CHAPTER 1

Weather Variables

Where should we start with our study of the atmosphere? How should we first approach the weather? Like many scientists, meteorologists first make observations. Then they raise questions, and try to answer them. In this first chapter, we will quickly describe four of the elements, also called variables, of weather that meteorologists regularly observe, measure, and chart on weather maps, before we return to each of them for a more thorough exploration in subsequent chapters. Three of these elements are fairly intuitive: when concerned with the weather, we like to know how warm or cold it will be (temperature), whether it will be windy or not (wind), and whether it will rain or not (precipitation). The fourth variable, atmospheric pressure, is less intuitive, but it may be the most important to a meteorologist, as we will soon discover.

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Weather results from atmospheric changes. These changes obey certain rules, dictated by the laws of physics. In meteorology, three laws are of particular importance: the law of conservation of energy, the law of conservation of mass, and the law of conservation of momentum. Each describes a particular aspect of the atmosphere, and each requires that we measure certain **variables** of the atmosphere. The objective of this first chapter will be to provide an overview of these variables, a starting point for our exploration of the atmosphere. We will then return to each variable in subsequent chapters for a more thorough description and analysis.

1.1 Temperature

Of primary interest to us is the law of conservation of energy (see Box 1.1). It states that energy is never created or destroyed, but only transferred between locations or transformed between different types of energy. In the atmosphere one form of energy is **heat**,

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Box 1.1. The Law of Conservation of Energy

The law of conservation of energy states that the total energy of a system remains constant, if we account for the gains and losses of energy from and to the outside. Energy can be *transferred* between different parts of the
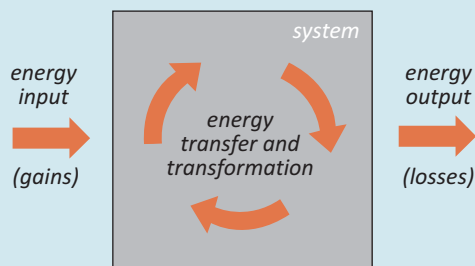

Figure 1.1.1. The law of conservation of energy applied to a system.

system, or *transformed* into different types of energy inside the system (e.g., from potential energy to kinetic energy, or chemical energy, or radiative energy), but energy cannot be created (out of nothing) or destroyed.

If we think of the atmosphere as a system, we can apply the law of conservation of energy to describe how energy, and in particular heat, is transferred or transformed in the atmosphere. And since heat transfers are related to temperature differences, we need to measure temperature and map these differences. That is one reason why temperature measurements are an integral part of weather observations.

We will return to the conservation of energy in Chapter 6, in the form of the first law of thermodynamics.

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and weather is largely the result of heat contrasts and heat transfers. Therefore we need to design ways of describing the amount and fluxes of heat throughout the atmosphere, which is accomplished by measuring **temperature**.

1.1.1 Heat and Temperature

We all have an intuitive feel for temperature, for whether things are hot or cold. But that feeling is very subjective – warm water feels cold after stepping out of a sauna. Science requires objectivity, which is sought by measuring and quantifying variables and processes. But how do we quantify that feeling of warm and cold? What is temperature, really? First, we need to return to the fundamental definition of **heat** as it is transferred to our body and the environment by contact and interactions at the molecular level.

We can think of the atmosphere as a mixture of gases made up of molecules in motion (Figure 1.1). And we can think of heat as the energy associated with this molecular motion. In warm air, molecules are moving more rapidly than in cold air, and therefore have more energy of motion. We call this energy of motion, **kinetic energy**. Temperature is an indirect measurement of the *average* kinetic energy of the molecules constituting the air. Here it is useful to
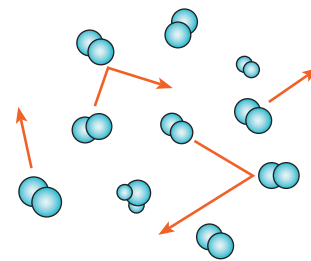

Figure 1.1. Random molecular motion in a volume of air.

think of the air around us as made of individual volumes of air of uniform characteristics, called *air parcels*. (Although the size of such volumes of air is somewhat arbitrary, the concept of an air parcel will be useful later to explain various processes at work in the atmosphere.) While individual molecules in the parcel might all have different speeds, their average state is indicative of a higher or lower energy level, hence a higher or lower temperature of the air parcel.

When fast-moving air molecules (at a higher temperature) are in contact with slow-moving air molecules (at a lower temperature), they impart some of their energy to the slow molecules, by contact, through collisions. In doing so, they *lose* energy, and therefore cool down, while the slower molecules *gain* energy, and therefore warm up. Kinetic energy, and

therefore heat, has been transferred from the faster air molecules to the slower molecules, from the warmer to the colder air. We will see in Chapter 4 that this is the basic mechanism for transferring energy at the molecular level through **conduction**, but it is not a primary heat transfer mechanism for most of the atmosphere.
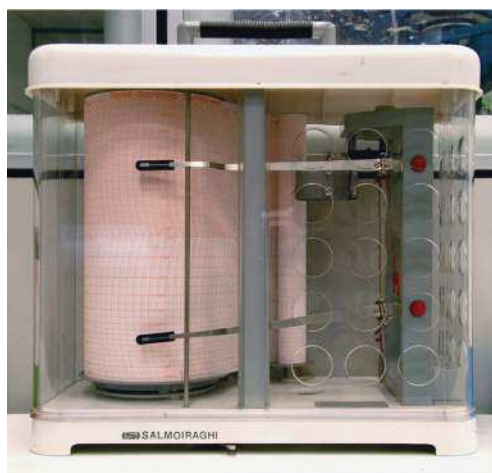
1.1.2 Thermometers

How do we measure temperature? We make indirect measurements of average kinetic energy using **thermometers**. Liquid-in-glass thermometers, for example, work on the basis that substances expand or contract when temperature increases or decreases. A glass tube is filled with a liquid such as alcohol or mercury. When placed in warmer air, the energy of motion of the air molecules is transferred to the liquid in the reservoir by conduction through the glass. The molecules constituting the liquid now have more energy than before. Being more active, they push each other apart, which makes the liquid expand. In particular, it makes the liquid rise in the tube. Exposed to the same temperature, the liquid will always rise to the same level. Thus, if we attach a scale and units to the glass tube, and calibrate the instrument against known temperatures (such as the freezing point and the boiling point of water), we obtain an instrument that can measure any temperature (Figure 1.2).

Many other types of thermometers exist, and all make use of a property of matter to determine temperature indirectly (Figure 1.3). A bimetallic strip, for

(a)

(b)

(c)

Figure 1.3. Different types of thermometers. (a) In a thermograph, a bimetallic strip bends and moves a pen up and down while a roll of paper turns underneath. (b) In a digital thermometer, the electrical resistance of a thermistor changes with temperature. (c) In an infrared thermometer, the radiation emitted by bodies is used to infer their temperature.

Figure 1.2. Thermometer in a standard weather shelter.

example, is made of two thin pieces of different types of metal attached to each other. Because the metals expand and contract at different rates, the combined strip bends when the temperature changes. If the tip of the metal strip is made to bend toward a temperature scale, the device can be calibrated and turned into an instrument.

Some electronic thermometers use a material whose electrical resistance depends on temperature (a thermistor). Others, called radiometers, measure the radiation emitted by bodies. Because the emitted radiation is a function of the temperature of these bodies, we can once again indirectly deduce their temperature, as we will describe in Chapter 4.

1.1.3 Temperature Measurements

As we will discuss at the end of this chapter, temperature measurements, like other weather measurements, are standardized, so that they can be compared and mapped. Weather stations always measure temperature in the shade, to avoid contamination by sunlight – recall that we are interested in measuring the kinetic energy of the air molecules surrounding the thermometer, and not the radiative energy contained in sunlight that might be absorbed by the thermometer. For this reason, temperatures are measured in a shelter, which is elevated at 1.25–2 meters (4–6½ feet) height above a vegetated surface, to avoid contamination by surface effects. Indeed, the air temperature can change rapidly near the ground, even within 2 meters (see Chapter 4 and Figure 4.14 in particular), and it is important that we measure the temperature of the *air*, not that of the surface. The shelter is painted white, and ventilated, to limit the absorption of sunlight and the concentration of heat inside the box, which could otherwise produce an artificially high temperature.

1.1.4 Temperature Scales

You are probably familiar with the Fahrenheit temperature scale, named after Gabriel Fahrenheit, a German scientist who constructed the first mercury thermometer in the eighteenth century and calibrated it against three fixed points: 0, as defined by a mixture of ice, water, and sea salt; 32 in water and ice; and 96

"in the mouth or armpit of a healthy man." Degrees Fahrenheit are commonly used in the United States. Scientists, and most other countries of the world, prefer to use the Celsius (or Centigrade) temperature scale, named after Anders Celsius, a Swedish astronomer who also lived in the eighteenth century and proposed the temperature of melting ice and the temperature of boiling water as fixed points (0 and 100, respectively). To convert from one scale to another, use the following formulas:

$$°C = 5/9 \times (°F - 32°)$$
$$°F = [9/5 \times (°C)] + 32°$$

(Note that, in the first formula, 32 is subtracted *before* multiplying by 5/9, while, in the second formula, 32 is added *after* multiplying by 9/5 – a common source of mistake.)

Scientists also use the Kelvin temperature scale, named after Lord Kelvin, a nineteenth-century scientist whose original name was William Thomson. The Kelvin scale is merely an offset version of the Celsius scale, translated in such a way that temperature measurements are always positive numbers.

$$K = °C + 273.15°$$
$$°C = K - 273.15$$

The smallest possible temperature, zero kelvin (0 K), is called "absolute zero," and corresponds to the theoretical state in which all molecular motion stops, in which case molecules have zero kinetic energy, and therefore zero temperature.

(Note that the symbol for kelvin (K) does not have the degree symbol – the little circle – in contrast with °F and °C.)

Here, in keeping with the International System of Units, we use degrees Celsius on weather maps and temperature profiles. If you are more accustomed to degrees Fahrenheit, however, it will be useful to be able to convert from one scale to another, using the formulas above, or Table 1.1 for quick reference. If you hear or read temperatures in degrees Celsius and want to convert them quickly to degrees Fahrenheit, you can take advantage of the fact that 9/5 is approximately equal to 2, and 32 is approximately equal to 30. Thus, you can use the following approximation:

$$°F \approx (2 \times °C) + 30°$$

Table 1.1.   Common temperature values in degrees Celsius and Fahrenheit

°C	°F	°C	°F
−40	−40	**0**	**32**
−35	−31	5	41
−30	−22	10	50
−25	−13	15	59
−20	−4	20	68
−15	5	25	77
−10	14	30	86
−5	23	35	95
0	**32**	40	104

Figure 1.4. Temperature profile obtained by radiosonde at Amarillo, Texas, on February 18, 2014, at 00:00 UTC. The red curve indicates temperature (T), and the green curve indicates the dew point temperature (T_d). Pressure levels (given in hectopascals, hPa) are shown in blue and are described in the text.

This is a very quick calculation that you can easily do in your head: multiply by 2 and add 30. For example, 20°C multiplied by 2 is equal to 40. Add 30 to obtain 70°F, which is close enough to the exact conversion, 68°F.

1.1.5 Radiosonde Profiles

As we will see later in the book, a large part of our weather is dictated by what takes place aloft, as weather systems often extend up into the atmosphere and surface phenomena are often driven by upper air currents. Therefore it is useful to obtain information about the vertical structure of the atmosphere (Figure 1.4). We do so by launching helium-filled balloons carrying instruments that record weather variables up to 35 km altitude as the balloons ascend – helium is used because it is a very light gas (Figure 1.5). These balloons are called **radiosondes**, since the measurements are radioed back to a receiving station at the surface. (See Box 1.2 about the need for upper-air data and the development of the radiosonde during the Second World War.)

Figure 1.4 shows the lower section of a temperature profile, in red, obtained by launching a radiosonde from Amarillo, Texas, on February 18, 2014, at 00:00 UTC. (Except where specified, we use Universal Time Coordinates (UTC) in the rest of the
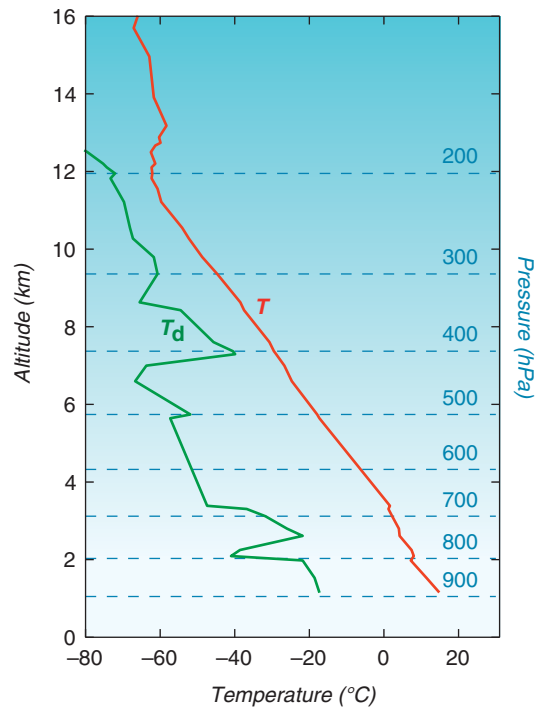

Figure 1.5. Radiosonde being launched in Hawaii.

Box 1.2.   Meteorology in Times of War

Many scientific advances have been made in response to military incentives. During the Second World War, understanding and predicting the upper-level atmospheric circulation became of prime importance as all parties involved sought control of air power, and air supremacy would be decisive in winning the war. The need for upper-air data greatly accelerated the development of radiosonde components and observational networks.

Figure 1.2.1. Florence van Straten.

A notable figure of the Second World War who contributed to this effort is Florence van Straten (1913–1992). She volunteered as part of the Women Accepted for Volunteer Emergency Service (WAVES) program, which had been created to meet the demands for trained military personnel. She was assigned to the Naval Aerology Service and trained as a weather forecaster at the Massachusetts Institute of Technology. She made significant advances to naval meteorology both during and after the war. She developed methods for taking advantage of weather phenomena during military operations, such as the timing of storms and the cover provided by clouds and precipitation. She contributed to the development of the constant-altitude weather balloon and the radiosonde, as well as the "tipping bucket" rain gauge and the weather station shelter. She also studied weather modification and the impact of atmospheric conditions on radar measurements.
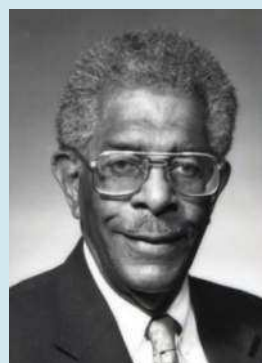

Figure 1.2.2. Charles E. Anderson.

Another notable meteorologist trained during the Second World War is Charles E. Anderson (1919–1994), the first African American to receive a PhD in meteorology (from the Massachusetts Institute of Technology). He was assigned to the meteorology division of the US Air Force and served as a weather officer during the war. He is remembered for his pioneering work on cloud dynamics (his dissertation was on the growth of cumulus clouds), forecasting severe storms, and reducing the contrails of high-altitude jet aircraft.

book, i.e., time referenced at the meridian of Greenwich, England.) The green curve indicates the dew point temperature, an important measure of the amount of water vapor in the air, which we will discuss in Chapter 5. Figure 1.4 also gives us an opportunity to describe a particular type of graph that we will encounter many times in this book. In such a graph, altitude is displayed on the *y*-axis (i.e., the vertical axis), while the variable of interest (temperature, pressure, wind speed) is displayed on the *x*-axis. Thus, the red curve in Figure 1.4 tells us that the temperature decreases from about 18°C at 1 km altitude to −60°C at 12 km altitude, which is in fact fairly typical. For reference, pressure levels are indicated by dashed blue lines, labeled in hectopascals (see Section 1.2.5). We can see that the temperature decreases up to about 200 hPa. We will soon learn that this level of the atmosphere (which changes with time and location) is called the *tropopause* and defines the upper limit of the *troposphere* – the layer of the atmosphere where essentially all weather takes place. The full vertical structure of the atmosphere will be described in Chapter 3.

1.2 Pressure

Atmospheric changes, and therefore weather, involve the movement of air, and the redistribution of air mass in the atmosphere. Such redistribution obeys the law of conservation of mass, which, in meteorology, is

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Box 1.3.   The Law of Conservation of Mass

Much in the same way as energy is conserved, matter, and therefore mass, is conserved in a system. It cannot be created out of nothing and cannot be destroyed: it can only enter or leave the system, or be transformed inside the system (by chemical reaction, for example, or by phase changes, as we will discuss in Chapter 5). This is described by the law of conservation of mass.

In meteorology, the distribution of mass in the atmosphere is described by the distribution of **density**, which is defined as mass per volume. However, because density is difficult to measure, we instead rely on measurements of pressure and temperature to infer density (see Box 1.4). As a result, we can consider pressure measurements as informing us about the distribution of mass in the atmosphere.

Notions of mass conservation and pressure will be important for understanding wind and weather systems in Chapters 8 and 10.

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expressed in terms of pressure (see Box 1.3). It is therefore important that we measure pressure to analyze and predict the weather. As we did for temperature, however, it is useful that we first understand the nature of pressure in the atmosphere before we describe instruments to measure it.

1.2.1 Force and Pressure

As we discussed for temperature, air is a gas made up of molecules in motion, and heat can be equated with the kinetic energy of the molecules. However, we also recognized that these molecules constantly bump into each other. Loosely speaking, the amount of bumping is what we call **pressure**. (We will come to a more exact definition shortly.) If the molecules bump into each other more frequently, or if the collisions themselves are stronger, there is more pressure. Thus, in a closed container, we could increase the pressure by either decreasing the size of the container at fixed temperature, which would bring the molecules closer to each other and would increase the number of the collisions, or increasing the temperature by heating the air, which would provide more kinetic energy to the gas molecules, and would result in more intense collisions. We can see here that pressure, temperature, and the volume of the container are indeed related, as expressed by the ideal gas law (see Box 1.4). However, this is not to say that pressure and temperature are the same thing. Temperature, as a measure of heat, relates to the average kinetic energy of the

molecules, independently of the collisions. Conversely, pressure relates to the collisions between molecules, which is not uniquely determined by the speed of the molecules, i.e., it is also a function of the number of molecules in the volume, and therefore the density of the air.

In the atmosphere, things are slightly more complicated, as the air is not enclosed in a container, but we will address that issue shortly. In the meantime, we can think of pressure as the amount of bumping between molecules. And since the molecules are moving in all directions, we can see that pressure is applied in all directions as well.

In practice, we often need to know what happens when air pushes on particular surfaces, either real, like the ground, or imaginary. For example, we sometimes think of the atmosphere as being made of air columns standing next to each other, and we are interested in knowing how much the air in two adjacent columns is pushing against the "wall" in between the columns. At other times, as described earlier, we will think of the atmosphere as being made of air parcels, i.e., volumes of air delimited by an imaginary envelope. Then we will be interested in knowing how much the air inside and outside the parcel pushes on the envelope. In all cases, pressure is applied to a surface, which results in a **force** pushing on that surface. As the area on which pressure is applied increases (i.e., more air molecules bump into that surface), the force increases proportionately. In symbolic notation:

> ### Box 1.4.   The Ideal Gas Law
>
> To very good approximation, the atmosphere behaves like an ideal gas, in which molecular collisions result only in a transfer of kinetic energy. As a result, pressure, temperature, and density do not vary independently of each other but are related by the ideal gas law,
>
> $$pV = nR^*T,$$
>
> where p stands for pressure, V denotes volume, n is the number of molecules in volume V, and T is temperature. R^* is a universal constant that applies to any gas.
>
> For the atmosphere, the ideal gas law is more conveniently written
>
> $$p = \rho RT,$$
>
> where ρ is the density (mass per volume) and R is a gas constant specific to the atmosphere.
>
> For example, if we heat a fixed volume and mass of air (i.e., the density remains constant), the kinetic energy of the molecules increases (i.e., the temperature increases), and the molecules bump into each other with more impetus (i.e., the pressure also increases).

$$F = p \times A$$

where p stands for pressure applied to a surface of area A, and F is the resulting force. Equivalently, pressure can be thought of as the force applied to a surface divided by the area of the surface:

$$p = F/A$$

which is, in reality, how pressure is defined in physics. In the above formula, if the force is distributed over a large area, the resulting pressure is small. But if the force is concentrated on a small area, the pressure is high.

A simple application comes to mind. If you are wearing regular shoes, your weight, which provides the force, is distributed over the entire sole of your shoes, which results in a small pressure being applied to the floor. With high heels, however, weight is concentrated on a very small area, which results in a much greater pressure, and could dent a soft floor surface.

1.2.2 Atmospheric Pressure

By analogy, we now consider the weight of the entire atmosphere applied onto Earth's surface at sea level. Air might seem weightless, as philosophers of antiquity, including Aristotle, used to think, but gravity pulls air downward, keeping the atmosphere, like everything else, around Earth. In a column of air extending from Earth's surface all the way up to the top of the atmosphere, the accumulated weight of all the air pushes down on the surface. Recall that weight is a force, so when we divide the weight of the air column by the area of the base of the atmospheric column, we obtain the pressure of the atmosphere, or **atmospheric pressure**.

Note that, in this explanation, we are thinking about atmospheric pressure as pushing essentially downward, but recall that pressure at a point acts equally in all directions.

1.2.3 Vertical Distribution of Pressure

We do not need to be at sea level for the concept of atmospheric pressure to make sense. We can repeat the same exercise at some altitude above sea level, say, 3000 m. Air pressure at that altitude is determined by how much the air is compressed due to the accumulated weight of the overlying layers of air. (Note that the amount of air below 3000 m is irrelevant in calculating pressure *at* 3000 m.) Since there is necessarily less air above 3000 m than there is above sea level, air pressure will be less than at sea level (Figure 1.6).

To make a familiar connection with this decrease in pressure with altitude, notice the feeling you experience when you swim to the bottom of a pool – you experience an increase in pressure in your ears and nose. The atmosphere is like a swimming pool, and we live at the bottom of the pool – a pool of air.

In the same way that a diver will experience higher pressure at depth in the ocean, atmospheric pressure is highest at ground level, and decreases upward (Figure 1.7 and Box 1.5 – pressure units are defined in Section 1.2.5).

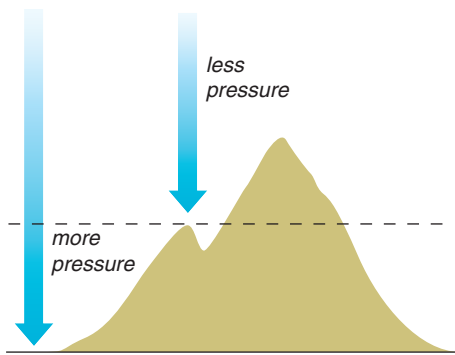In summary, it is useful to think of atmospheric pressure as the weight of air above a point, or above a unit area. If we are at sea level, we are experiencing the weight of the total atmospheric column, and therefore higher pressure. If we are higher up in the atmosphere, the column above us contains less air than at sea level, and as a result the pressure is lower.
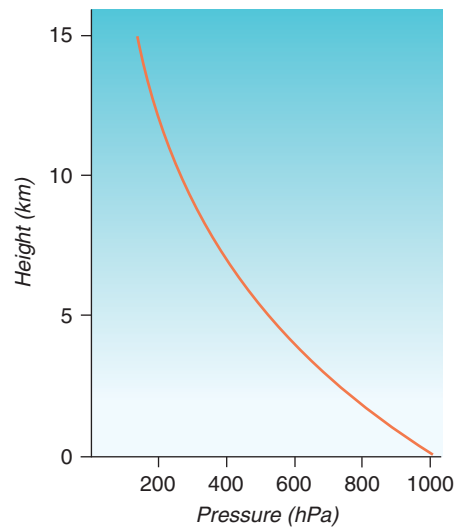

Figure 1.6. Atmospheric pressure is tantamount to the weight of the atmosphere above us.

Figure 1.7. Pressure decreases with height.

Box 1.5.   A Compressible Atmosphere

You might have noticed in Figure 1.7 that the rate of decrease of pressure with height is greater near the ground than higher up in the atmosphere. For example, atmospheric pressure drops by about 200 hPa from 0 to 2 km altitude, whereas it drops by less than 50 hPa from 13 to 15 km altitude. This is due to the fact that, in addition to experiencing the accumulated weight of the overlying atmosphere, air is also compressible. Air at lower altitudes experiences more weight from the air above than at higher altitudes, and is compressed more than air at higher altitudes. As a result, a layer of air containing the same amount of air (weight) is thinner at lower altitudes. If air were not compressible, the graph in Figure 1.7 would be a straight line: it would show a linear decrease of the weight of the overlying air column with height (see the dashed lines representing layers of equal thickness in Figure 1.5.1(a)). Because a layer of air is more compressed in the lower atmosphere, however, the decrease in pressure occurs over a smaller height than at higher levels of the atmosphere, where the air is not as compressed (compare the thickness of the layers in Figure 1.5.1(b)). Technically, the rate of pressure decrease
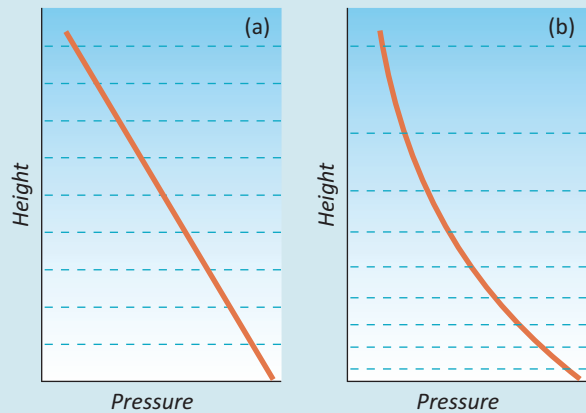

Figure 1.5.1. (a) Linear pressure decrease with height (or linear pressure increase with depth, as would happen in a pool, because water is incompressible). (b) Exponential pressure decrease with height.

with altitude itself decreases approximately linearly with pressure, which results in an *exponential* decrease of pressure with altitude.

1.2.4 Barometers

As the weather changes, air moves around, resulting in atmospheric pressure changes at a given location. We monitor these changes with an instrument that measures atmospheric pressure, which we call a **barometer**.

In the seventeenth century, Evangelista Torricelli, an Italian scientist who was greatly influenced by the writings of Galileo, invented the **mercury barometer**. He filled a glass tube with mercury, inverted it in a cistern, and observed that the level of the liquid in the tube would not completely drop, but would stabilize at a certain height above the cistern (about 760 mm). He concluded that the air pressure applied downward onto the mercury in the cistern was forcing the liquid up into the tube to that level. He correctly speculated that the weight of the column of mercury in the tube was balanced by the weight of the atmosphere pressing down on the mercury outside the tube; i.e., the weight of the mercury column is balanced by atmospheric pressure (Figure 1.8).

Torricelli's design remained the basic principle behind most barometers for more than 300 years. If atmospheric pressure increases, the mercury level rises in the tube. Conversely, when a storm approac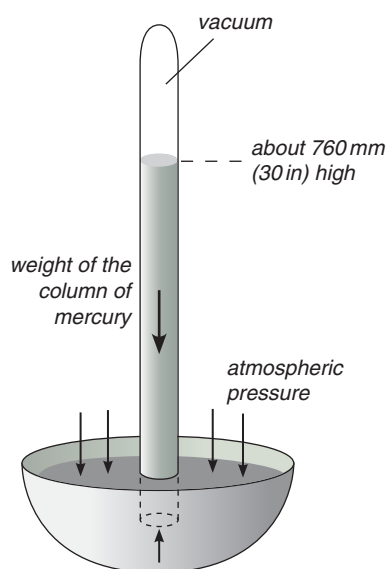hes and atmospheric pressure drops, a shorter column of mercury is required to balance the weight of the atmosphere, and the mercury level drops. That is why atmospheric pressure is often reported as the height of the mercury column, namely about 76 cm, 760 mm, or 30 inches Hg ("Hg" being the chemical symbol for mercury).

Note that, in reality, Torricelli first experimented with water. However, about 10 meters of water are required to balance the weight of the atmosphere, i.e., about 33 feet, which made the experiment quite cumbersome! Since mercury is much denser, and thus heavier than water, a much shorter column of mercury is required to achieve the same weight.

If your barometer at home is relatively small and round, you must be wondering how it can possibly contain 30 inches of mercury in a tube. And it does not, of course. You probably own an **aneroid barometer**, which is made of a small empty chamber (evacuated of some air to create a partial vacuum) that can expand and contract. When atmospheric pressure increases, it squeezes the chamber and forces it to contract. When atmospheric pressure decreases, the chamber expands. By connecting the chamber to a moving needle, the expanding chamber can be turned into a precise instrument (Figure 1.9).
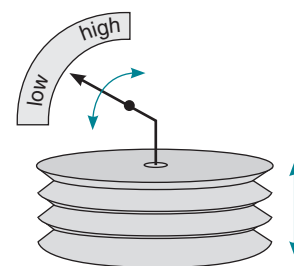

Figure 1.8. Concept of a mercury barometer.

Figure 1.9. Aneroid barometer.