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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 object of this first chapter will be to provide an overview of these variables, a starting point for our exploration of atmospheric changes. 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

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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 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.

locations or transformed between different types of energy. In the atmosphere one form of energy is **heat**, 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**.

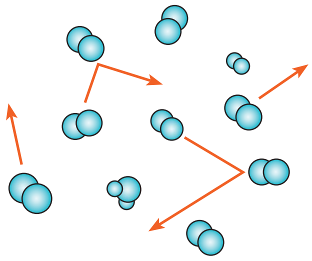

Figure 1.1. Random molecular motion in a volume of air.

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 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 thermometer exist, and all make use of a property of matter to determine temperature indirectly. A bimetallic strip, for 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.

Figure 1.2. Thermometer in a standard weather shelter.

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 2 meters (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:

$$°\mathrm{C} = 5/9 \times (°\mathrm{F} - 32°)$$
$$°\mathrm{F} = 9/5 \times (°\mathrm{C}) + 32°$$

(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.

$$\mathrm{K} = °\mathrm{C} + 273.15°$$
$$°\mathrm{C} = \mathrm{K} - 273.15$$

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:

$$°\mathrm{F} \approx 2 \times °\mathrm{C} + 30$$

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.

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

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 1.3). 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 1.4). These balloons are called **radiosondes**, since the measurements are radioed back to a receiving station at the surface.

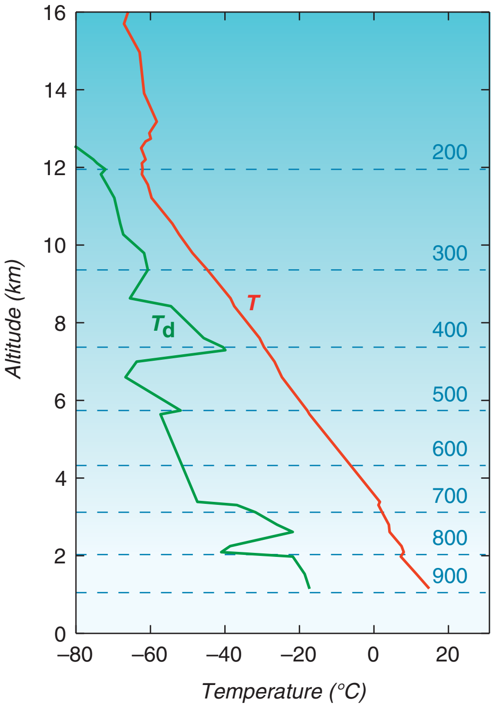

Figure 1.3. Temperature profile obtained by radiosonde at Amarillo, Texas, on February 18, 2014, at 00:00 UTC. The red curve indicates temperature (T) and the green curve indicates the dew point temperature (T_d). Pressure levels (given in hectopascals, hPa) are shown in blue and will be described in the text.

Figure 1.3 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 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 1.3 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 1.3 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.

Figure 1.4. Radiosonde being launched in Hawaii.

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 expressed in terms of pressure (see Box 1.2). 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.

Box 1.2. 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 1.3). 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.

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 1.3). 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:

$$F = p \times A$$

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:

Box 1.3. 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).

$$p = F/A$$

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 1.5).

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

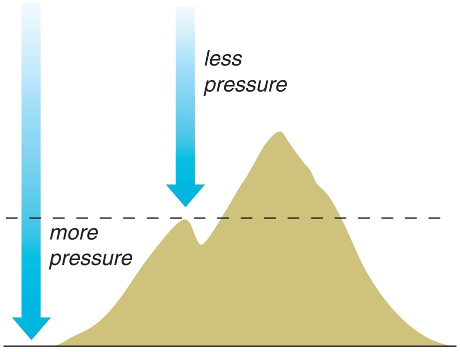

Figure 1.5. Atmospheric pressure is tantamount to the weight of the atmosphere above us.

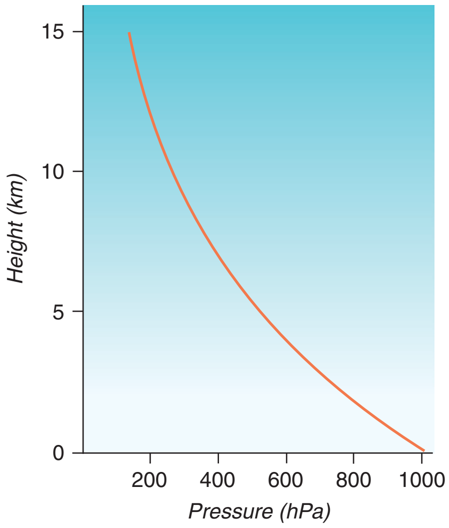

Figure 1.6. Pressure decreases with height.

at ground level, and decreases upward (Figure 1.6 – 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.

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 principle of 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 1.7).

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 approaches 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 1.8).

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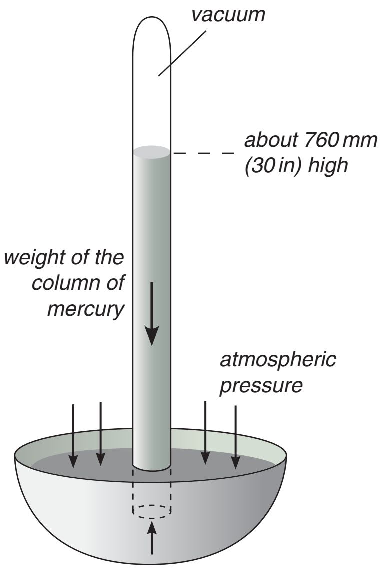

Figure 1.7. Concept of a mercury barometer.

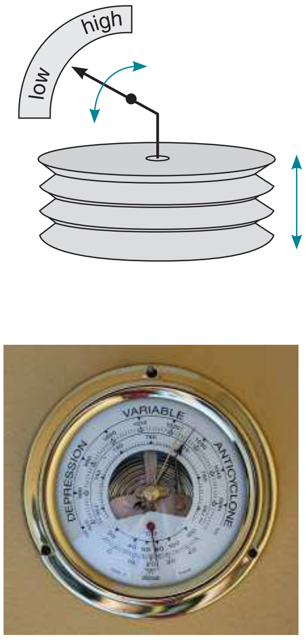

Figure 1.8. Aneroid barometer.

1.2.5 Pressure units

In the International System of Units, pressure is expressed in pascal (Pa), where 1 Pa = 1 N/m^2 (one newton per square meter), or alternatively in kilopascal (1 kPa = 1000 Pa). Note that a newton is a unit of force and a square meter is a unit of surface area, which is consistent with our earlier definition of pressure as force divided by area. In meteorology, however, since we think of atmospheric pressure as the weight of the atmosphere exerted on a surface, it can be more intuitive to express it in pounds per square inch (psi), or kilograms per square centimeter (kg/cm^2) in the metric system. In that context, we can picture a column of air extending from Earth's surface up to the top of the atmosphere. If the base of the column has a surface area of one square inch, then the weight of that air column is on average 14.6 pounds (at sea level); thus, atmospheric pressure is, on average, 14.6 psi.

Meteorologists, however, go one step further. Since we are interested in the pressure exerted by the atmosphere, we may express pressure... in atmospheres! The average atmospheric pressure would then be close to "one atmosphere." For historical reasons, one atmosphere is called "one bar," after the Greek "baros," meaning weight. Since daily variations in pressure are on the order of a thousandth of an atmosphere (i.e., a thousandth of a bar), meteorologists have historically used millibars (mb), where 1 bar = 1000 mb. The exact value of a bar has been chosen as 100 kPa for practical purposes, which is equal to 1000 hectopascal (hPa, where hecto- is the prefix for 100), so millibar and hectopascal can be used interchangeably. Putting it all together, the average pressure at sea level is approximately:

$$1013\,\text{hPa} = 1013\,\text{mb} = 29.92''\,\text{Hg} = 14.6\,\text{psi}$$

To be consistent with the International System of Units, we will use hectopascals (hPa) in the rest of this book.

1.2.6 Some Useful Numbers

If you observe the variations in atmospheric pressure at a given location over time using a barometer, you will find that the typical range of pressure at sea level is about 980 to 1030 hPa. The pressure will only drop below 980 hPa during intense storms. Similarly, it will only rise above 1030 hPa during exceptionally high pressure events.

If you carry your barometer upward, you will find that pressure drops very quickly with height, as we saw in Figure 1.6. As a rough rule of thumb, pressure

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near the surface drops about 8 hPa for every 60 m of altitude – this is easily noticeable, and measurable, in elevator trips in tall buildings. Recall the pool analogy: pressure is greatest at the bottom of the pool, or ocean of air, and decreases as we rise in the pool. Since we will be interested in what happens higher up in the atmosphere, it is useful to have some standard levels of pressure in mind and know their approximate altitude.

Figure 1.9 shows some useful numbers to which we will refer often. We saw that the average pressure at sea level is 1013 hPa. We will often look at maps of weather variables at 850 hPa, which is around 1.5 km altitude (about 5000 ft), and near the top of what we call the "boundary layer" (i.e., near the top of the layer where the effects of the daily variations in temperature near Earth's surface are felt). At 500 hPa, about half of the atmosphere is above us and the other half below (in terms of mass and weight), and this happens at about 5.5 km, or 18,000 ft. Aircraft cruising altitude is at about 250 hPa (roughly 11 km, or 34,000 ft), which is also the top of the **troposphere**. Thus, the troposphere, i.e., the first layer of the atmosphere, where most of our weather takes place, contains about 75% of the mass of the atmosphere. Note that 250 hPa also corresponds to the altitude where we find the jet stream, as will be discussed in Chapter 9. Above the troposphere, we enter the stratosphere, where the ozone layer is found, as we will describe in more detail in Chapter 3.

Since water is 1000 times more dense than air, the increase in pressure with depth happens much faster in the ocean than in the atmosphere. Recall that the weight of the atmosphere is balanced by a 10 m column of water in a water barometer. In other words, a 10 m column of water has the same weight as a column of air extending from sea level to the top of the atmosphere. Therefore, we will add the equivalent of another atmosphere of pressure (i.e., one additional bar) if we dive 10 meters down into the ocean – another good number to remember. And for each additional 10 m layer of ocean, we will add another bar of pressure. That is why it is important for divers to pause regularly, especially on their way up to the surface, to let their blood and other internal fluids adjust to the change in external pressure.

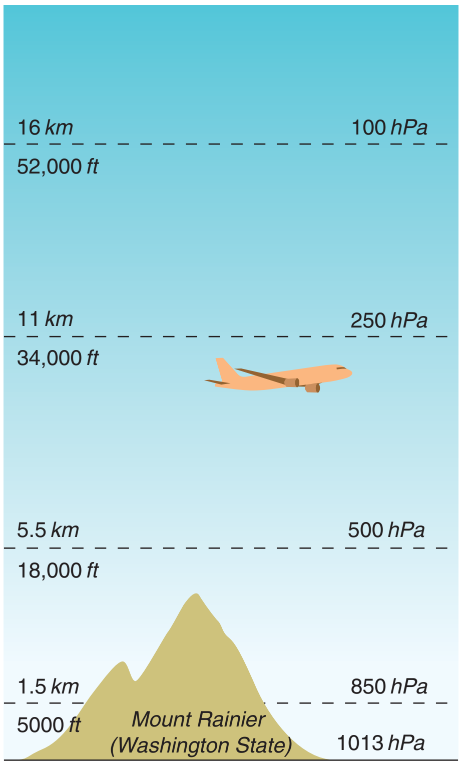

Figure 1.9. Altitude and pressure in the atmosphere: some useful numbers to remember.

1.3 Wind

As mentioned before, weather results from atmospheric motions, and air motion is also constrained by the laws of physics. As for energy and mass, air motion is constrained by a conservation law: the law of conservation of momentum (see Box 1.4). Momentum is the product of mass and velocity. The density of air is estimated indirectly from measurements of pressure and temperature, using the ideal gas law (see Box 1.3). The wind velocity, however, needs to be measured. It has two components, direction and magnitude, which need to be measured separately.

1.3.1 Measuring Wind

Wind direction is measured using a **wind vane**. The vane points toward the direction that the wind is coming from (Figure 1.10), which is why meteorologists report wind direction *as the direction that the wind is*