded fear of steam burns. This difference in energy is also the reason you can run a locomotive on steam, but not on hot water.)

The latent heat of a phase change depends on the temperature at which the phase change occurs. In nature, freezing always occurs at 0°C (or a little below that for seawater) so you only need to know one value of the latent heat of fusion for water. However, evaporation can occur at any temperature between 0°C and 100°C so there is a continuous range of latent heats of vaporization. Eagleson's text "Dynamic Hydrology" (p. 60; OGI library; my office) gives a formula for computing the heat of vaporization at any temperature.

One other phase change is possible besides melting/freezing and evaporation/condensation. Ice can sublime directly into water vapor (as it does in a "frost-free" freezer or in a freeze-drier). Sublimation occurs in nature, and snow and ice can convert directly to the vapor phase. The latent heat of sublimation depends on the temperature of the ice and a formula is given in Eagleson's book on p. 60.

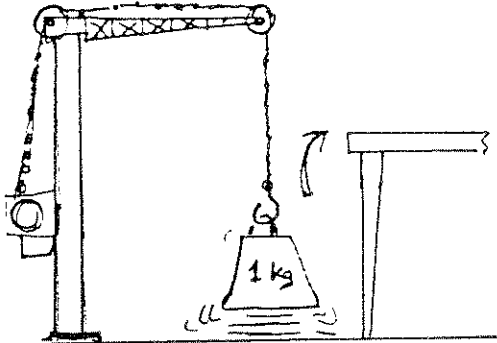
From the standpoint of this class, you want to keep in mind that when water evaporates from the surface of the earth, it carries a tremendous amount of energy with it out of the water and up into the atmosphere as latent heat. Conversely, when water precipitates as rain or snow, it releases enormous quantities of energy to the atmosphere.

Make sure that you have a clear understanding of the difference between sensible heat and latent heat. The analogy depicted below may be useful.

1.

Put some *mechanical energy* into an object:

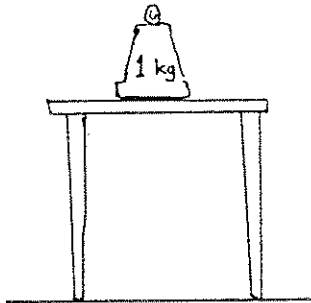
$$E_m = \frac{1}{2} mv^2 = F \cdot d$$



2.

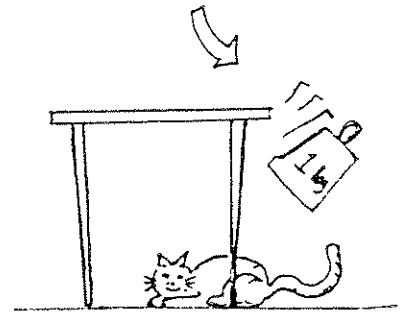
Potential energy is now stored in a motionless object.

$$E_p = E_m$$



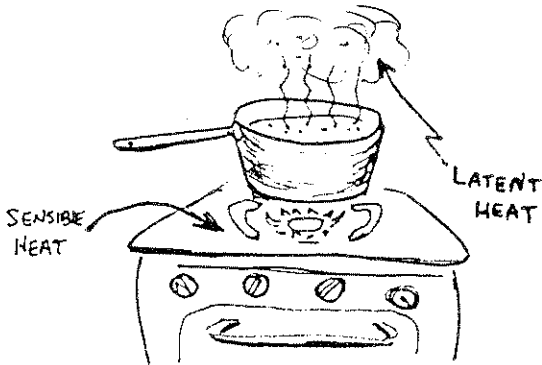
3.

We can generate *mechanical energy* and simultaneously lose the *potential energy*.



1.

Put some *thermal energy* into a pot of water. Above 100°C the temp. does not rise further and all energy goes into converting liquid into gas. (We are adding *mechanical energy* to the molecules).



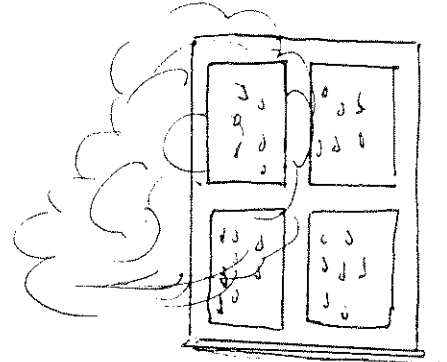
2.

The cloud of gas (water vapor) stores the heat energy you put into it. That energy is now *latent heat*: you cannot feel it. The cloud can be room temperature but the *latent heat* is still there.



3.

If the gas hits a cold window the *latent heat* is converted back to *sensible heat*: the window gets warmer (you can "sense" the heat) while the gas converts back to the liquid state (condensation).



- Evaporation requires heat: it cools things off.
- Condensation releases latent heat: it warms things up.