UNIT C

HEAT, TEMPERATURE, AND CLOUD FORMATION

DETAILED CONTENTS

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Nature, it seems, is the popular name
For milliards and milliards and milliards
Of particles playing their infinite game
Of billiards and billiards and billiards.
--Piet Hein

0 OBJECTIVES

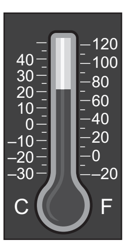
1. To observe and describe temperature differences and compare these with our sensations of hot and cold.
2. To understand why a thermometer behaves the way it does and to understand the relationship between different temperature scales by constructing our own thermometer.
3. To develop a model for how thermal energy is transferred between objects and to understand how the transfer of thermal energy causes the temperatures of objects to change.
4. To build an understanding of phase changes of matter, such as the melting of ice and the boiling of water.
5. To recognize the difference between wet-bulb and dry-bulb temperature readings and understand how this difference provides a measure of humidity.
6. To use the concepts of evaporation, condensation, and relative humidity to understand the principles of cloud formation and investigate the conditions under which clouds will form.
7. To learn more about the nature of heat, temperature, and the process of scientific research by undertaking an independent investigation.

0.1 OVERVIEW

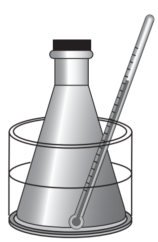
Everyday we engage in activities that are influenced by heat and temperature from drinking a glass of iced tea to feeling chilled after swimming on a hot day to stepping through a fogged bathroom after a hot shower. Even so, we rarely think carefully about the concepts of heat and temperature and how they affect our daily lives. If you want to cool down a glass of tea, how much ice do you need to add? If you add two ice cubes, will it cool the tea down twice as much as if you added only one ice cube? How much will drinking the iced tea actually cool your body. Why do you get chilled after swimming even on a hot day? What makes the temperature you "feel" seem different even if the actual temperature hasn't changed? Why does fog form in your bathroom when you take a hot shower? Will opening the window in your bathroom prevent the fog from forming or make it worse? In this unit, we will investigate the concepts of heat, temperature and humidity with the aim of learning how to answer these kinds of questions.

In addition, you will learn how these concepts fit together to explain other everyday phenomena. In the process, you will have an opportunity to work on experiments to help you understand the nature of these concepts. In the end, you will even make your own cloud. Some of your work will be done independently and some in small teams. You will likely learn the most when you are engaged in discussions with your partners. Debating your ideas will lead all of you to a more solid understanding of the concepts under study. So don't blindly work through the experiments. If you don't understand something, speak up and challenge your partners to explain it to you. The ensuing discussion will benefit you all.

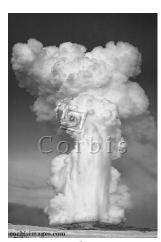
The path that we will be taking in this unit is a bit complicated so it is useful to break it down into smaller, more manageable pieces. First we will look at temperature and how we measure it. Then we will investigate how two objects of different temperatures interact. Next we will study how a substance, in our case water, changes form as it is heated. This kind of transformation is known as a *phase change*. Finally, we will tie all of this together to explain one of the most ephemeral of phenomena, the cloud. This is all summarized in Figure C-1.



Sections 1 & 2: How does a thermometer work?



Sections 2 & 3: How are "heat" and temperature related?



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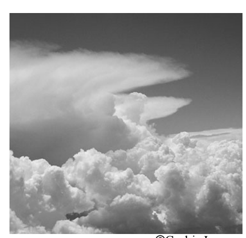
Sections 3 & 4: How are evaporation and humidity related?



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Section 4: What makes a cloud form?

Figure C-1: The main questions we will be tackling in this unit.

As we carry out our investigations, it helps to have some specific questions we would like answered. As already mentioned, we all know when something is "hot," but this doesn't describe things specifically enough to be useful. A natural first question might be *how does a thermometer work and what does it tell us?* As you know, you can increase the temperature of an object by "heating it up" (in an oven, for example). But how that happens may be a bit of a mystery. Therefore, another question is *how does heating something up produce a temperature change?* A third common experience for most people is the boiling of water. Water naturally "evaporates" into the air, but when we boil it, evaporation is more rapid. What exactly is taking place and does this extra water in the air have anything to do with humidity? A third question might be *how are evaporation and humidity related?* Finally, as you watch a pot of water boil you see "steam" rising into the air (some of you may even be familiar with the "steam" rising from geysers and hot springs). This looks vaguely like fog or a cloud. So a final question might be *what exactly is this "steam" and is it the same as a cloud?*

Since each of these smaller pieces builds on the previous one, it is important for you to have a solid understanding of each one before moving on to the next. This is important, and if you feel as though you do not quite understand something, talk to your instructor. You may find it useful to refer back to Figure C-1 on occasion, to see where we are and where we are going. This is particularly useful if you are having difficulty with a particular concept and you find yourself getting frustrated and bogged down on small details.

1 *EXPERIMENTING AND HYPOTHESIZING*

We begin this unit by looking at something we are all familiar with—thermometers. As you know, thermometers are used to determine the temperature of an object. What you may not be aware of is that thermometers do not always give a correct reading for the temperature of an object. Thus, we will begin by learning how to make an accurate measurement of temperature with a thermometer. After that, we will explore how the temperature of a glass of water changes when it is mixed with water of a different temperature. In particular, we will devise a method to determine the final temperature of the mixture after we combine a cup of hot water with a cup of cold water.

You may need some of the following equipment for the activities in this section:

- Thermometers and temperature sensors [1.1, 1.2]
- ~500 ml beakers [1.1, 1.2]
- ~100 ml graduated cylinders [1.1, 1.2]
- Styrofoam cups [1.1, 1.2]
- Hot and cold water [1.1, 1.2]
- MBL system [1.1, 1.2]

1.1 THERMOMETERS AND TEMPERATURES

Why does it take so long to take your temperature with an old-fashioned thermometer? What determines how long you must wait before the thermometer returns an accurate reading? You will be using thermometers in most of the activities in this unit. Therefore it makes sense to spend a few minutes investigating how they work.

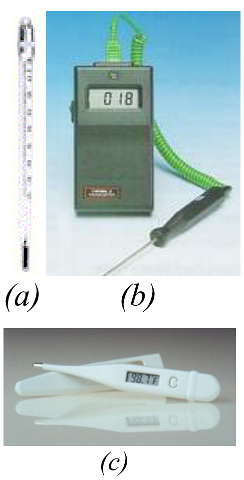


Figure C-2: Two types of thermometers: *(a)* A standard liquid-filled thermometer *(b)* and *(c)* Electronic-thermometers. (*(a)* and *(c)* courtesy of Pocket Nurse Enterprises, Inc. *(b)* courtesy of Electronic Temperature Instruments, Ltd.)

Activity 1.1.1 Thermometers and Thermal Equilibrium

a) Use a “standard” thermometer to measure the temperature of the air in the room. Next, get a cup of hot water from the tap and while watching the temperature readings on the thermometer, place it in the cup of water. Describe what you observe.

b) Is the thermometer making an accurate reading of the temperature of the water from the moment it is placed in the cup? How long do you think the thermometer needs to sit in the water before it is reading an accurate temperature?

c) If the thermometer is not making an accurate reading of the water when it is first placed in the water, what temperature do you think it is measuring?

d) Now take the thermometer out of the water and let it sit in the room air. Again, watch the temperature readings and describe what you observe. How long do you think the thermometer needs to sit in the air before it is reading an accurate temperature? Why do you think this is so different from the time it needs to sit in the water?

There is a subtle lesson to be learned here. Before the thermometer is placed in the water, it is at the same temperature as the air. Once it is placed in the water, its temperature increases because it is in contact with the warmer water. Once it has been sitting in the water for a while, the thermometer reaches the temperature of the water. We say that it is in *thermal equilibrium* with the water. Thus, we view this situation as follows. When two objects at different temperatures are in thermal contact with each other, their temperatures will change. This will continue until the objects have reached thermal equilibrium, at which point their temperatures are equal. As simple and obvious as this may sound, scientists use this type of reasoning to define the concept of temperature.

Using a Temperature Sensor

The previous activity can be made much more visual with the use of an MBL system and a temperature sensor. The computer can make many temperature readings and plot them on a graph. This will allow us to view how the temperature is changing with time. This kind of graph is called a temperature-time graph. A temperature-time graph can be much more useful than a large table of numbers because you can see very quickly how the temperature varies over time. Your instructor may provide you with more specific information on how to use the temperature sensor and MBL system.

Activity 1.1.2 Temperature Sensors

a) Imagine the following experiment. A temperature sensor has been sitting out in room air for a while. The sensor is then placed in a cup of hot water until it is in thermal equilibrium with the water. Then the temperature sensor is removed from the water and held in the air. Make a rough sketch of what you think the temperature-time graph will look like for this experiment. Briefly explain your prediction. **Note:** Make sure to indicate when the sensor is placed into the water and when it is removed from the water.

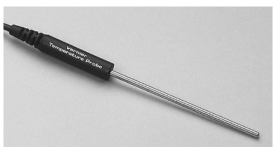


Figure C-3: Temperature sensor for use with computer based laboratory systems. (Courtesy Vernier Software and Technology)

b) Now, carry out the experiment. Set the software to run for 10 minutes and start the experiment after the sensor has been sitting in the air for quite a while. Then, after about 30 seconds, place the sensor in a cup of hot water. After the temperature no longer appears to be changing, take the sensor out of the water and let it sit in the air for the remainder of the experiment. Print out a copy of your temperature-time graph to include in your Activity Guide and label the portions when the sensor is in the water and when it is in the air.

c) Is the behavior of the temperature sensor similar to that of the standard thermometer? Explain briefly

d) Describe how the graph looks when the temperature is changing rapidly compared to when it is changing more slowly.

It should be fairly obvious from the previous activity that thermal equilibrium is reached much more quickly when the temperature sensor is in water compared to when it is in air. It should also be pretty clear that the electronic temperature sensor behaves very much like a standard thermometer.

Accuracy of the Temperature Sensor

The fact that an electronic temperature sensor behaves much like a standard thermometer does not mean that it should be trusted without a second thought. As we saw in Activity 1.1.1, even a standard thermometer will give an inaccurate reading if it is not in thermal equilibrium with the object whose temperature you are trying to measure. Furthermore, there is some inherent imprecision when making *any* measurement (not just temperature). In the next activity, we will look more closely at how precisely we can measure temperature with our temperature sensors.

Activity 1.1.3 What Temperature is it?

a) Take a cup of cool water and place the temperature sensor in the water. After about 30 seconds, start the software while continuing to hold the temperature sensor in the water. Stop the software after 30 seconds and comment on the temperature-time graph.

b) Now, rescale your temperature graph so that the temperature range is about 1°C. For example, if you measured a temperature of 23.9°C, try scaling the graph to go from 23.5°C to 24.5°C°. Then try scaling it to an even smaller ranger say 0.5°C or 0.2°C. Comment on what you observe.

c) Based on your observations in part b), how much can you trust your temperature reading? Is it precise to within 1°C? To within 0.5°C? To within 0.1°C? Explain your reasoning.

As demonstrated in the previous activity, your temperature measurements will not be perfectly accurate. There will be some (we hope small) amount of uncertainty associated with any measurement process and it is wise to always keep this in mind. Knowing how much uncertainty is inherent in your measurements will help you determine when two measurements are the same. For example, if your temperature sensor is precise to within 0.5°C, and you measure the temperature of two cups of water to be T_1=22.3°C and T_2=23.6°C, you can be pretty sure that these temperatures are, in fact, different. If, on the other hand, you measure these temperatures to be, T_1=22.3°C and T_2=22.9°C, you cannot know for sure whether or not these two temperatures are different because they could each be wrong by as much as 0.5°C.

Most electronic temperature sensors are precise to within 0.5°C and many are precise to within 0.1°C (or even better). In this unit, we will not be interested in measuring temperature differences that are smaller than 1°C.

1.2 MIXING WATER

The next two activities constitute a game. The object of the game is to be able to predict the final temperature when two cups of water (initially at different temperatures) are mixed together. As you might guess, this is fairly easy if the two cups have the same amount of water, but it is not so easy when they don't. Before beginning an experiment it is useful to make a prediction about what you expect to see. It is important to make the prediction as specific as possible. For example, in the next activity you will predict the final temperature that results from mixing together a cup of hot water and a cup of cold water. You might predict that the final temperature will be somewhere in between the two starting temperatures, and you'd be correct. But a better prediction will be more specific about the final temperature. Will the final temperature be exactly halfway between the two temperatures? One third of the way?

Developing such a prediction involves considering what factors are likely to affect the experiment (e.g., the amount of water) and what factor are unlikely to affect the experiment (e.g., day of the week). Then the experiment tests your hypothesis. If the results of the experiment are inconsistent with your hypothesis, you know that it is false. On the other hand, if your experimental results are consistent with your hypothesis, your hypothesis is supported. It is important to remember that a hypothesis can never be proven correct. Even though a hypothesis may be supported by many observations and experiments, only one careful experiment is required to show that it is incomplete or incorrect it.

Activity 1.2.1 Final Temperature—Predictions

a) Suppose you mix two cups of water that are at different temperatures, one at a high temperature, T_H, and one at a lower temperature, T_C. What do you think the final temperature of the mixture will be? (i.e., will it be higher than T_H, lower than T_C, right in the middle, or something else?) Explain briefly

b) What factors do you think will influence the final temperature of the mixture? Explain.

c) To make your ideas a bit more concrete, predict the final temperature of a mixture of 50 grams of water[1] at 60ºC mixed with 100 grams of water at 20ºC. Try to explain the reasoning behind your prediction as best you can. **Note:** If you have to make a bit of a guess at this point, that's fine.

Testing Your Predictions

In this next activity you'll have a chance to test the predictions you just made. Remember that the goal is to come up with a method of predicting the final temperature when two *arbitrary* amounts of water at different temperatures are mixed together. So be prepared to make a number of different measurements.

[1] A gram is measurement of *mass*, which is a measure of "how much stuff" you have. One milliliter (ml) of water has a mass of 1 gram (gm)

Activity 1.2.2 Final Temperature—Observations

a) Typically, when beginning a scientific investigation, it is best to start simple. Therefore, you should begin by mixing together two cups of *equal* amounts of water (say, 50 or 100 grams each) that are at different temperatures. It might be difficult to get specific temperatures, so just use hot tap water and cool tap water. To record your observations, place an electronic temperature sensor in each cup and begin recording data. Then, after waiting until the sensors are recording a fairly steady value, pour the cold water into the cup with the hot water and place both temperature sensors in the final mixture. Print out a copy of your data to include in your activity guide. Is the final temperature halfway between the two initial temperatures?

b) You should have noticed that both temperature sensors end up at the same final temperature (within a few tenths of a degree Celsius). If they differ from each other by more than 0.5°C, ask your instructor how to calibrate your sensors. You should also have noticed that the final temperature was not exactly halfway between the two initial temperatures. Most often, the final temperature will be a bit below the halfway point (usually less than 1°C). Explain why this happens. **Hint:** What happens to the temperature of a cup of hat water that sits in a room for a long time?

c) The fact that the final temperature was (probably) not exactly halfway between the two initial temperatures is worth thinking about briefly. There could be some outside influence that affects the experiments (as alluded to above), or there could be other measurement errors that lead to the unexpected result. Besides temperature measurements, what other measurements did you make in this experiment? How might errors in these measurements lead to a different final temperature than you expected?

d) Now try the same experiment again, except, this time, combine 100 grams of cold water and only 50 grams of hot water. Again, write down the initial and final temperature and print out a copy of your graph to include in your activity guide. **Note:** Remember to pour the cold water into the cup of hot water.

e) Using the data from this experiment, work with your group to determine an equation that allows you to calculate the final temperature you measured above. **Hint:** What fraction of the total mixture was initial in the cold water cup and what fraction was initially in the hot water cup? Do these fractions have any bearing on where the final temperature will lie between the two initial temperatures? Remember that your final temperature measurement might be a little lower than expected.

f) When you have an equation that works, generalize it so that it is valid for arbitrary amounts of water and arbitrary initial temperatures. Show your work and write your new equation below.

g) Using your newfound equation calculate the final temperature when mixing 50 grams of water at 60°C and 100 grams of water at 20°C. Do you results agree with your predictions from Activity 1.2.1? (It's okay if they don't)

h) Try one more experiment with 100 grams of cold water and 25 grams of hot water. Compare your experimental measurements with your theoretical predictions. They should be quite close!

Congratulations! You have successfully developed a method for predicting the final temperature when mixing two cups of water together. This kind of quantitative reasoning plays an important part in the scientific process. It is equally important to consider sources of error and outside influences that may have affected your experiment. Each measurement that is made is a potential source of error so that making careful measurements is essential. Furthermore, even if perfect measurements could be made, outside influences (such as the cooling of the hot water) can also have an impact on the final results of an experiment. Often it is only possible to understand a particular result after including the outside influences and sources of error.

Checkpoint Discussion: Before proceeding, discuss your ideas with your instructor.

2 *THE DISTINCTION BETWEEN HEAT AND TEMPERATURE*

While most people are reasonably good at distinguishing hot from cold, our bodies are actually not very trustworthy when it comes to trying to determine temperature. In fact, it is not too difficult to fool yourself into thinking an object is hotter or colder than it really is. This happens, for example, if you place one hand in a tub of hot water and the other hand in a tub of cold water and leave them there for a little while. At this point, most people will be able to correctly distinguish which water is hotter. However, if you now take your hands out and put them both into water of the same temperature, one hand will feel hotter than the other. The hand that was in the hot water will sense the water as being colder than the hand that was in the cold water. If you didn't know better, you would incorrectly predict that the temperatures were *not* the same. This effect is quite dramatic and if you have never experienced it before, you should try it!

One of the things we will be considering in this section is exactly what we mean by how "hot" something is. Since we cannot trust our senses, we would like to develop a method of assigning a value for the "hotness" of something that is independent of our body's senses. Clearly, a thermometer is what we seek. However, although they are quite commonplace, many people do not understand exactly why a thermometer works the way it does. To gain an understanding of the inner workings of a thermometer, you will have an opportunity to build your own.

You may need some of the following equipment for the activities in this section:

- Liquid soap and ground black pepper [2.1]
- Styrofoam (or other insulated) cups [2.1 - 2.4]
- Hot and cold water [2.1- 2.3]
- Brownian motion demonstration (or use video clip) [2.1]
- Thermal expansion of a solid demonstration [2.2]
- Blocks made of metal, Styrofoam and wood with a hole drilled in them to accept the tip of a temperature sensor [2.2]
- Glass syringes and small flasks [2.2]
- Rubber tubing, one-holed stoppers, connectors [2.2]
- "Standard" thermometers [2.2]
- Thermos (preferably vacuum insulated) [2.3]
- An MBL system [2.3, 2.4]
- Temperature sensors [2.3, 2.4]
- Heat pulser [2.3, 2.4]
- Immersion Heater [2.3, 2.4]
- Anti-freeze [2.4]

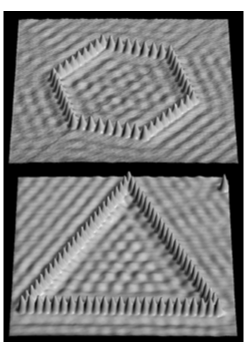


Figure C-4: Individual iron atoms are arranged in different patterns on a copper substrate and imaged with a scanning tunneling microscope. (Courtesy of International Business Machines Corporation)

2.1 GASES, LIQUIDS, AND SOLIDS

Much of our knowledge about gases, liquids, and solids come from delicate experiments that are not easily reproduced in the classroom. One reason for this is the extremely small size of the constituents of matter. For example, the approximate size of an atom is about 0.0000000001 meters (10^{-10} m) in diameter. That means it would take roughly 100 million (10^8) atoms, side by side, to reach a length of one centimeter. Even more dramatic is that it would take about 10^{24} atoms, all clumped together, to occupy a cubic centimeter (a region about the size of a board-game die). That's a million billion billion, an almost incomprehensibly large number. Even the best optical microscopes in the world do not allow us to see things that small with our eyes. It is only with the use of exotic instruments, such as electron microscopes, atomic-force microscopes, or scanning-tunneling microscopes, that we can image individual atoms (see Figure C-4).

Before continuing, it will be useful to give a brief description of atoms and molecules. An atom is a stable configuration of three different tiny particles—electrons, neutrons, and protons. These particles are arranged in such a way as to behave like a single particle. There are many different configurations of electrons, neutrons, and protons that will form stable entities, and each one has its own unique properties. These entities are referred to as chemical *elements* and are listed in the periodic table, which groups them by weight and the similarity of their properties. A *molecule*, on the other hand, is a stable grouping of two or more atoms that behave as a single entity and has its own unique properties that might be very different from the elements that make it up. All of the materials you come into contact with everyday are made of large numbers of atoms or molecules, the properties of which are determined by the individual atoms or molecules.

As a simple example, oxygen is an element that most people are familiar with. However, what most people don't realize is that the oxygen that we breathe is actually *molecular* oxygen, consisting of two oxygen atoms bound together. We denote elemental oxygen with the symbol O and molecular oxygen with the symbol O_2. The subscript 2 refers to the fact that there are two oxygen atoms contained in this molecule. Similarly, three oxygen atoms can be combined into a stable configuration called ozone, with the symbol O_3. Furthermore, even though there are only about 100 elements, some molecules (such as DNA) can contain thousands, or even millions of atoms. While not every combination of atoms will form a stable entity, there seems to be no limit to the number of molecules that can be created. Inventing new materials that have special properties is an active area of research called *materials science*.

We won't worry too much about the distinction between atoms and molecules. For our purposes, what is important is the fact that they are extremely small, stable particles that usually are not deformed or altered. With this in mind, let us now loosely define the three most common phases of matter: solids, liquids, and gases.

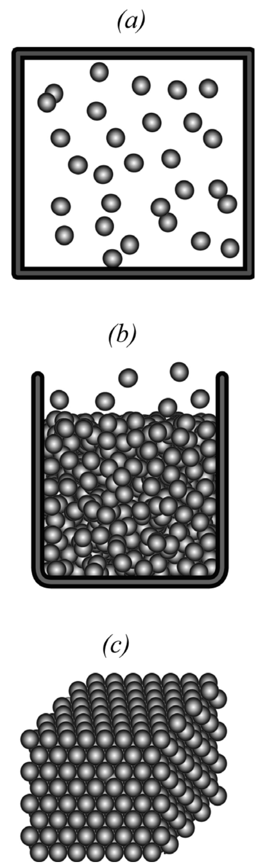


Figure C-5: Depictions of atomic models for (*a*) a gas, (*b*) a liquid, and (*c*) a solid.

Gases—A gas is a substance in which all the particles (atoms or molecules) have (basically) no attraction to each other. We can think of a gas as a collection of little hard spheres, that are whizzing around in random directions at very high speeds (about 1,000 mph). When two of these particles run into each other, they collide and bounce off one another, sort of like a collision between billiard balls. The important point is that there is no interaction between the individual particles unless they smash into each other. Another important point about gases is that because the particles move so fast and collide so often, the effects of gravity are almost completely negligible when discussing the individual gas particles. Thus, when a gas is placed into a container, the gas molecules quickly fly around and completely fill whatever container they are in.

Liquids—A liquid is a substance in which the particles (atoms or molecules) have an attraction for one another. (If the liquid is made up of molecules, such as water, then it is important to note that the attraction between the atoms in the molecules is *much* stronger than the attraction between the molecules.) This means that the individual water molecules want to be next to each other. Because of their attraction to each other, liquid particles behave "en-mass," with billions upon billions of individual particles acting together. Thus, unlike gases, gravity has a large effect on liquids. In addition, because the attraction between liquid particles is not all that strong, individual particles can move between other particles, allowing the liquid to "flow." Thus, when a liquid is placed inside a container the particles fall to the bottom and conform to the shape of the container.

Solids—A solid is a substance in which there are very strong attractions between the individual particles (atoms or molecules). (The strength of this attraction is less than the attraction between atoms in a molecule, but it is a much stronger one than the attraction between liquid particles.) In fact, this attraction is so strong, that the individual particles cannot move around like they can in a liquid. In some sense, you can think of a solid as an extremely large molecule, whose constituents may be either atoms or molecules. Now, although the individual particles are fixed to their neighbors, they can still wiggle around a little. It is as if the particles are attached to each other by little springs, which allow them to bounce around a little without moving away from each other.[2] Because of the strong bonds between neighboring atoms, a solid retains its shape regardless of what kind of container it is placed in (it doesn't flow like a liquid).

Measuring the Size of a Molecule

It might seem as if molecules are too small to be investigated, let alone have any importance in our day to day lives. We will soon see that thinking of substances as composed of atoms and molecules allows us to develop a very powerful theory for heat and temperature. And while it is true that molecules cannot be seen directly, it is possible to design experiments that do reveal information about lone molecules. In fact, using only a drop of soap, some water, and some pepper you can determine the approximate size of a single soap molecule!

Activity 2.1.1 Estimating the Size of a Molecules

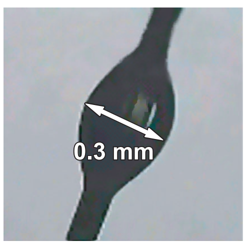


Figure C-6: A human hair with a small drop of liquid soap on it as seen under a microscope.

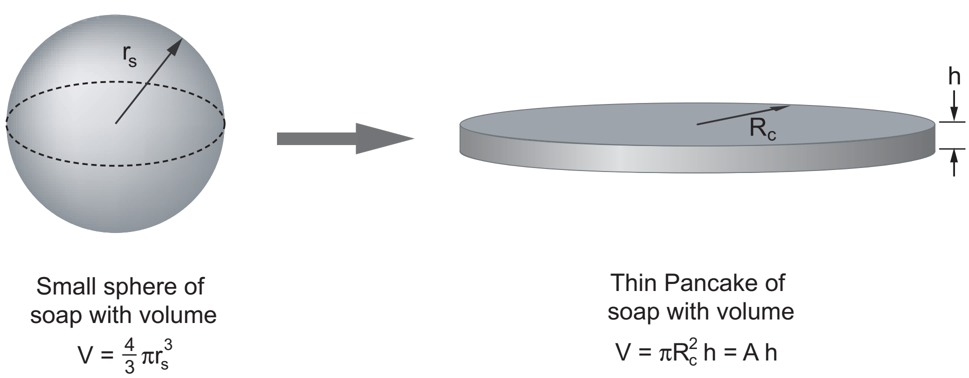
a) Take a small cup of water and gently sprinkle some pepper on the surface. Next, pull a hair out of your head and lightly dip one end in some liquid soap. You should end up with a small bead of soap on the end of the hair. Your bead of soap should look like a tiny sphere with a diameter about twice the width of your hair (see Figure C-6). Now poke this end of the hair into the center of the cup of water and describe what you observe.

b) What do you think happened to the drop of soap?

[2] In fact, the bonds between atoms in a molecule also behave like little springs.

c) You should have seen that an approximately circular area has been "cleared" of pepper. Estimate the radius of this circle and calculate the area in square millimeters. Recall that the area of a circle is $A = \pi R_c^{\ 2}$, where R_c is the radius of the circle. If you measure the radius in millimeters, your area will be in square millimeters.

d) So what happened in this experiment? The small bead of soap that we began with spread out into a thin pancake when it hit the surface of the water. That is what pushed all the pepper out of the way. Basically, the soap tried to make itself as thin as possible. Clearly, the pancake of soap cannot be thinner than one molecular layer (although there's no guarantee that it isn't more than one molecule thick). Thus, in order to estimate the size of a soap molecule, we need to equate the volume of soap that we started with (the small sphere) to the volume of soap that we ended with (the thin pancake). This is shown schematically below. **Note:** It might help to visualize this using modeling clay. Roll some clay into a ball. This is like the soap droplet on the hair. Then squash the ball with your hand until it is a thin pancake. Notice that the shape of the clay has changed dramatically, but the amount (volume) of clay has remained the same.



Estimate the volume of soap that was initially on the hair. (Remember that this is an estimate. You don't have to know the *exact* size of the sphere of soap, just the *approximate* size). **Hint:** The width of a human hair is about 0.1 mm.

e) Now equate this volume to the volume of a small pancake, whose area A, you estimated in part c), and solve for the height h, of the pancake This is your estimate for the smallest thickness that a single soap molecule can have.

Keep in mind that the size of the molecule you just calculated is only an *estimate* for the size of a soap molecule. In fact, since it is possible that the thin layer of soap is more than one molecule thick, all we can say for sure is that the size of a soap molecule can be *no larger* than the value you calculated in the last activity. However, this estimate should make it clear that the size of a soap molecule is many thousands of times smaller than the width of a human hair. Furthermore, since soap molecules are composed of many atoms, the size of an atom is smaller still.

Seeing Molecular Motion

Because the small size of atoms and molecules prohibits us from seeing them individually, we can't just look at them to understand how they behave. However, because molecular motion is important in understanding much of what is to come, we will attempt to give you some direct evidence for the motion of molecules. The following activity re-creates an observation that was first made in 1827 by the English botanist Robert Brown. Brown was looking through a microscope at pollen grains and other small inanimate objects suspended in water and was surprised to see them moving. This is now referred to as *Brownian motion*.

Activity 2.1.2 Brownian Motion

a) If your instructor has a Brownian motion demonstration, be sure to take a look. Keep in mind that the particles you are seeing are thousands of times *larger* than the water molecules themselves. It is also important to notice that the drop of water you are looking at appears to be completely motionless, like a glass of water that has been sitting for a long time. Describe the motion you observe as best you can.

Figure C-7: What do you see when you look at particles suspended in water under a microscope? Rather than moving smoothly, the particles jiggle around erratically. What do you think causes the particles to move in this way?

b) Does the motion of these particles tell you anything about the motion of the water molecules? Explain why or why not.

It is important to note that the water molecules are much smaller than the particles you are observing. What you are seeing when the particle moves is the effect of thousands of water molecules all hitting the big particle. An analogy would be many people together pushing a very large "Earth ball" or medicine ball. One person alone cannot move the ball. However, if many people push at various places on the ball then the ball might move in a certain direction. That direction will change, however, as the people move around in order to find room to push.

2.2 TURNING UP THE HEAT

Now that we have some evidence for the natural motion of molecules in a liquid, we can try to observe some differences between hot and cold objects. This is not as simple as it sounds. For example, imagine walking into a room where there are two glasses of water that look identical but they are at different temperatures. Without measuring their temperatures or feeling them in any way, would you be able to determine which one is hotter? Thus, we need to answer the question, "Is anything measurably different between these two glasses of water besides their temperature?" This is the subject of the following activity.

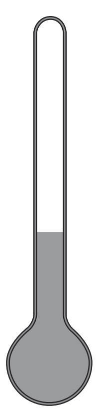
Activity 2.2.1 Heating Objects

a) Besides temperature, do you think there will be any measurable change in a solid if it is heated? If so, explain briefly what you think will happen.

b) Your instructor may have a demonstration of a solid object being heated. Watch carefully and describe what happens to the object when it is heated.

c) Explain how this behavior might be exploited to make a solid thermometer.

d) Explain how this same behavior in a liquid could be used to construct a liquid thermometer.



e) At the beginning of this activity, you saw that when a solid is heated, it expands. In terms of the actual molecules that make up a solid (or liquid), can you think of two possible explanations that would account this observation?

One very reasonable hypothesis that explains the expansion phenomenon (called *thermal expansion*) observed in the previous activity is that the molecules themselves increase in size when heated. However, there are some materials that actually contract when heated. Even more confusing is that certain materials expand when heated at some temperatures and contract when heated at other temperatures. For example, liquid water contracts when heated between 0°C and 4°C, but expands when heated above 4°C. Thus, although the idea that the molecules get larger when heated is appealing, it is difficult to use this idea to explain much of what is actually observed.

Molecular Motion and Temperature

An alternative idea that seems quite reasonable to most people is that the molecules gain energy and mover faster when heated. As the molecules jiggle around more and more, it seems plausible that they might "demand" a little more space. In fact, it is more complicated than this because the intermolecular forces can be such that the molecules actually "demand" *less* space when moving faster. While the details of this are beyond the scope of this class, numerous experiments have confirmed that when a material is heated, the sizes of the molecules do not change, whereas the average speed of the molecules increases (regardless of whether the material expands or contracts).[3] Therefore, in this class, we will consider *temperature* to be a measure of the underlying molecular motion of the material.[4]

Checkpoint Discussion: Before proceeding, discuss your ideas with your instructor.

The Need for a Thermometer

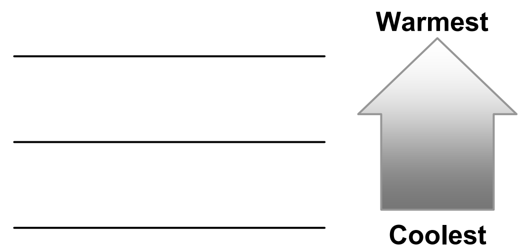
At the beginning of this section, it was mentioned that our bodies are not necessarily very good at determining the temperature of objects. The following activity explores this point a little further.

Activity 2.2.2 Which is Hotter?

a) You should have at your table a block of Styrofoam, a block of wood, and a block of aluminum. Briefly touch each of the objects

[3] In fact, in some situations, it is possible for molecules to gain energy without increasing their average speed.
[4] A more detailed study of temperature would reveal that this definition is somewhat limited. However, for our purposes it is more than sufficient.

and rank them in order from coolest to warmest. **Note:** Don't handle the blocks too long or you will warm them up with your hands.



b) Do you think these objects have the same or different temperature? If you think they are the same, explain why. If you think they are different, indicate which has the highest temperature and which has the lowest temperature and estimate the temperature difference between them in °C.

c) Now, measure the temperatures of each object by inserting a temperature sensor into the hole in each object. Note: Remember not to handle the objects too much because the heat from your hands will change their temperature.

Block Material	Temperature (°C)
Metal	
Styrofoam	
Wood	

d) Do your results surprise you? Explain briefly.

The last activity should make it very clear that what we feel as hot and cold does not necessarily coincide with the temperature of the object. That is, our bodies are not good thermometers, So how does one go about building a good thermometer? Well, we saw in Activity 2.2.1, that some solids expand when heated. You proposed a method for building a thermometer based upon this property common to many solids and liquids. Unfortunately, since solids and liquids undergo only a very small expansion when heated, it is not very easy to construct this type of thermometer in the classroom. Instead, we will build a constant pressure *gas* thermometer. But first, we need to consider the behavior of gases when heated.

Do Gases Expand?

Because it is the intermolecular forces that determine whether a material will expand or contract when heated, you might wonder what will happen to a gas when it is heated. Since we have described a gas as having no intermolecular forces, it may not be clear what should happen with a gas. The following activity examines this question.

Activity 2.2.3 Heating a Gas

a) Recall that a gas expands to fill the container it is placed in. Image taking a fixed amount of gas and placing it into a large glass container with a lid that sealed the container and was immovable. Now, imaging heating up the gas, assuming that the size of the container cannot change. Would the gas expand? Explain briefly.

b) Now if the temperature of the gas has increased due to the heating, will this have any effect on the collisions the gas molecules make with the walls? Explain briefly. **Hint:** What is temperature a measure of?

It should be clear that unlike a liquid or a solid, a gas does not have a well defined volume. Instead, the volume of a gas depends entirely on the size of the container it is placed in. Thus, a gas will expand or contract only if the volume of the container changes, independent of whether it has been heated or not. Now, if the container is flexible and has the ability to shrink or grow, then heating the gas will result in the gas pushing harder on the container which will cause the volume of the container to increase. This is the principle of the constant pressure gas thermometer.

Activity 2.2.4 Building a Thermometer

a) To build the thermometer, place a one-holed rubber stopper in the mouth of a small flask or other container. Next, attach a piece of rubber tubing to the hole in the rubber stopper with connector or adapter. For the last step, you will need to choose a glass syringe that is approximately 20-25% of the size of the container that you are using.[5] Thus, if you are using a 125-ml flask, a 20-30 ml syringe works well. Finally, connect the free end of the rubber tubing to the glass syringe, making sure that the piston on the syringe is only pulled out approximately 25% of the way (if you are using a 20-ml syringe, the barrel should be pulled out to about 5-ml). A schematic diagram of a finished thermometer is shown in Figure C-8 below. Check to see that your thermometer works by placing the flask (not the syringe) in hot water and then in cold water. Make sure that the piston on the syringe is always pointed in the "up" direction.

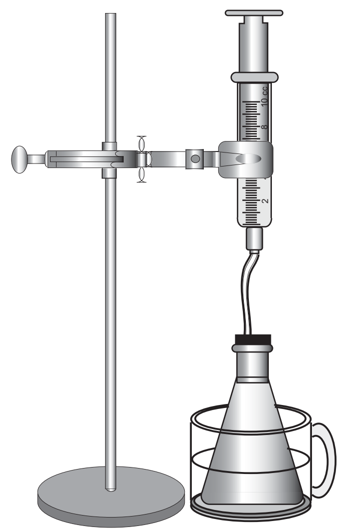


Figure C-8: A thermometer built from a flask and glass syringe sits in a cup of water.

[5] Using a syringe that is too small will result in the piston of the syringe being pushed all the way out of the barrel, which often leads to broken glass.

b) Now, the numbers on the side of your syringe correspond to actual temperatures, but you first need to calibrate your thermometer to a universally accepted temperature scale. Do this by placing your thermometer into a hot water bath of known temperature and labeling the position of the piston with this temperature using masking tape. Then do the same thing for a cold water bath of known temperature. Using these two markings, you should be able to determine what temperatures correspond to the other major markings on your syringe. Record these in the table below.

Syringe labels (ml)	**Temperature (ºC)**

c) Determine the maximum and minimum temperatures you can measure with your thermometer. Would your thermometer be useful for measuring very high or very low temperatures? Explain.

Although the thermometer you just built works reasonably well, it has a couple of major drawbacks. One is that this thermometer can only measure a limited range of temperatures. More problematic is that a change in atmospheric pressure will change the reading on the thermometer, even if the temperature does not change. This is one reason why people do not use such crude thermometers. However, this type of thermometer demonstrates something fundamental about all thermometers. They take advantage of a property of a material that changes with temperature. In this case, the gas causes the container to expand as its temperature is increased, and that expansion can be measured and calibrated. Although the expansion property of materials is a convenient feature to exploit in making a thermometer, it is by no means the only one. Other types of thermometers may rely on other properties of substances. The temperature sensors we have been using illustrate one common example. These sensors exploit the electrical properties of materials to produce an electrical current that changes with temperature. This current is then calibrated and displayed as a temperature on the computer screen.

2.3 THERMAL ENERGY TRANSFER AND TEMPERATURE CHANGES

Armed with our understanding of the concept of temperature, we will now turn our attention to trying to understand "heat." Clearly, we can increase the temperature of an object by heating it up, but what exactly does "heating it up," mean? One way of heating something up is to place it on a hot burner. But are there other ways? That is the topic of the next activity.

Heating Water with Mechanical Work

If temperature is a measure of how quickly molecules are moving, then perhaps we can increase an object's temperature by physically making the molecules move faster. In fact, pioneering experiments in heat and temperature tried to do just this. Scientists would spin paddle wheels in liquids and try to measure a change in temperature. You will do this in the next activity, albeit in a less controlled manner.

Activity 2.3.1 Shake it, Baby!

Figure C-9: A thermos is an insulating container designed to prevent heat from entering or leaving. (Corbis Images)

a) Your instructor will place a small amount of water (50-100 g) in a thermos, which is an insulated container designed to prevent heat from entering or leaving the container. Write down the initial temperature of the water that is placed in the thermos.

b) After the thermos is sealed, your instructor will begin shaking it vigorously. This is difficult work, so after about a minute or so, it will be passed to one of the students in class. This student should continue to shake the thermos vigorously for about a minute or so, then pass it on to someone else. When the thermos has gone around the whole room, pass it back to the instructor, who will open it and measure the temperature of the water inside. (Since this will take some time, your instructor might want you to work ahead on the next activity while all the shaking is going on.) Write down the final temperature of the water after having been shaken.

c) From a molecular perspective, what is different about the water after having been shaken and what caused the change? (You should think a bit more deeply than just replying, "the shaking caused the change.")

Heating Water with Electric Work

This following activity involves heating up some water while monitoring its temperature in a way that might be a bit more familiar. You will be using an immersion heater, which is similar to a stove's burner. When on, electricity causes the immersion heater to get hot and become a "source of heat." We use the immersion heater instead of a stove because it is easier to control. It is also a more delicate piece of equipment and you need to be careful with it.

Caution: ***Never plug the heat pulser into the wall when it is not in the water.*** This will damage the heating unit. It is all right to plug the heat pulser into the computer, but make sure it is immersed in water before actually pulsing heat. Another thing to remember is that when the heater is on, ***the metal portion of the heating unit shouldn't touch anything except the water***. This includes the air in the room, you, and your partner!

Activity 2.3.2 Pulsing Heat in a Cup of Water

a) Begin by putting 100 grams of water into a Styrofoam cup and placing a heat pulser in the water with a stand. Set the total experiment time to 250 seconds and set the heat pulse length to 10 seconds. When you think you are ready, have your instructor check your set-up ***before*** starting the experiment! You will need to stir the water with the temperature sensor throughout this experiment, being careful not to touch the heat pulser with the temperature sensor. Begin the experiment and after about 40 seconds (don't forget to stir the water continually), click on the heat button. This will add a 10-second pulse of heat, which should increase the temperature of the water a little bit. After the temperature of the water has stabilized (which should take about 40 seconds) hit the heat button again. Repeat this process until the experiment ends. When you are done, you should have hit the heat button five times, once at 40, 80, 120, 160, and 200 seconds. Briefly describe the graph you obtained, and print out a copy for your activity guide.

b) Fill in the following table by finding the temperature change for each pulse of heat that was added.

Heat Pulse	Temp. before pulse (ºC)	Temp. after pulse (ºC)	Temp. Change (ºC)
#1			
#2			
#3			
#4			
#5			

c) Explain what a 10-second pulse of heat does to your 100 grams of water. **Hint:** Is the temperature change for each pulse approximately the same?

d) How much would you expect the temperature of 50 grams of water to rise when subjected to one 10-second heat pulse? What about 10 grams of water?

e) From a molecular perspective, what is different about the water before and after receiving a pulse of "heat?"

Heat as an Energy Transfer Process

As these last two activities showed, you can increase the temperature of water by placing a "heat source" in the water (the heat pulser), but you can also increase the temperature of water just by shaking it (with no heat source at all). This surprises many people. Remember, an increase in temperature just means that the molecules are moving more vigorously, so anyway that you can get the molecules moving more vigorously will result in an increase in temperature. By shaking the water, you are adding energy to the water which speeds up the motion of the molecules. In scientific terms, *work* is being done on the water.

Notice that whether we used an electrical source to "heat up" the water or we did work to "heat up" the water, the underlying result is the same in both cases; there is an increase in the motion of the water molecules. Viewed from this perspective, it may be clearer to think in terms of the "energy transferred" to the water. This *thermal energy* either comes from collisions with the molecules that make up the heating unit or from the energy that you used to shake up the thermos.

Quantifying Heat

Now, since the result is the same in both cases (i.e., the temperature of the water increased), we might choose to define "heat" (or *thermal energy*) in terms of the temperature change of water. In fact, this is precisely how scientists first defined heat: One *calorie* is the amount of thermal energy needed to raise the temperature of *one gram* of water by *one degree Celsius*. **Note:** The term *calorie* (with a small c) as defined here is not the same as the term *Calorie* (with a capital C) in regards to food consumption. One "food" *Calorie* is actually equal to one thousand *calories* as defined here. In this activity guide, unless specifically stated, we will always be referring to "non-food" *calories*. The following activity should help you get familiar with this unit of thermal energy.

Activity 2.3.3 Counting Calories

a) If it takes one calorie to raise the temperature of one gram of water by one degree Celsius, how many calories does it take to raise the temperature of 100 grams of water by one degree Celsius?

b) How many calories would it take to raise the temperature of 100 grams of water by 5 degrees Celsius?

c) How much thermal energy (in calories) is required to change 50 grams of water from 30 degrees Celsius up to 50 degrees Celsius?

d) How much thermal energy did each pulse of your heat pulser produce in the Activity 2.3.2? Show the details of your calculation.

2.4 SPECIFIC HEAT

Because we have defined a calorie in terms of its effect on one gram of water, it is particularly easy to determine how much thermal energy has been transferred to a specific amount of water if we know its temperature change. However, there are many situations when one might be interested in heating up a substance besides water. So the question naturally arises, if we transfer one calorie of thermal energy to one gram of a substance that is not water, will its temperature also increase by one degree Celsius? That is the topic of the following activity.

Activity 2.4.1 Activity: Pulsing Heat Again

a) Begin by placing 100 grams of anti-freeze into a Styrofoam cup and placing an immersion heater in it. Set the total experiment time to 250 seconds and prepare to deliver 10-second heat pulses to the water. **Note:** When you think you are ready, have your instructor check your set-up ***before*** starting the experiment! You will need to continually stir the anti-freeze with the temperature sensor throughout this experiment, being careful not to touch the heat pulser with the temperature sensor. Begin the experiment and after about 40 seconds, click on the heat button. After the temperature of the anti-freeze has stabilized (which should take about 40 seconds) hit the heat button again. Repeat this process until the experiment ends. When you are done, you will have clicked on the heat button five times, beginning at 40, 80, 120, 160, and 200 seconds. Briefly describe the graph you obtained, and print out a copy for your activity guide. Is there any difference between this graph and the graph you made in Activity 2.3.2?

b) Fill in the following table by finding the temperature change for each pulse of heat that was added.

Heat Pulse	Temp. before pulse (ºC)	Temp. after pulse (ºC)	Temp. Change (ºC)
#1			
#2			
#3			
#4			
#5			

c) Knowing how much thermal energy is transferred to the anti-freeze in one 10-second pulse (see Activity 2.3.2), determine how much thermal energy it takes to increase the temperature of one gram of anti-freeze by one degree Celsius. Show your work.

The quantity just calculated—how much thermal energy it takes to change the temperature of one gram of anti-freeze by one degree Celsius—is an important and useful quantity. It is a property of the material that does not change as you add more material. Scientists call this quantity the *specific heat* of a substance and denote it with the symbol c. We already know that it takes one calorie to raise the temperature of one gram of water by one degree Celsius. This means that water has a specific heat of $c = 1 \text{ cal} \big/ \text{g}^\circ\text{C}$. Since the specific heat tells us how much energy it takes to increase the temperature of *one* gram of a substance by one degree Celsius, the quantity mc (where m is the mass in grams) is the amount of energy it takes to increase the temperature of m grams by one degree Celsius.

Stated another way, $mc\Delta T$ is the amount of thermal energy gained by an object of mass m when its temperature changes from T_i to T_f. If $T_f > T_i$, then $\Delta T > 0$ and the energy *gained* by the object is positive. But if $T_f < T_i$, then $\Delta T < 0$ and the energy gained by the object is negative. This indicates that the object actually *lost* energy. Now, when two objects are placed in thermal contact with each other, the energy *gained* by one object will be exactly equal to the energy *lost* by the other object. We can use this fact to determine the final temperature of two cups of water when they are mixed together.

Activity 2.4.2 Thermal Energy Transfer

a) Suppose we have m_h grams of hot water, initially at a temperature of T_h and m_c grams of cold water initially at temperature T_c. After mixing, the final temperature is given by T_f. Write down an expression in terms of m_c, c, T_c and T_f for the heat *gained* by the cold water during the mixing process?

b) Write down an expression in terms of m_h, c, T_h and T_f for the heat *lost* by the warmer water during the mixing process. **Hint:** The quantity will be positive.

c) Now, equate the heat *gained* by the cooler water to the heat *lost* by the warmer water and solve for the final temperature T_f.

d) Compare your equation to the equation you deduced in Activity 1.2.2 and show that these two equations are identical.

Checkpoint Discussion: Before proceeding, discuss your ideas with your instructor.

The equation you derived here should be the same as the equation you deduced in Activity 1.2.2 from your experimental observations. Hopefully, you can appreciate this result even more now, knowing that it comes from equating the energy *lost* by the hot water to the energy *gained* by the cold water. In fact, this is actually one of the most important principles in all of the sciences. It is better known as the principle of energy conservation. Although we have not discussed it in detail, energy comes in many different forms, motion, thermal (heat), sound, chemical, solar, electrical, nuclear, etc. The conservation of energy principle states that energy can never be created or destroyed, it can only be changed from one form to another, or transferred from one object to another. We will not try to confirm this statement, but it is worth mentioning that it has been well verified experimentally and is one of the most fundamental ideas in the all the sciences.

3 *EVAPORATION, CONDENSATION, AND HUMIDITY*

At this point, you should have a pretty good understanding of temperature and thermal energy transfer. In this section, we will be taking a look at some situations where a transfer of energy does *not* result in a temperature change. This will lead us to consider how substances can change from one phase into another. In particular, water is a substance that readily changes from a liquid into a gas. In this section we will investigate how this phenomenon can be understood in the context of energy transfer. Specifically we will try to answer the question: is the transfer of energy necessary for a substance to change phase?

You may need some of the following equipment for the activities in this section:

- Styrofoam (or other insulated) cups [3.1 – 3.2]
- Temperature sensors [3.1 – 3.3]
- Immersion heater [3.1]
- Heat pulsers (optional) [3.1]
- MBL system [3.1 – 3.3]
- Mass balance [3.1]
- Hot and cold water [3.1 – 3.3]
- Paper towels [3.2]
- Relative humidity sensors [3.3]
- Plastic, seal-able containers [3.3]

3.1 HEAT OF TRANSFORMATION

We saw in the last section that adding one calorie of thermal energy resulted in one gram of water changing temperature by one degree. Thermal energy is almost always discussed in the context of *transfer* from one object to another. Based on this observation, we might conclude that if we continue to add energy to water its temperature will continue to rise. Let's find out if this is true. We know that if we add thermal energy to a container of water it will eventually begin to boil. The question we want to ask is what happens to the temperature of the water if we continue to add energy to it while is boiling? **Caution:** Be very careful in the following activity. Boiling water can cause severe burns.

Activity 3.1.1 Watching Water Boil

Figure C-10: When water is heated to 100 °C, it begins to boil. What happens to the water when it boils? (Eye Ubiquitous /Corbis Images)

a) Begin by measuring the mass of an empty cup. Now put about 150 grams (it doesn't have to be exact) of room temperature water in the cup and measure the mass of the cup/water system. Report your results below.

b) Set up the temperature sensing software so that it will take readings once per second for about 15 minutes and make sure the temperature range goes from 0°C to 110°C. Next, place the temperature sensor and immersion heater (without plugging it in) into the water. Have your instructor check out your setup ***before*** starting the software. Once you have been checked out, start the software and begin stirring the water with the temperature sensor, being careful not to touch it to the heating unit. After 20 seconds or so, plug in the heating unit and continue stirring (without splashing) until the water begins to boil vigorously. At this point, you can stop stirring the water, but you still need to make sure that the temperature sensor doesn't come into contact with the heating unit.

After the water has been boiling vigorously for at least five minutes, unplug the heating unit and remove the temperature sensor from the water. Then, as quickly as possible (but being as careful as possible), remove the heating unit from the cup and measure the mass of the cup/water system. Report your result below. *Do not throw out the water remaining in the cup!* If there is less water in your cup than before you boiled it, where do you think the water went?

c) After about 10 minutes (when the water has cooled off substantially) measure the mass of the cup/water system again. In the meantime, you should print out a copy of your data for your activity guide. How much water was lost while cooling down? Do you think any water was lost as the water was heating up (but not yet boiling)? Using this information, determine how much liquid water was lost during the boiling process alone. Show all you work.

d) Briefly describe the main features of your graph, specifically what happened to the temperature of the water once it started boiling. What does this mean about how quickly the water molecules are moving? Are they moving faster and faster as you add energy to the boiling liquid? Explain briefly.

e) Although you continued to add energy into the water, its temperature stopped increasing. Where did the energy that you added to the water go?

At this point, you might have noticed that our definition of "heat" is a bit flawed. We defined our unit of thermal energy (the calorie) as the amount of energy needed to increase the temperature of one gram of water by one degree Celsius. Thermal energy was studied by its affect on the temperature of a substance. As the previous activity just showed, the temperature of water did not increase beyond 100°C, even though we were still adding energy. Thus, adding one calorie of thermal energy to one gram of water that is already at 100°C will *not* increase its temperature by one degree, which makes it important to specify what temperature the water is at when we try to increase its temperature. The accepted definition for one calorie is the amount of thermal energy needed to increase the temperature of one gram of water from 14.5 ºC to 15.5 ºC.

The fact that the temperature of the water stopped rising at 100 degrees Celsius may not have been a real surprise to you. However, you may not have realized that this temperature actually tells you something about the interaction between the molecules of water. Since temperature is a measure of the average speed of the molecules, it seems like there will be a point when the water molecules are moving so vigorously that they can "break away" from each other. This is what's referred to as the *boiling point* of the material. At this point, the liquid molecules are breaking their bonds with each other and becoming gas molecules. The bubbles you see when boiling water are bubbles of *gaseous* water (not air), also called steam.[6]

This is an example of what scientists call a *phase change*; the state of matter is being changed from one phase into another. In this case, liquid water is being turned into gaseous water. This happens at a well-defined temperature called the *boiling point*.

[6] It is worth pointing out that when people say they see "steam" rising from a pot of boiling water, they are not actually seeing steam. We will discuss this in more detail when we talk about cloud formation.

Another example is when solid water (ice) is turned into liquid water. This happens at a well defined temperature called the *melting* (or freezing) point. Because the inter-molecular forces are different for different molecules, each substance has its own unique melting and boiling points.

Determining the Energy of a Phase Change

The fact that it takes some energy to "break apart" a solid into a liquid or a liquid into a gas should not be completely surprising. This energy is called a *latent heat,* although latent energy is probably a better term. The latent heat of substances is very important in many areas of science and engineering. With our data from the last activity, we can calculate the *latent heat of vaporization*, which tells you how much energy must be transferred to one gram of water to break the molecules apart and create steam. In the next activity you will determine the latent heat of vaporization of water.

Activity 3.1.2 Heat of Vaporization

a) Examine your graph of the boiling water again. Notice that once the temperature reaches about 60 ºC, the line starts to curve down a little bit (can you explain why?) Before that, it looks like a very straight line. Fit a straight line (either by hand or with the computer) to the portion of the graph that looks like a straight line (up to about 50 ºC). Write down the value for the slope of this line (including units) and interpret what this slope means physically. **Hint:** Remember that the slope is the "rise" over "run." What does the "rise" (the change in *y*) tell you *physically*? What does the "run" (the change in *x*) tell you *physically*?

b) Knowing how much water you had at the beginning of the experiment, use the slope of this line to determine how much energy is being transferred to the water every second (or minute). **Hint:** How many calories does it take to increase the temperature of your cup of water by 1ºC and how many degrees Celsius is the temperature increasing every second (or minute)? Show your calculations!

c) Now that you know how much energy is being delivered to the water every second (or minute), determine how much total energy was delivered to the water during the time it was boiling. **Note:** You will need to determine from your graph exactly how long the water was boiling.

d) Finally, knowing how much energy was delivered during the boiling process and how much liquid water was turned into steam during the boiling process, you should be able to determine how much energy it takes to turn *one gram* of liquid water into steam. **Hint:** Remember that some water might have been turned into steam before or after the water was actually boiling.

e) This quantity is called the *latent heat of vaporization* of water, and is given the symbol L_v. The accepted value for water is 539 cal/g. Compare your result to this and comment on any difference?

Your measurements for the heat of vaporization of water should be reasonably good. There may be some small experimental errors, but if you are careful, you should get a result that is within 5% of the accepted value. It should be pointed out that there is an analogous quantity called the *latent heat of fusion*, which tells you how much energy it takes to *melt* one gram of a substance. Although we won't actually measure this quantity (it would make a nice project), the procedure would be similar to the experiment done here. As already mentioned, the latent heat of vaporization, L_v, tells you how much energy it takes to vaporize one gram of water. Thus the quantity mL_v tells you how much energy it will take to vaporize m grams of water. For example, if you have 10 grams of water, (10 g)(539 cal/g) = 5390 calories will be required to vaporize the water.

3.2 EVAPORATION

Most people are familiar with the fact that a small puddle of water on the ground will eventually disappear. This happens because of a process called evaporation. To begin our investigation of evaporation, we will first examine a quantity called the “wet-bulb” temperature.

Activity 3.2.1 What is Room Temperature?

a) Suppose two electronic temperature sensors are sitting in a room (22°C). One is initially surrounded by a piece of paper towel that has been soaked in warm (30°C) water, the other is open to the air. Sketch below your prediction for the temperature of the two thermometers over ten or so minutes.

b) Now try the experiment. Set up the software to run for at least ten minutes. Then, get a cup of water that is warmer than room temperature (at least 5ºC warmer) and wrap the end of one of the temperature sensors with a small piece of paper towel and tape it securely in place. Place this sensor in the water and stir it around till the temperature stabilizes. When the temperature is stable and only about 5 degrees higher than the thermometer that’s not in the water, start the program and take the thermometer out of the water and lay it down on the table (on top of some paper towels so the table doesn’t get wet). When the program has finished running, print out a copy of your graph. Describe what happens. What is the lowest temperature reached by either thermometer?

c) Which of the thermometers do you think gives a better indication of “room temperature?” Why.

d) Based on your observation, what would happen if a thermometer wrapped in a towel soaked in 22°C water were left open to 22°C air. Do you think the temperature would drop below 27°C? Discuss this with your group and come to a consensus. **Hint:** was the paper towel in the last experiment ever at 22°C?

e) Using your knowledge of temperature, what can you conclude about the motion of the molecules in the wet paper towel versus the room air?

"Wet" and "Dry" Temperatures

The temperatures you just took are referred to as "dry-bulb" and "wet-bulb" temperatures. You will have noticed that the wet-bulb temperature reading drops down below the dry-bulb temperature reading and then levels off to a constant value. It is this constant value that it called the wet-bulb temperature. Most students find it surprising that the wet-bulb temperature drops below the dry bulb. Why does this happen and what stops it from continuing to drop? The answer must have something to do with the wetness of the paper towel, since that is the only difference between the two thermometers. The following thought experiment will help you understand why this happens.

Activity 3.2.2 Evaporating Water

a) Consider a glass of water that has no lid on it. If you wait for a really long time, the water in the glass will have *evaporated*. Where must the water molecules be going.

b) Since boiling water also disappears (although at a much faster rate), do you think these processes (boiling and evaporating) are related? Explain briefly.

c) What causes specific water molecules to leave the water? Can *any* water molecule leave or only certain ones? Explain. **Hint:** When water is boiling, lots of molecules are leaving the water.

d) Remember that temperature is actually a measure of molecular motion. If the wet-bulb reading is lower than the dry-bulb reading, what does that tell you about the average motion of the molecules? Use this to explain how evaporation could be the cause of the wet-bulb reading being lower than the dry-bulb reading.

Checkpoint Discussion: Before proceeding, discuss your ideas with your instructor.

We have developed the idea that temperature is a measure of the average speed of the molecules. This means that the lower wet-bulb reading must be a result of slower moving molecules. However, the wet-bulb reading was originally at a higher temperature, which is a result of faster moving molecules. Clearly the average speed of the molecules around the wet-bulb thermometer is decreasing. Now, one way for the faster moving (water) molecules to slow down, is though collisions with the slower moving (air) molecules. The wet paper towel and the air are in *thermal contact* with each other, so at some point they should reach *thermal equilibrium* (i.e., they would reach the same temperature). However, this is not what we observed. Something *else* must be happening to cool the wet-bulb reading below the dry-bulb reading. This "something" is called evaporative cooling and is an example of a dynamic equilibrium (as opposed to thermal equilibrium).

Evaporative Cooling

Evaporation occurs when molecules leave the bulk liquid and move off into the air. As we saw when boiling water, the ability of molecules to leave the liquid is enhanced when the liquid temperature is greater. This should make sense, since faster moving molecules (and temperature is a measure of molecular speed) are more likely to have the necessary energy to break away from surrounding molecules. But the temperature of a liquid is only the measure of the *average* speed of the molecules. This means that some molecules are moving *faster* than average, while others are moving more slowly. Which are more likely to break free from the liquid? In fact, only the fastest moving molecules are in a position to break away. You observed this earlier when boiling the liquid. If all of the molecules could break free, then the entire cup of water would suddenly vaporize into steam. This in fact, is not what happened. It took some time for the water to boil away. So what happens as the faster moving molecules leave? The *average* speed goes down. Since temperature is a measure of this *average* speed, the temperature goes down as well. This is called *evaporative cooling*, since the evaporation process itself lowers the temperature.

Since heating the water hastens evaporation, one might wonder if qualities of the air (other than temperature) also affect evaporation. The answer, as you will soon discover, is yes. A clue is found in how we cool ourselves during exercise. Sweating uses the idea of evaporative cooling to lower your body's temperature. If you are from New Orleans or another humid city, you might know from first-hand experience that this doesn't always work. Sometimes sweating only makes you wetter. In particular, on a very humid day, sweating is a very inefficient means of cooling off. We'll see why in the next section.

3.3 RELATIVE HUMIDITY AND DYNAMIC EQUILIBRIUM

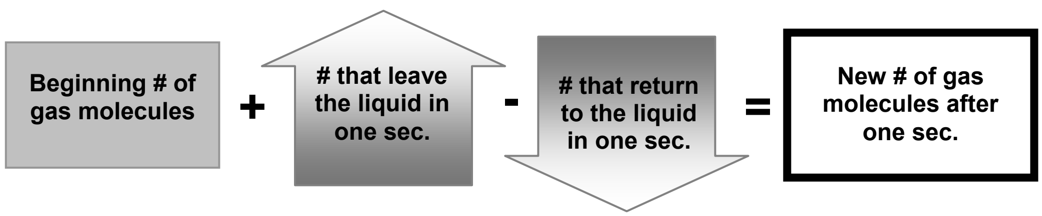
It is well known that placing a lid on a glass of water will prevent it from evaporating. The act of covering the water prevents molecules from leaving the glass. This seems reasonable enough, however, the cover is not in contact with the surface of the water so how can it possibly prevent molecules from leaving the water? Shouldn't the water continue to evaporate? This question is the addressed in the following activity.

Activity 3.3.1 Dynamic Equilibrium

a) If the cover on the glass of water cannot prevent molecules from leaving the water, what do you think happens to the molecules that *do* leave the water? Where do they go?

b) We observe that in a sealed container the water level does not change. This means that the total number of molecules in the liquid is not changing. However, some water molecules are leaving the liquid (the fastest moving ones). How can you reconcile these two seemingly contradictory facts? **Hint:** What must be happing in order for the number of molecules in the liquid to remain constant?

c) If the total number of molecules in the liquid is not changing, yet some molecules are *leaving* the liquid, then it must be the case that some of the gaseous water molecules are *returning* to the liquid. In fact, if the total number of molecules in the liquid is not changing, then the number that leave must be equal to the number that return. If more molecules leave the liquid than return, then the water will slowly disappear. If we know how many molecules are leaving and returning to the liquid each second, then we can determine how the number of gaseous water molecules changes as shown schematically in the figure below.



Now, imagine a glass of water that is sealed with a lid. Imagine further that at the instant the lid is placed on the glass, there are no water molecules in the gaseous state (they are all still in the liquid). After a short time, some of the liquid molecules have evaporated and become gaseous water molecules above the liquid. Since the surface of the liquid water does not change its size, it is reasonable to assume that the same number of water molecules leave the liquid every second. Let's assume for the moment that 100 molecules leave the liquid every second.[7] We need to devise a model for how the molecules might return to the liquid. We could assume that this process is similar to evaporation. That is, every second a certain number of molecules return to the liquid, regardless of other factors. On the other hand, we could assume that the number of molecules that return to the liquid depends on the number of gaseous water molecules above the liquid. Discuss with your group which of these two ideas seems more realistic. Explain your reasoning below.

[7] In fact, the number of molecules evaporating each second depends critically on how much of the water is exposed to the air. Still, a more realistic estimate of this number would be one hundred million billion or 10^{17} molecules every second.

d) Most students find it more realistic to assume that when there are a small number of gaseous water molecules, there will be a small number returning and when there are a large number of gaseous water molecules, there will be a large number returning. A simple way to model this is to postulate that a certain percentage of the gaseous water molecules will return to the liquid each second. For simplicity, let's assume that 50% of the gaseous molecules return every second. Using these assumptions, we are now in a position to determine the number of gaseous molecules as a function of time. This calculation has been started in the table below. Complete this table, rounding your results to the nearest whole number. Note that the number of gas molecules one second later (right column) becomes the beginning number of gas molecules for the next second.

Time (sec)	Beginning # of gas molecules	# leaving (constant) # returning (50% of # at left)	New # of gaseous molecules one second later
0	0	100 0	100
1	100	100 50	150
2	150	100 75	175
3	175	100	
4			
5			
6			
7			
8			
9			
10			
11			
12			

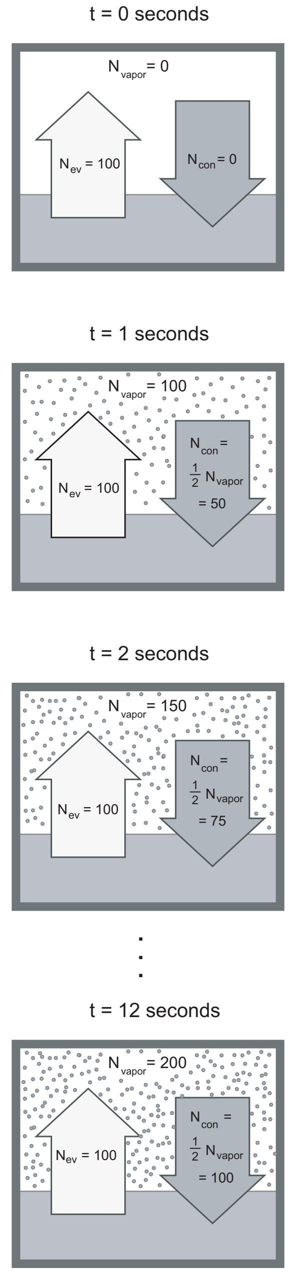


Figure C-11: If molecules leave (evaporate) the liquid at a constant rate N_{ev} each second and return (condense) to the liquid at a rate N_{con} per second proportional to the number of molecules in the gaseous phase, the system will eventually reach a dynamic equilibrium in which the total number of molecules in the gas (vapor) phase (N_{vapor}) is constant.

e) Describe what happens after about 10 minutes, and using the data from your table, make a rough sketch of the number of water molecules in the gaseous state as a function of time.

Gaseous water molecules are also referred to as *water vapor*. Thus, the graph you just made depicts the amount of water vapor in a sealed container as a function of time. Initially, the amount of water vapor increases quite rapidly, but then levels off to a constant value. This happens when the *evaporation rate* (number of molecules leaving the liquid each second) is exactly equal to the *condensation rate* (number of molecules returning to the liquid each second). This is what scientists call a *dynamic equilibrium*. This simply means that although the amount of water in the glass does not appear to be changing, it is actually losing and gaining molecules at exactly the same rate so that the average number of molecules in the liquid does not change. Although the mathematical model we used in the previous activity is quite simple, it captures the main features of evaporation in a closed container. In reality, the actual percentage of molecules evaporating and returning both depend on many factors, such as temperature and the amount of air above the liquid.

When the rate of evaporation is equal to the rate of condensation, we say that the air is *saturated*. This simply means that no more water molecules can coexist with the air. If more water molecules are placed into the air, they simply *condense* back into liquid water. Thus, there is no *net* evaporation. This is why a glass of water that is covered will not disappear. After a little time, there is enough water vapor coexisting with the air above the liquid so that an equal amount is being condensed back into the water as is being evaporated away. The amount of water vapor that can co-exist with the air at saturation is called the *equilibrium value* and must be experimentally determined. It has been observed that the equilibrium value for water vapor in air increases as the temperature increases.

When the amount of water vapor is far from its equilibrium value, the rate of evaporation is much larger than the rate of condensation and there is a net evaporation. But as the amount of water vapor in the air gets closer and closer to its equilibrium value, the number of molecules that condense back into the liquid state increases and this slows down the net rate of evaporation. Thus, things evaporate more quickly when there is a small amount of water vapor in the air and things evaporate more slowly when there is a large amount of water vapor in the air.

Evaporation, Humidity, and "Wet" and "Dry" Temperatures

Remember that this investigation started with the observation that a thermometer surrounded with a damp towel read a *lower* temperature than a dry thermometer. This led to the idea of evaporative cooling and the factors that affected evaporation. We now return to our wet and dry thermometers for a thought experiment.

Activity 3.3.2 Water Vapor Indicator

a) We saw earlier that wet-bulb and dry-bulb temperature measurements are not equal. Would you expect a higher temperature difference between these readings when there is more or less water vapor in the air? Explain.

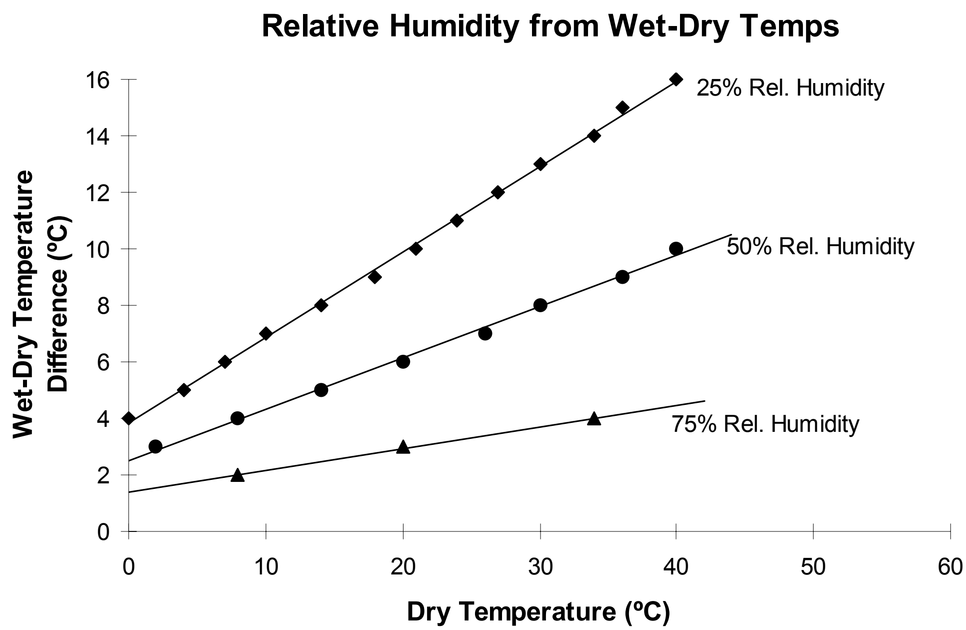
b) How do you think the wet-bulb and dry-bulb readings would compare with each other if the air were saturated? Explain.

c) Explain how you could use wet-bulb and dry-bulb temperature readings to tell you something about the amount of water vapor in the air.

The difference between wet-bulb and dry-bulb temperatures cannot tell you exactly how much water vapor is co-existing with the air. It can only tell you how much water vapor there is relative to the amount at saturation. This is an important point because as already mentioned the saturation value depends on temperature. Thus, the wet-bulb and dry-bulb readings will be the same whenever the air is saturated, even though the actual amount of water vapor in the air might be very different. This relative measurement of water vapor can be used to define a quantity called the *relative humidity*. Relative humidity is defined as the amount of water vapor in the air divided by the amount of water vapor in the air at saturation (the equilibrium value). Thus, relative humidity tells you, on a percentage basis, how close the water vapor content of the air is to it's equilibrium value.

Because the amount of water vapor at saturation depends on temperature, relative humidity itself depends on temperature. In fact, this temperature dependence plays an important role in the formation of clouds, as we will see shortly. Unfortunately, it makes the determination of relative humidity a little bit more complicated because the wet-bulb and dry-bulb temperature difference (sometimes just called the wet-bulb depression) will change depending on both temperature and relative humidity! However, the meaning of relative humidity is always the same. A reading of 50% relative humidity means the water vapor content of the air is half its equilibrium value. A 90% relative humidity reading means that the water vapor content of the air is 90% of its equilibrium value.

The following graph shows how you can use the wet-bulb and dry-bulb temperatures to determine the relative humidity. Notice that the same wet-dry temperature difference can correspond to different values of relative humidity depending on the (dry) temperature of the air. Again, this is because the equilibrium value of water vapor in the air depends on temperature. You can determine the relative humidity for values not shown on the graph by estimating where they would lie on the graph.



Using an Electronic Sensor to Measure the Relative Humidity

In addition to wet-bulb and dry-bulb temperature readings, relative humidity can also be measured using electronic components that undergo a physical change depending on the humidity (the same way our temperature sensors use electronic components that undergo physical changes depending on the temperature). The following activity will have you measuring the relative humidity of a sealed container. When using the relative humidity sensor, be careful to keep it out of water.

Activity 3.3.3 Measuring Relative Humidity

a) Place the relative humidity sensor on a small stand inside a large plastic container (a small cooler works well). You may want to tape the sensor to the stand so that it doesn't fall off into the water. Now set the experiment length to 15 minutes and use a data rate of 1-2 points per second. Before pouring any water into the container, begin the experiment and let it run for a minute or two so that you have a good reading of the relative humidity of the room. Then, pour in a cup or two of hot tap water (enough to completely cover the bottom of the container) and seal the container. Describe below what you expect will happen to the relative humidity readings as time goes on. Make a sketch of what a graph of relative humidity versus time would look like.

b) When the experiment ends, print out a copy of your data for your activity guide. Does the relative humidity reading reach 100%? Explain why this makes sense. If it doesn't reach 100%, can you explain why?

c) Compare your graph to the prediction you made as a result of your dynamic equilibrium model from Activity 3.3.1. Do the graphs have the same shape? Explain why you might expect these two graphs to look similar.

Checkpoint Discussion: Before proceeding, discuss your ideas with your instructor.

The last few activities have introduced a number of new terms and concepts such as relative humidity, saturation, wet-bulb depression, equilibrium value, etc. It is beneficial to discuss these terms and concepts with your group to make sure that you understand what they refer to. We will be using these terms in the next section when we discuss cloud formation.

4 CLOUD FORMATION

Have you ever wondered exactly what a cloud is? Many people are surprised to find out that clouds are nothing more than a collection of tiny droplets of water (or possibly ice crystals). Indeed, looking up into the sky and seeing huge white billowy clouds often makes one think more of cotton than droplets of water. On the other hand, you know from experience that it only rains when there are clouds present, so maybe it is not all that surprising that clouds are nothing more than water droplets. In this section, we will be taking a closer look at how clouds form. By the time we are finished, you should understand exactly what ingredients are necessary for clouds to form, and you will get a chance to produce your very own cloud.

You may need some of the following equipment for the activities in this section:

- Petri-dish, small stand, and salt [4.2]
- Styrofoam cup [4.2]
- Fire syringe demo [4.2]
- Small can of compressed air [4.2]
- Small mass (~50 g) [4.2]
- Flask with one hole stopper [4.3]
- Rubber tubing and connectors [4.3]
- Small hand pump [4.3]

4.1 GAS COOLING AND DEW POINT

In order to understand cloud formation, it is essential that you know how the relative humidity changes with temperature. You should recall that the equilibrium value of water vapor in air increases with temperature.

Activity 4.1.1 Humidity and Cooling

a) Using the definition of relative humidity, explain why the relative humidity decreases when the temperature rises, even if no water vapor is added to the air.

b) Now imagine a parcel of air that is being cooled with no change in its water vapor content. In other words, the air is getting cooler but the number of evaporated water molecules in the air is unchanged. Explain how its humidity will change.

c) If the temperature of this parcel of air continues to drop, describe what happens to the relative amounts of evaporation and condensation.

Relative Humidity and Clouds

It is important to remember that the relative humidity does not tell us absolutely how much water vapor is in the air. Instead, it indicates the ratio of the amount of water vapor in the air to the maximum amount that *could* be in the air. The fact that the equilibrium value of water vapor in cooler air is smaller means that the relative humidity of the air will go up as the air is cooled. The graph below shows the equilibrium value of water vapor as a function of temperature. **Note:** Recall that after Activity 3.3.1, we defined the *equilibrium value* of water vapor in the air to be the amount of water vapor that will co-exist with the air once the system has reached equilibrium (i.e., the evaporation and condensation rates have reached the same value)

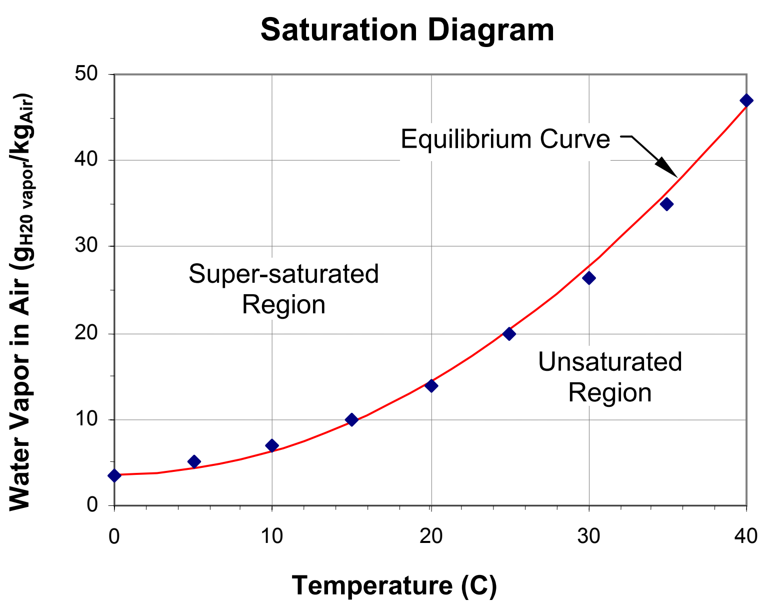


Figure C-12: A saturation diagram depicts the equilibrium curve for water-vapor. That is, the amount of water vapor that will co-exist with the air at equilibrium, as a function of temperature. The region below the curve includes systems that are unsaturated (the evaporation rate is larger than the condensation rate). The region above the curve includes systems that are saturated (the evaporation rate is smaller than the condensation rate). Systems lying on the equilibrium curve are referred to as saturated because the evaporation rate and the condensation rate are equal.

To get this *equilibrium curve*, we have plotted the maximum possible water vapor content of air at different temperatures. Note that the water vapor content is measured in units of grams of water per kilogram of air, so a measurement of 5 g/kg means that there are 5 grams of water vapor in every kilogram of air. Unlike the relative humidity, this is an *absolute* measurement of how much water vapor is in the air. The equilibrium curve represents the line of 100% relative humidity.

Consider a system (of air and water vapor) at a temperature of 15°C with 10 grams of water for every kilogram of air. Looking at the graph, you see that this air, represented by a point at 15°C and 10 g/kg, is sitting on the equilibrium curve. This means the water vapor is at its equilibrium value, i.e., the system is *saturated* and has a relative humidity of 100%. If this system suddenly warmed to 30°C without gaining any extra water, there would still be 10g of water vapor for every kg of air. This point, at 30°C and 10g/kg, lies below the equilibrium curve, which has a value of 27.5 g/kg at 30°C. This means that another 17.5 grams of water can co-exist with each kilogram of air (at this temperature). The region below the equilibrium curve is called the *unsaturated* region, since an air-water system in this region is not fully saturated. The relative humidity for this system at 30°C is calculated by dividing the actual water vapor content of the system by the equilibrium value at 30°C. That is,

$$\frac{10g/kg}{27.5g/kg} = 0.364 = 36.4\%\,.$$

Note: Strictly speaking we should always refer to "the air-water vapor system" when talking about relative humidity and saturation. In practice, however, it is often more succinct to refer to this system simply as *the air*. Thus, for convenience, we will often refer to a mixture of air and water vapor as a *parcel of air*.

The above calculation raises an interesting question. What if a parcel of saturated air gets colder? The next few activities explore this question.

Activity 4.1.2 Calculating Relative Humidity Changes

a) If you started with air that had a relative humidity of 50% at 40 degrees Celsius and you cooled that air down to 30 degrees Celsius (without changing the water vapor content), determine how the relative humidity changes.

b) If you started with air that had a relative humidity of 20% and a temperature of 20 degrees Celsius, and you added (for example, by boiling water) 5 grams of water vapor to the air *without* changing its temperature, determine how the relative humidity would change.

c) The *dew point* is defined to be the temperature at which a given parcel of air becomes saturated. This is the temperature at which the air has a relative humidity of 100%, assuming no water vapor is gained or lost. What is the dew point if the temperature is 25°C and the relative humidity is 50%?

d) What is the dew point if the air is 30°C with a relative humidity of 30%?

e) What do you think happens to the water in the air as it cools below the dew point?

Checkpoint Discussion: Before proceeding, discuss your ideas with your instructor.

Supersaturation

We have defined the saturation curve as representing the maximum possible amount of water vapor that air can hold. This is in fact not completely correct. If you cool air very quickly then the water vapor does not have time to condense out of the air. This system is then said to be *supersaturated* and is represented by points in Figure C-12 above the saturation curve. In this state, however, the water vapor in the air is very *unstable*, meaning the slightest change causes condensation. When this happens, the water vapor often, but not always, condenses into very small water droplets that we see as fog or as a cloud.

4.2 UNDERSTANDING CLOUD FORMATION

There are many different kinds of clouds and many different ways for them to form. We will limit ourselves to two situations. The first is when the temperature of air suddenly changes, as can happen in the atmosphere when thermal energy is absorbed from the sun or when air moves to colder regions. The other situation of interest is when two different parcels of air come into contact with each other. We call this a *mixing* cloud because the two air masses mix together, and the properties of the combined air mass can be quite different than either of the constituent air masses.

What Factors Contribute to Cloud Formation?

It is interesting to note that a cloud does not always form when the air reaches saturation. Consider a closed bottle of water. If the water has been sitting for a long time, the relative humidity of the air above the liquid will be 100%. This means the system is saturated. Nevertheless, you don't often see little clouds in bottled water (you do however, see clouds when opening a bottle containing a carbonated beverage). Clearly, saturated air is only one ingredient in cloud formation. The water vapor that exists in a saturated (or supersaturated) system is ripe to condense into water droplets. The following activity introduces you to the essential ingredient that leads to the condensation of water in the form of a cloud.

Activity 4.2.1 Condensation Nuclei

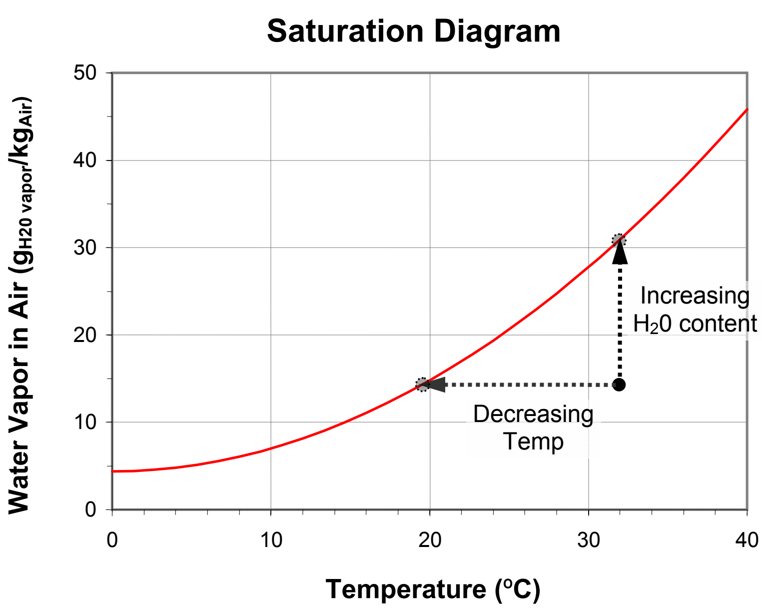
a) Take a small Petri-dish or other small container and place a little room temperature water in it. Next, place a small platform (the bottom ¼" or so of a Styrofoam cup works well) in the water so that there is a dry, flat surface above the surface of the water. Now, half of the groups should place about 10 grains of salt on the small platform and the other half should leave theirs empty. Lastly, half of the "salt" groups and half of the "non-salt" groups should cover their platforms with a Styrofoam cup (you may need to place a small mass on top of the cup to keep it from floating). Gently put this little contraption aside so that it won't get bumped, we will be checking back on it later

b) After about 45 minutes, take a look at the platforms that were left uncovered. Describe what you observe. Make sure you check experiments that had salt and those that didn't. Is anything unusual occurring on the platforms?

c) Now remove the cups from the other experiments and look at the platforms. Describe what you observe. Again, make sure to observe those that had salt as well as those that didn't.

d) For those experiments that were covered, what do you think the relative humidity was inside the cup? **Hint:** Do you think there was any net evaporation in this situation? Do you have any ideas why there was condensation on the grains of salt but no where else? Explain.

As seen in the last activity, the presence of the salt particles helps *encourage* condensation, but only when the relative humidity was 100%. The salt particles (or any other particles that act in this way) are called *condensation nuclei*. We will see later that condensation nuclei can have a large impact on whether or not a cloud can form.



Now, before attempting to make a cloud, we need to understand how a system becomes saturated. There are two basic mechanisms by which a parcel of air can become saturated. One is by increasing the water vapor content of the system and the other is by decreasing the temperature of the system. These two processes are sketched on the saturation diagram above. Of course, real clouds usually result from a combination of these processes.

We will be making a cloud in a bottle by decreasing the temperature of the air. But first, we need to understand how to cool down a parcel of air. This is the topic of the next activity.

Activity 4.2.2 Expanding and Compressing Air

a) What do you think would happen if we were to rapidly compress some air? Would its temperature rise, fall, or stay the same? What makes you think so?

b) Now observe what happens to the temperature of a gas that is compressed. Take a well-sealed compression piston, known as a "fire syringe," and *very rapidly* compress the air in the tube. Describe what you observe (have your instructor help you with this). Can you explain this behavior using the concept of molecular motion?

c) What do you think would happen if you were to allow the air to rapidly expand instead of compressing it? Explain.

d) Now take a small can of compressed air and let some of the air escape. This air is rapidly expanding. Describe what you observe when you let some air out of the can. How does the can feel?

The results of the previous activity are pretty dramatic and may have surprised you a bit. Recall that in the beginning of this unit, we increased the temperature of water just by shaking it. That is, we did *work* on the water to increase its temperature. When compressing the air, we are similarly doing work on the air and likewise, its temperature increases. Of course, the temperature of the air increases a whole lot more than the water, but that's because there is so much less air than water. On the other hand, when the gas is escaping from a can of compressed air, the gas itself is doing work as it pushes air out of the can. For our purposes, we are only interested in the fact that when air is rapidly released from a container the temperature of the air left inside drops dramatically.

4.3 MAKING A CLOUD

We now have all the ingredients we need to actually create our very own cloud! This is the topic of the next activity.

Activity 4.3.1 A Cloud in a Bottle

a) Take a small flask and put about 10-20 grams of water in it. Next, connect a rubber tube from a small hand pump to the flask with a rubber stopper. You are now ready to try and make a cloud as follows: Pump some air into the flask with the hand pump and then let the flask sit for a moment. After letting the flask sit for a few minutes, what are the temperature and relative humidity of the gas inside the flask?

b) If you were to quickly allow the air to escape, how would the temperature of the air in the flask change? What would this do to the relative humidity? Would you expect to see a cloud?

c) After waiting for a minute to allow the gas to equilibrate, quickly pull the stopper out of the flask. Did you see a cloud? If you didn't, can you think of any reason why not?

Figure C-13: With an understanding of how temperature and humidity interact, it is possible to make a cloud in a bottle.

d) This first attempt at making a cloud may have been a dismal failure. The reason is that there may not have been enough condensation nuclei around. Try the experiment again, only this time, drop a lit match into the flask (with water) before putting on the rubber stopper. This will put many smoke particles in the flask to act as condensation nuclei. Once again, pump some air into the flask and then let it sit for a brief time. Explain what happens this time as you allow the air to expand suddenly.

Checkpoint Discussion: Before proceeding, discuss your ideas with your instructor.

Congratulations! You have just created your very own cloud. Now, in case this activity seems contrived to you, it is important to point out that this process is actually quite realistic. That's not to say that clouds actually form in bottles, but that air masses often cool down and become saturated.

Cloud Levels

Under normal conditions, the temperature of air decreases by about 2°C with every 1,000 ft. of altitude gained. Thus, as air rises it will cool. There are a number of reasons why air might rise. Wind may push an air mass against the side of a mountain forcing the air mass to rise. Two winds blowing in opposite directions might collide and force air upward. Whatever the reason, it is common for air to rise and if, as a result, the air cools to the dew point, there is a strong possibility that a cloud will form.

Activity 4.3.2 Cloud Levels

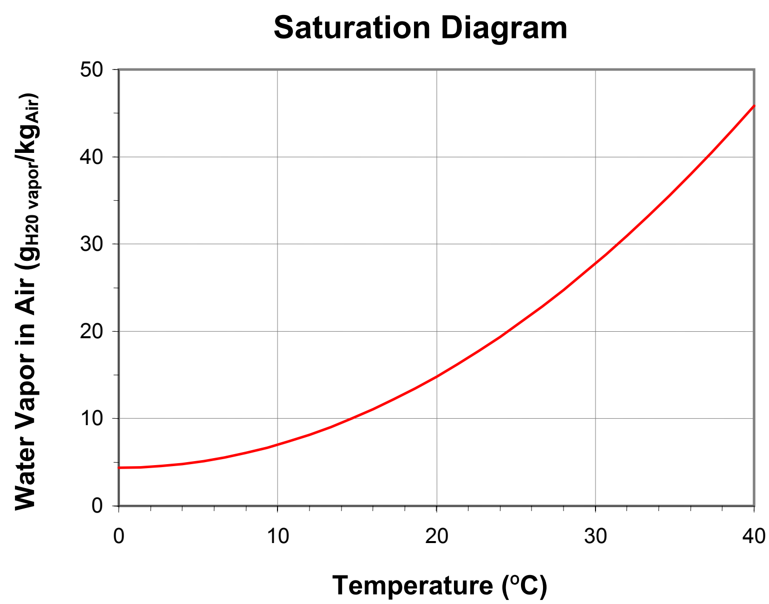
a) If the temperature at the surface of the Earth is 30°C and the relative humidity is 50%, at what altitude are you likely to find clouds? Explain. **Hint:** Air at the surface will probably get pushed upwards

Cloud Formation Due to Air Mixing

The final topic that we will undertake in this unit is that of a *mixing cloud*. The most common example of a mixing cloud is the cloud that forms from your breath on a cold winter day. This type of cloud forms when two masses of air mix together. The two masses of air might have different properties or they might be very similar. For example, one might be hot and humid and the other might be cold and dry. To understand the transformation of the air as it mixes together, it is instructive to consider carefully what happens to the temperature and the vapor content of the air.

Activity 4.3.3 Mixing Air Masses

a) Let us consider two parcels of air. The first one is at 20°C and has a water vapor content of 5 g/kg and the second one is at 30°C and has a water vapor content of 15 g/kg. Plot these two air masses as points on the following saturation diagram.



b) Now consider the case where both packets of air are the same size. For example, say both contain one kilogram of air. When they mix together, there will be two kilograms of air. What is the total amount of water vapor if these two packets combine? What is the water vapor content of the mixture in grams per kilogram of air?

c) Determining the final temperature of the new air mass is exactly the same as determining the final temperature after mixing two cups of water together. The final temperature will depend on the amounts and temperatures of the initial parcels of air. In this case, the initial amounts are the same, so the calculation is easy. Find the final temperature of the mass of air and, using the water vapor content from the previous question, plot the point that represents the final air mass on the saturation diagram. Describe where it lies in relation to the two initial points.

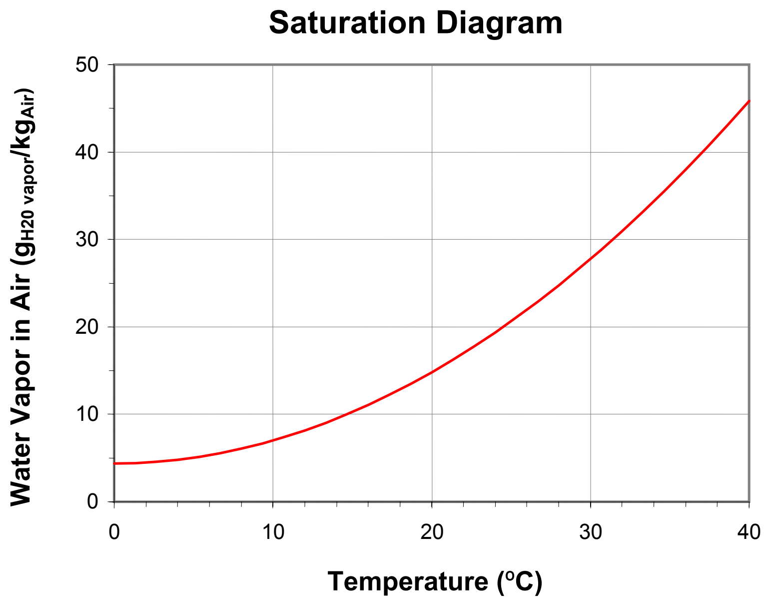
d) What would the final relative humidity be if, instead of one kilogram of each air packet, we had one kilogram of 30 °C air and two kilograms of 20 °C air? Where would it lie on the saturation diagram?

Notice that in the above activity, the final state after mixing lies directly between the initial points. The fact that it lies exactly in the middle is because we began with the same amount of air in the initial parcels. If we had started with one parcel being larger than the other, the final point would not be exactly in the middle. However, the final point would like *somewhere* on the line that joins the two initial points. In fact, it will be proportionally closer to the initial point that had more mass. Thus, if one parcel of air had three times the mass of the other, then the final point would lie three times closer to that point than the other, but still on the line that joins the two initial points.

Now consider the case where the two initial parcels of air are such that when you draw a line connecting them on a saturation diagram, that line goes into the super-saturated region. That suggests that there is the possibility of observing a cloud. Of course, it will depend on the initial masses of the air parcels and also on whether there are any condensation nuclei around. So, while there is no guarantee that a cloud will be visible, it is at least a possibility.

Activity 4.3.4 Mixing Clouds

a) Consider being outside on a cold winter day when the temperature is 2°C and the relative humidity is 50%. When you breath, you blow out air at that is near your body temperature of about 37°C with a high relative humidity of about 95%. Find these points on the following saturation diagram and connect them with a line. Do you think there will be a cloud on your breath? Explain your reasoning.



Checkpoint Discussion: Before proceeding, discuss your ideas with your instructor.

It is surprising just how often mixing clouds will form. In fact, now that you understand where they come from, you will probably start noticing them all over the place. They happen at the exhaust pipes of cars, jet aircraft sometimes make them in the sky, you see them when you take a hot shower, and also when boiling water to make tea. The next time you see a mixing cloud, it may be difficult *not* to think about the water vapor content of the air masses and how their mixing leads to the air being in the super-saturated region of the saturation diagram.