6

Conceptions and Misconceptions of Pressure

> *"I'm becoming frightened, Doctor," she said, in a low voice. She smiled as she spoke, but I could see that she was seriously worried. "Do you know what the barometer reads? Twenty-eight seventy! And the wind is growing stronger every minute."*

Charles Nordhoff and James Hall surely would not have included this scene in their rattling good tale, *The Hurricane,* had they doubted that a mere barometer reading would convey an apprehension of impending disaster. For who, however rudimentary his knowledge of meteorology, does not associate the rise and fall of barometric pressure with changes in the weather? Indeed, there is no physical quantity the variation of which in time and space is more important to weather forecasting. What, then, is pressure? Although it is true that sea level atmospheric pressure is a measure of the weight of the atmosphere, this is not sufficient to explain all phenomena associated with atmospheric pressure and its variations. Moreover, a fluid's pressure and its weight are two different concepts; they are not synonymous, for if they were then one of them could—and should—be discarded, at no loss to physical understanding and at a gain in economy.

Like temperature, pressure in its everyday sense is a prescientific concept, and you do not have to know what it is to experience it. This is obvious to anyone who dives under water and feels his surroundings squeeze him more tightly the deeper he descends. This experience of pressure is as old as humankind. Yet only comparatively recently did it become evident that the atmosphere, like water, is also a fluid capable of exerting a pressure. This realization was long in coming because we do not ordinarily sense atmospheric pressure. Awareness of atmospheric pressure comes with the ability to measure it, and instruments to do so were not invented until the seventeenth century.

DEMONSTRATING ATMOSPHERIC PRESSURE

There are so many demonstrations of atmospheric pressure that it would be futile to labor mightily on devising a new one or varying somehow an old one. Instead, I shall describe the one I consider to be the best of the lot. It was reported by Pieter Visscher of the University of Alabama in a short note to the *American Journal of Physics* (1979, Vol. 47, p. 1015).

Very little is needed for this demonstration: a hotplate, an empty soft (or hard) drink can, and a shallow pan filled with water. Put a small amount of water (about a spoonful) into the can and place it on the hotplate. The water

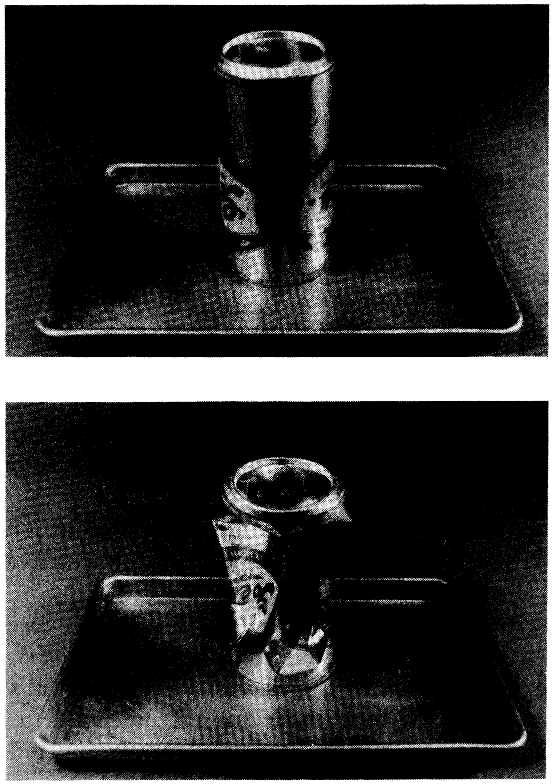

Figure 6.1 When a can filled with hot water vapor is placed in a shallow pan filled with cooler water, the can is crushed because of the pressure difference between its inside and its outside. That outside is atmospheric pressure. That inside is initially atmospheric pressure but drops as the water vapor condenses onto the cooled walls of the can.

temperature will rise quickly and the can will fill with water vapor. Unlike the gaseous constituents of the atmosphere, water vapor is readily condensed (this is the distinction between a gas and a vapor). To do so just snatch the can from the hotplate and plunge it mouth first into the pan. In a few seconds, if all goes well, the can will be crushed (see Fig. 6.1) accompanied by a dull pop. Very dramatic. It sometimes takes more than one try depending on how successful you are at finding something with which to grasp the can.

Tongs of the proper span would be best, but these never seem to be available so I often improvise them using whatever is at hand. As a consequence, I sometimes drop the can or it ends up on its side in the water—a fizzle. But never mind, the demonstration is easily repeated, and even the most fumbling efforts are usually met with success. Be sure to use a *shallow* pan. If you do not, all that is likely to happen is that the can will sink into the water without being crushed.

In Chapter 2 I discussed the concept of saturation vapor pressure—the pressure of a vapor in equilibrium with its liquid—and in Chapter 5 I emphasized how rapidly it increases with increasing temperature. For example, in going from 40° to 100°C (the normal boiling point) the saturation vapor pressure of water increases by more than a factor of ten. It is this large change that crushes the can.

The can sizzling on the hotplate was filled mostly with water vapor at a pressure about that of the surroundings (at the boiling point the saturation vapor pressure is that of the surroundings). When this hot can was plunged suddenly into room-temperature water it cooled, hence the rate of evaporation from its inner surface decreased. The rate of condensation onto it, however, was slower to respond. So there was *net* condensation of water vapor for a moment, and the pressure inside the can consequently dropped to well below that of the surroundings. Thin aluminum cans are not able to withstand such large pressure differences so they collapse. You may wonder why the pressure difference doesn't merely force water into the can, like mercury in a barometer. With a much stouter can—one made of cast iron, for example—this is what would happen. And it would also happen even with a thin aluminum can were it not for the inertia of the water: before the water has sufficient time to flow in, the can collapses.

This demonstration of the *existence* of atmospheric pressure is unambiguous. It is only when we *interpret* what we observe that the ambiguities and contradictions arise.

It is often asserted that atmospheric pressure is just the weight of the atmosphere above a unit area. While this is true in the absence of vertical acceleration, it does not come to grips with the concept of pressure in a fundamental way. For gas pressure and weight are, in general, independent. To convince yourself of this consider the following thought experiment. Fill a sealed container with a gas, which could be air but need not be. At a given location the weight of this gas—that is, the force exerted on it by the gravitational attraction of the earth—is fixed. Yet it is a matter of experience that the gas pressure would rise if the container were heated. And if the gas were taken beyond the pull of the earth's gravitational field it would become weightless but not pressureless. Thus the relationship between pressure and weight is not unique. Moreover, the weight interpretation of pressure does little to help us understand the demonstration. There is more to the pressure concept than that it is an alternative guise for weight.

INTERPRETING PRESSURE

To our coarse senses all matter is *continuous:* we can discern no gaps, no lumps. This is particularly true of the common fluids air and water. Despite this, the hypothesis that matter is ultimately *discrete,* the atomic (or molecular) hypothesis, has had many adherents almost since the beginning of recorded thought. And its appeal is not to be wondered at, for we too are atoms. This does not mean that we are very small: the root meaning of atom, from the Greek *atomos,* is indivisible. Neither we nor atoms can be divided, at least not without losing our salient characteristics: cut a man in two and he just won't be the same.

Although the existence of atoms was hotly debated for centuries, it was only during the first decade of this century that enough evidence had been amassed to silence the critics, among whom were counted some of the most eminent scientists of the day. Today, anyone who seriously denies the existence of atoms and molecules (although not strictly indivisible, a molecule is a more or less stable aggregation of atoms) is classed with the flat-earth believers. But in one sense the critics of the molecular hypothesis were right: it is sometimes neither necessary nor even desirable to take account of the discreteness of matter. For many purposes it is more expedient to proceed *as if* matter were continuous.

Few people have not at least heard of molecules. Unfortunately, what they have heard is often so fragmentary, so superficial, so distorted, that it would have been better had they not. A little bit of knowledge is sometimes less than none at all. For example, if one adopts the continuum interpretation of matter, then one has little choice but to consider the pressure within a gas as resulting from adjacent portions of a continuous medium pushing against one another. If one now switches to the molecular interpretation it is only natural to carry over this notion, that is, to consider gas pressure as resulting from gas molecules pushing against one another. As attractive and widespread as this notion is, it is nevertheless false. The error arises in going from one interpretation to another. Translation errors, sometimes the cause of embarrassment, are not uncommon in going from one language to another. Concepts in one language do not necessarily have exact counterparts in another. So also is it with different languages for describing the physical world. Let us explore how one such language, that of atoms and molecules, describes gas pressure.

MOLECULAR INTERPRETATION OF GAS PRESSURE

Even after one adopts the molecular viewpoint, one is confronted with two aspects of the concept of gas pressure: as a *measurable* quantity and as a *property* of the gas. I shall consider each of these in turn.

Consider a gas in a rigid container. In even a tiny volume of this gas there is an enormous number of molecules: in one cubic millimeter, more than a million billion of them (under normal conditions). Molecules that collide with the walls of the container will suffer a change in *momentum*. Momentum, which is the product of mass and velocity, is a directed quantity. Thus even if the speed of a molecule—how fast, but not where, it is going—does not change upon collision, its momentum does: before collision it was moving *toward* the wall, after collision it is moving *away* from it.

Newton's second law states that a body will not change its momentum unless a force acts on it. And from his third law—to every action there is an equal and opposite reaction—it follows that if the wall exerts a force on a molecule that collides with it, each such molecule exerts an equal and opposite force on the wall (this is painfully obvious to anyone who has ever run headlong into a wall). Because of their great number and speed (at room temperature the average speed of an air molecule is about that of a rifle bullet) molecules collide with the wall at a tremendous rate: in a millionth of a second, each square centimeter of wall is struck by billions upon billions of molecules. And each of them contributes its mite to the pressure, the total force acting on a unit area of the wall. This is the force that causes the liquid in barometers to rise and that caused the can to collapse: the total force due to molecular collisions with the inside of the can was overwhelmed by that due to collisions with the outside.

Note that there is an important distinction between the target (i.e., a surface) and the projectiles (i.e., gas molecules) impinging on it. The average separation between molecules in a gas (at normal temperatures and pressures) is about ten times that in a solid or liquid. The density of solids and liquids is therefore about a thousand times greater than that of gases. As a consequence, the rate at which gas molecules collide with one another cannot be anywhere near the rate at which they collide with a liquid or solid surface. Indeed, to good approximation air may be treated as an *ideal* gas, another name for which is a *collisionless* gas. The very name gives the game away. It is not that collisions between gas molecules do not occur, or even that they are not important in determining *some* properties of gases, it is just that they are not the determinants of such quantities as gas pressure. As far as pressure is concerned, the molecules might just as well be collisionless—with one another, of course. And the same holds true for temperature. Two gases with greatly different rates of intermolecular collision (e.g., two gases with greatly different densities) can be at the same temperature.

When gas pressure is measured, it is always by virtue of molecules interacting with a solid or liquid surface (e.g., a barometer). If we were concerned only with the measurability of pressure then we would need say no more than this. But what meaning is to be attached to pressure construed as a property of the gas? That is, how should we interpret pressure within a gas as opposed to at its interface with a liquid or solid?

The dimensions of pressure are force per unit area, as are those of *momentum flux,* the rate at which momentum is transported across a unit area. This is not just a coincidence. Recall that pressure was associated with the rate at which gas molecules transfer momentum to unit area of a material surface at the *boundary* of a gas. Now imagine a point *within* a gas. Molecules in the neighborhood of this point are whizzing about in all directions, but their total momentum is zero if the gas is at rest: for every molecule with a given momentum there is another with the opposite momentum. Contrary to what you might expect, however, the total momentum flux is not zero. Consider an arbitrary direction, call it the positive direction (the opposite direction is the negative direction). Each molecule has a component of momentum, either positive or negative, along this direction. Moving molecules carry their properties, one of which is momentum, with them wherever they go; that is, molecules transport momentum. The reason that the total momentum flux is not zero, even though the total momentum is, is that positive momentum transported into a region is equivalent to negative momentum transported out. A simple analogy comes to mind: getting rid of a debt (an outflow of negative money) is equivalent to a raise in salary (an inflow of positive money). Lest I mislead anyone, I must emphasize that I have in mind momentum transported across a unit area of an imaginary *flat* surface in a gas. The net rate at which momentum is transported across a *closed* surface in a gas is zero if there are no forces acting on the molecules.

According to the molecular interpretation, the pressure in a gas is the total rate at which its molecules transport their momentum, along a given direction, across a unit area perpendicular to that direction. Regardless of the direction chosen, the corresponding momentum flux is the same. Consequently, the pressure acts equally in all directions. Pressure has essentially nothing to do with the rate at which gas molecules collide with one another. This rate is small compared with the rate at which they collide with a solid or liquid surface, and it is by means of the latter that pressure is measured.

How does the molecular interpretation elucidate observations of atmospheric pressure, such as its decrease with height? This is usually explained by saying that the weight of the atmosphere above a given elevation decreases with increasing elevation. This is true enough, but the observation has an alternative explanation. At each successive elevation lower pressure implies lower momentum flux. This means that for any horizontal layer of atmosphere more vertical momentum goes in at its bottom than comes out at its top. What happened to this momentum? According to Newton's second law, momentum changes (e.g., decreases) only if caused to do so by a force. In this instance, the force is that of gravity. And weight is, after all, merely the force of gravity. So the two interpretations—microscopic and macroscopic—are in harmony. But they can be made to clash unnecessarily by mixing concepts appropriate to one with those appropriate to the other. Consistency is what is wanted. Once you have adopted a particular viewpoint, stick with it. If you must change, beware of translation errors.

The choice of interpretation is to some extent a matter of taste. More often it is guided by considerations of expediency. It would be foolish, for example, to try to do a molecule by molecule forecast of the weather. For this purpose the view that matter is continuous is the most sensible one. For other purposes, however, understanding is acquired only by recognizing the ultimate discreteness of matter.

7

Dew Drops on a Bathroom Mirror

> *Nothing . . . is absolutely true. The earth is not quite round. The sky is not quite blue. Rain isn't altogether wet.*
> Stephen Leacock

Several years ago, after class one day, a student brought me a small vial similar to a purse-size perfume atomizer; its label proclaimed that it contained a wondrous compound which when applied to eyeglasses would keep them from fogging. He wanted to know how it worked and whether it was worth its steep price. I told him that I would show him how to obtain a lifetime supply of antifogging compound for a few dollars.

When class next met I had with me two Petri dishes and a container of liquid dishwashing detergent. First I breathed hot and heavily on one dish, the way you do when you want to warm your hands on a cold day, and immediately put it on an overhead projector to show that it had been fogged by my breath. Then I smeared a very thin film of liquid detergent onto the other dish. A fraction of a tiny droplet is more than enough. This dish, with its imperceptible film of detergent, I could not fog with my breath, no matter how hard I tried. You might think that the film prevented water from condensing onto the dish—a natural, but incorrect, supposition. To convince yourself of this you can do a simple (and relaxing) experiment in your own home.

TAKE A HOT BATH

This experiment is done in the bathroom, a home laboratory second in importance only to the kitchen. With liquid soap or detergent make a figure on the dry bathroom mirror; the film need be only very thin. Then hop into the tub

and take a good hot bath, preferably on a cold day. Take your time. Enjoy your soak. And as you blissfully stew in your own juices cast a glance now and again at the mirror.

The moist air in the bathroom is much warmer than the dry mirror. From the mirror's point of view, its environment is supersaturated (see Chapter 5). Water will therefore gradually condense onto it, as evidenced by fogging—except where it is covered by the soap film. To convince yourself that the spot which appears to be free of water is not, press small pieces of tissue paper against clear and fogged spots. Paper will stick to both of them (shown in Fig. 7.1), but not to the mirror when it is dry. So the soap film has not prevented water from condensing.

The fogged part of the mirror is covered with small water droplets, whereas the clear part is covered with a more or less continuous film of water. Scattering has not been prevented, only its nature has been changed: the water droplets scatter light in all directions; the water film scatters light mostly in a few directions (for more on this see Chapters 13 and 14).

There is a way to prevent (net) condensation, which I had to use in order to take the photograph in Figure 7.1. On my first attempt I got out of the tub, took my camera and tripod from a cold bedroom, set them up outside the bathroom, opened the door poised to shoot—and couldn't see a thing. My lens had fogged. Of course, I should have anticipated this, but perhaps my critical faculties had been softened by the warm waters of my bath.

To prevent the lens from fogging, I took off the haze filter and heated it by suspending it on a roasting rack over a lamp. Water vapor is still incident on

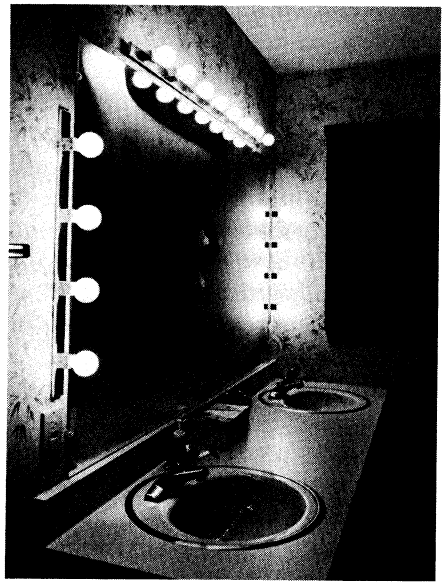

Figure 7.1 This mirror is fogged everywhere except in a spot over which a thin soap film had been smeared. That the film does not prevent water from condensing onto the mirror is evident from the two pieces of tissue paper sticking to the wet mirror, both where it is fogged and where it is clear. Paper will not stick to the mirror when it is dry.