11. The Molecular Model of Matter

Earlier we discussed some of the properties exhibited by matter. No attempt was made to explain these properties in terms of more fundamental structure; our purpose was to outline some of the interesting features that have been observed and are important for scientific purposes.

We will now begin to develop more detailed models by which these and other features of matter can be understood. The word model in this context means an image of how we imagine matter to be organized. We have already encountered one such model, the Continuous Model, which is useful and adequate for many purposes. However, the Continuous Model fails to explain many things, so we must make other assumptions about the composition of matter. These will be accepted to the extent that they can successfully explain the behavior of matter in ever more sophisticated experiments.

Despite the many useful models science has created, the puzzle of the nature of matter has not yet been solved completely—no one knows what the ultimate building blocks of nature are like. The problem is like a collection of nested eggs: when each egg is opened, another smaller one is found inside. As the eggs become smaller, they are more difficult to open. Each opening provides greater insights, but also raises new questions. Someday scientists may come to the last egg, one that will prove to be the basis for everything, but so far we have only hints about what the ultimate nature of matter is like.

Thus, each successive model of matter provides new insights and opens new possibilities. These insights and possibilities simultaneously demonstrate the correctness of our understanding and provide useful new ways to organize and use the materials found in nature. The models appear to be valid and useful, even though they are not as complete as we might wish.

Molecules

In the Molecular Model, matter is assumed to be composed of tiny particles called molecules. Each kind of matter has a different kind of molecule—one for hydrogen, another for water, still another for sugar, and so on (Fig. 11.1). Molecules are so small that they cannot be seen even with the most powerful optical microscopes, although some of the larger molecules have been “seen” in recent years with electron microscopes (Color Plates 1 and 2). Indirect measurements, however, have given us considerable information about the size, shape, and internal makeup of many molecules.

The molecules in matter are in constant motion, interacting with each other and their surroundings through electromagnetic interaction. Because of these motions, the theoretical framework of the Molecular Model of Matter is often called the Kinetic Theory of Matter, and it is used to explain large-scale properties of matter.

Molecules move and interact, in accord with the laws of motion and the laws of force as well as the conservation principles. The great success of the Molecular Model lies in its ability to explain the properties of matter in terms of physical laws, which are already understood and which can be tested using larger chunks of matter.

Brownian Motion

When small objects in fluids—such as dust particles in water—are observed through an optical microscope, they appear to be in constant, irregular motion in random directions (Fig. 11.2). This is called Brownian Motion, in honor of Robert Brown, who first described it in 1827.

![Hydrogen](image1.png)
![Water](image2.png)
![Methane](image3.png)

Figure 11.1. Representations of the shapes of some simple molecules.
1.2. Dust particles seen through a microscope are in constant, irregular motion. Why? At first, Brownian Motion seems mysterious—the particles do not appear to touch, yet they are obviously interacting with something. The motions are typical of collision-type contact interactions. The motion must be a result of the interaction between the particles and the fluid that surrounds them.

The Molecular Model provides an obvious explanation. The visible particles are being bombarded by the “invisible” molecules of the fluid in a random way. The motion of the visible particles reveals the presence of the molecules.

Although this phenomenon had been known since early in the 19th century, only in 1905 did Albert Einstein properly explain it in terms of the Molecular Model. Einstein’s successful explanation of Brownian Motion is one of the important proofs of the validity of the Molecular Model and the existence of molecules.

The States of Matter

The Molecular Model suggests a straightforward explanation for the existence of the various states of matter. The molecules in solids must be attached rigidly to one another. The attractive forces are due to the electrical interaction. Further, the arrangement of charged particles within the molecules determines the physical arrangement of the molecules in solids (Fig. 11.3).

Molecules in liquids are still in contact with each other, but they are no longer held rigidly in place. Instead, they are free to roll around each other and to move from place to place. The arrangement of molecules in a liquid state is much like that of many small ball bearings in an open box. These can be poured from one box to another and assume the shape of their new container. They exert forces on each other and on the sides of the container and are free to move about if they are agitated, perhaps by someone shaking the box. Many of the properties of liquids can be understood by imagining the behavior of such small balls. But unlike the ball bearings, molecules are attracted to each other by the same forces that would hold them together in the solid state.

In gases, the molecules are usually some distance from each other. They move through space almost as if they were free particles. They may collide with each other and with the walls of their container, but they spend most of the time free from outside influences other than gravity. Whereas the molecules in solids and liquids fill the volume of the material almost completely, the molecules in a gas may fill only one-tenth of one percent of the total volume—most of the volume is just empty space. This is why and how gases can expand to fill any container.

The Molecular Model also suggests a simple explanation for the changes of state that occur in all materials. The key is to remember that changes of state are always associated with changes in energy. Energy must be added to a solid if it is to become a gas. Imagine the molecules in a solid, perhaps with little energy. The molecules are almost at rest, vibrating rather slowly with respect to one another. As the molecules receive energy, they vibrate more rapidly than before, even though the forces holding the molecules together may be strong enough to keep them in position while they are moving. If the molecules move enough (i.e., if they receive enough energy), the attractive forces between them will not be able to hold them near fixed positions. When this happens, the solid melts and becomes a liquid. The molecules do not have enough energy to separate from one another, but they do have enough energy to...
to break the rigid bonds associated with fixed arrangements. If the molecules in the liquid continue to receive energy, their kinetic energy and their speed continue to increase. If they move fast enough, they finally can escape completely from the attractions of their neighbors. They then leave the liquid and travel as free particles until they encounter the walls of the container. In this gaseous state the molecules are traveling quite fast—faster than the speed of sound (about 720 miles per hour).

The Molecular Model of a liquid can also explain evaporation at temperatures lower than boiling (Fig. 11.4). The molecules are moving about, sliding over one another and colliding with one another in random ways. As a result of these collisions, the molecules do not have the same speed. Sometimes one will be bumped in such a way to give it a rather high speed, while another will have a much lower speed. The total energy stays the same, but some molecules have more energy than others.

![Figure 11.4](image)

Figure 11.4. How can molecules escape from a liquid when it is not hot enough to boil?

Imagine what would happen if one of these high-speed molecules were near the surface of the liquid, particularly if it were moving away from the bulk of the molecules the liquid is composed of. If the molecule were going fast enough, it would escape from the attractions that hold it to its neighbors and keep going, leaving the liquid surface and entering the space above it. This is exactly what happens when a liquid, such as water, evaporates. The faster moving molecules escape from the liquid surface.

The molecules of a gas frequently collide with each other. At ordinary energies, these collisions are elastic. The molecules bounce off each other just as if they were little, hard rubber balls. If they are moving fast enough, however, these collisions can actually break the molecules apart, sometimes separating the charged particles they are composed of. When this happens, the gas becomes a plasma and begins to emit light as the charged particles again combine into neutral combinations.

Finally, the Molecular Model suggests why some materials are easier to melt or vaporize than others. Molecules that are strongly attracted to each other would need considerable energy to escape that attraction, whereas those that are weakly attached would be separated with less energy. The model is not sophisticated enough to help predict in advance which molecules would have the stronger attractive forces, but it does explain that melting and boiling temperatures are related to the strength of the attraction of molecules for each other.

**Internal Energy and Temperature**

We previously identified one of the important forms of energy as internal energy. It was given this name because the energy seemed to be somehow locked up inside matter. The Molecular Model can explain internal energy in terms of more familiar kinds of energy.

Our first observation is that molecules must be moving. Their motion is associated with kinetic energy, just as is the motion of a moving baseball or a moving truck. The molecules are so tiny, however, that each one has only a short distance to travel before it experiences a collision and changes direction. Any collection of molecules large enough to see contains trillions of billions of molecules, all moving in different directions and each having a small amount of kinetic energy. For example, the molecules in a table are vibrating violently. Each is traveling faster than any automobile or most airplanes. Yet the table seems to be at rest, with almost no manifestation of its internal turmoil. This internal kinetic energy represents an important part of the internal energy of any material. Internal kinetic energy is usually called **heat**, or **thermal energy**, for reasons that should become clearer as we proceed.

The second form of internal energy is electrical potential energy associated with the interaction between molecules. We know from our molecular understanding of melting and boiling that molecules attract each other. Remember that attracting objects have more potential energy when they are farther apart. The same is true of molecules. The potential energy associated with the attractive forces between molecules is another important part of internal energy.

Additional internal energy is associated with the motions and forces within the molecules themselves. We are not quite ready to discuss the details at this point, but this is the energy involved in all chemical reactions.

Now we can see that internal energy is really not a unique form of energy at all; it is just the same kind of kinetic energy and electrical potential energy that we have previously identified. The only difference is that internal energy occurs on a molecular scale, each molecule making a small contribution to the total energy of
the whole. Certainly, we would not think of trying to measure the kinetic energy of each molecule and then adding all of these to calculate the total energy in any sample of matter. It is much easier to measure the large-scale manifestations of internal energy, such as temperature and physical state. Nevertheless, in principle at least, internal energy is the same as the ordinary kinds of energy we deal with when we study the motions of larger objects.

With this insight, we can complete Figure 7.10, which represents the various kinds of energy. The result is Figure 11.5. Potential energy is associated with all the forces of nature, but there are no significant microscopic manifestations of gravitational potential energy and no macroscopic examples of nuclear potential energy. With those exceptions, these kinds of energy may be summarized by saying that there are microscopic and macroscopic forms of kinetic energy, three kinds of potential energy, and rest-mass energy. The figure lists some common examples but is not exhaustive.

We have previously noted the connection between temperature and internal energy. Generally, higher temperature is associated with higher internal energy. We now see that internal energy has several components, each representing a particular type of microscopic energy. Molecular kinetic energy is associated with temperature. Temperature is a measure of the average molecular kinetic energy of any collection of molecules. (The average is used because the molecules in any sample of matter do not all have the same speed and, therefore, do not all have the same kinetic energy.)

These ideas can be combined to help understand a fairly common phenomenon: cooling through evaporation. Wet one finger and hold it in the air. As the water evaporates, your finger feels cooler. This occurs because the faster molecules in the water are the ones that escape; the slower ones with less kinetic energy are left behind. The average kinetic energy of the remaining molecules is less than the average kinetic energy before the faster ones left. Thus, the temperature of the water is reduced as evaporation takes place. Thermal energy decreases, but what form of energy increases to conserve energy? The electrical potential energy of the molecules that evaporate is increased as they are separated from one another. This increase in electrical potential energy is shown in Figure 11.6 as a change in physical state from liquid to gas.

Figure 11.6. How is the internal energy of water vapor different from that of ice?

Heat Conduction

We have previously described heat conduction as the process by which internal energy is transferred from

![Figure 11.5. The forms of mass-energy taking the details of internal energy into account.](image-url)
one object to another because of differences in temperature; two objects of different temperature eventually will come to the same temperature if allowed to remain in contact long enough. The Molecular Model explains what is happening and why.

The molecules of the hotter object have more kinetic energy on the average than those of the colder object. At the points of contact, the molecules of the one are colliding with those of the other. In those collisions, energy is transferred from molecules with more energy to molecules with less energy. The speeds of the molecules in the cooler object are increased, and the temperature of the object rises. The molecules in the warmer object, having lost some of their energy, are slower than before. Since the average kinetic energy of its molecules is reduced, its temperature drops. The process continues until the average kinetic energy of the molecules in the two objects is the same. At this point collisions still occur at the contact points, but no average transmission of energy from one object to the other occurs. Thus, heat conduction is just the transfer of molecular kinetic energy from molecules with more energy to molecules with less energy.

To visualize these ideas, consider the heat conduction through a glass window on a cold day (Fig. 11.7). The outside air and the glass window both have lower temperatures than the inside air. The inside air molecules, with more kinetic energy, collide with the surface layers of molecules in the glass. These experience an increase in kinetic energy that is transmitted, by collisions between molecules within the glass, through the entire layer of glass. Collisions occur between the energetic molecules near the outside surface of glass and the less energetic air molecules adjacent to that surface. These collisions transmit kinetic energy to the outside air.

The average kinetic energy of molecules, not their speed, comes to equilibrium in such situations. The glass molecules in this example have considerably more mass than do the air molecules. When they reach the same temperature, the air molecules are moving much faster than the glass molecules because of the dependence of kinetic energy on mass as well as speed. The less massive air molecules simply must move faster to have the same kinetic energy. It is this equality of average kinetic energy at thermal equilibrium that leads us to associate temperature with kinetic energy, not with speed or some other measure of the molecular motion.

Properties of Gases

Matter in the gaseous state exhibits several properties that can be understood using the Molecular Model. The molecules in gases are far enough apart that their mutual attractive and repulsive forces are unimportant, except during the brief intervals while they are colliding with each other and with the walls of their container. Thus, gas molecules move almost as free particles between collisions. Further, as long as the temperature is low enough so no permanent damage is done, the details of the collisions are not important when considering the external behavior of the gas.

Molecular kinetic energy is the most important form of internal energy in the physical behavior of gases. As temperature increases, the internal energy increases in a predictable way. These predictions have been verified by measuring the masses and speeds of a variety of gas molecules at different temperatures.

Since the speeds of molecules at various temperatures can be measured and predicted, one can estimate a temperature at which they would have no speed at all—at which they would be completely at rest. This is called the absolute zero of temperature and is found to be just below -273 °C (-460 °F) for all gases. Absolute zero represents a lower limit of temperature at which all materials have the least possible internal energy.

The forces that gases exert on their surroundings can also be understood in terms of the Molecular Model. Such forces are often described in terms of pressure, or force per unit of area. For example, the air above a table pushes down on the table with a force of about 15 psi (pounds per square inch), or about 1 ton per square foot. The air in an automobile tire pushes outward with an additional force of about 30 psi, thus supporting the car and preventing the tire rim from falling to the road surface.

Gas pressure is caused by the collisions of the gas molecules with the surface experiencing the force. Suppose you were to throw a tennis ball against a wall. During the collision there would be a small force exerted on the wall and the ball that would last only a short time and probably would have no appreciable effect.
Suppose, however, that you had some mechanism for throwing thousands of tennis balls at the wall, picking up each and rethrowing it after rebounds. The force from all the balls would add together to provide a continuous and considerably strong push on the wall. This is the way that gases exert force—through the combined effect of billions of billions of small collisions (Fig. 11.8).

Figure 11.8. How do the hot gases in an automobile cylinder exert pressure on the piston?

With these ideas you can understand another common occurrence. Gas pressure increases with temperature. For example, if you measure the pressure in your car tires when they are cold and after traveling several miles so they are hot, you will find the pressure to be significantly higher in the hot tires. In terms of the Molecular Model, the hotter molecules move faster and collide harder with the tire walls than do the slower molecules in cooler air (Fig. 11.9).

It is possible, using the Molecular Model and the laws of motion, to predict the pressure that would be exerted by any sample of gas molecules on the walls of its container. The prediction is, as we would expect, that the pressure should increase with temperature.

Another prediction of the model is that equal numbers of molecules in equal volumes at the same temperature would exert the same pressure no matter what the mass of the individual molecules might be. (This result, known as Avogadro’s Hypothesis, was first suggested in 1811 by Amadeo Avogadro to explain some of the details of chemical reactions.) Suppose, for example, that the molecules of one gas have 100 times the mass of another and that both gases have the same number of molecules confined inside containers that have the same volume. As long as the temperatures are equal, the pressures exerted by the two gases on the walls of the containers are predicted by this theory to be the same.

At first this result seems a bit surprising. Anyone knows that, other things being equal, a collision with a more massive object involves more force. We have to remember, however, that all other things are not equal between these two gases. The two sets of molecules have the same temperature and, therefore, the same average kinetic energy. That means that the ones with more mass are moving slower—in this case, about one-tenth the speed of the lighter molecules. Each collision of a heavy molecule exerts ten times as much force as does each collision of a lighter one.

We can understand the theoretical prediction that the pressures are the same by noticing that the smaller molecules, moving ten times as fast, strike the container ten times as often. These factors taken together allow us to predict that the total forces exerted by the two sets of molecules on the walls of their containers have exactly the same strength.

**Change of Physical State**

When matter changes from one physical state to another, there must be a systematic reorganization of the molecules within the matter. As a result, there are accompanying changes in the internal energy of the matter. For example, if we want to change the physical state of water from solid to liquid, we must add internal energy to the ice. However, as matter changes physical state, a peculiar thing happens.

Imagine that we have created a mechanism to remove internal energy from water at a slow but constant rate. If we begin with water at room temperature, we notice that the temperature falls as we remove internal energy (see Fig. 11.10). The falling temperature tells us that at least some of the internal energy that is being
removed comes at the expense of internal kinetic energy of the molecules because the decreasing temperature is a measure of the average kinetic energy of the molecules. However, when we reach the freezing point of water, the temperature no longer changes even though we continue to remove internal energy at the same rate from the water! This means that the internal energy being removed at this point must come at the expense of the electrical potential energy of the molecules as they reorganize themselves from liquid to solid state.

The amount of internal electrical potential energy that we remove during this change of state at constant temperature is called \textit{latent ("hidden") heat}. When the water changes from gas to liquid or from liquid to solid, the molecules must “dump” the latent heat to the surrounding environment. On the other hand, when the water changes from solid to liquid or from liquid to gas at constant temperature, an amount of energy equal to the latent heat must be \textit{added} to the internal electrical potential energy of the water molecules. The change of physical state from liquid to gas, which occurs at the constant boiling point temperature, is not to be confused with evaporation which may occur in principle at any temperature and, in fact, is often associated with a cooling effect which changes the temperature.

When water droplets freeze to form snowflakes in a winter storm, the water molecules release the latent heat to the surrounding molecules of air (nitrogen, oxygen, carbon dioxide). The air molecules experience an increase in kinetic energy, and a corresponding warming trend occurs. (However, the warming trend may be masked by movements of warm and cold air which usually accompany the storm.) Similarly, the release of latent heat at the earth’s inner core as molten iron solidifies to form the inner core provides energy to keep the outer core hot enough to remain in the liquid state.

It is estimated that about one-sixth of the solar energy reaching the earth is stored as internal energy of water vapor (gas) in the atmosphere. When the water condenses to form raindrops (liquid), the latent heat must be released. This enormous amount of energy is collected over the oceans but dumped over the land as

Figure 11.10. Temperature changes that accompany the removal of internal energy from water. Observe the horizontal portions of the graph where the temperature does not change. Is the average kinetic energy of the molecules changing along these horizontal portions of the graph? What kind of internal energy is being changed during changes of physical state at constant temperature?
the water condenses, releasing energy to drive the monsoon weather patterns that provide moisture and sustenance to nearly half the people of the earth who live in India, Asia, and parts of Africa.

Summary

The Molecular Model of matter is rather simple. There are no sophisticated assumptions about the structure of molecules, nor about their interactions with each other. Even so, the model successfully explains such widespread phenomena as changes of state, changes in internal energy, the relationship between internal energy and temperature, heat conduction, gas pressure, and Brownian Motion. We have not shown the mathematical details, but the quantitative agreement between the predictions of the model and experimental observations is impressive indeed. No other model of matter can explain such a wide range of phenomena with such precision. The Molecular Model by itself, however, does not explain all the details of chemical reactions. For these, it is necessary to have more information about the structure of molecules themselves. We will see that chemical reactions involve changes in these structures.

Historical Perspectives

The idea of an “atom” can be traced to the Greek philosophers Democritus and Leucippus, who lived about 450 B.C. Atoms were conceived to address the question of what it was about the world that was eternal. Atoms (as the word itself means) were “uncut” and had no parts, and hence were not subject to deterioration, change, or internal rearrangement. Atoms were eternal things in a world of change.

Atoms were also very tiny. So for well over 2000 years they persisted as an idea, but always with the skepticism aroused by something that cannot be seen. Our modern concept of atoms and molecules began to come into focus as the result of the study of the behavior of gases. How is it that gases exert pressure? If you compress the gas, why does its temperature rise? If the compression is done while holding the temperature constant, why does the pressure increase?

Robert Boyle (1627-1691) was the seventh son and fourteenth child of the Earl of Cork, Richard Boyle. As a child he learned Latin and French. At eight years of age he was sent to school at Eton, and at age 11 he traveled abroad with a French tutor. Visiting Italy in 1641, he studied “the paradoxes of the great stargazer,” Galileo. Besides being a prolific natural philosopher, Boyle was interested in theology. He learned Hebrew, Greek, and Syriac to pursue his scriptural studies. At his death he endowed in his will the Boyle lectures for proving the Christian religion against the infidels—Jews, Moslems, atheists, and pagans. He added the proviso that controversies between Christians were not to be mentioned in the lectures.

By 1662 Boyle had discovered that, at constant temperature, the product of the pressure times volume of a gas is a constant. Decrease the volume and the pressure increases by a prescribed, predictable amount and vice versa. He offered two possible explanations in terms of atoms, though neither was original with him.

First, the air might consist of compressible particles like little tufts of wool that were at rest and touching one another. When compressed, the little tufts would act like springs and exert repulsive forces. Heating the gas corresponded to adding a substance called “caloric,” which flowed in to surround the atoms. The caloric caused an increase in the intensity of the repulsion and caused the gas to expand or to exert more pressure. It was a little harder in this model to account for the ability of the atoms to expand to fill whatever space they are allowed, as gases are observed to do.

Boyle’s alternative model (later developed by Daniel Bernoulli in 1738) was that atoms were not continually touching, but rather were suspended in a fluid and were in violent motion, exerting pressure as a result of collisions with one another and with the walls of the container. But it seemed difficult to imagine getting Boyle’s simple relationship between pressure and volume from chaotically moving particles.

The concept of atoms was simplistic until John Dalton (1766-1844)—a poor, largely self-taught Englishman—rose to fame for his basic studies in what was to become the science of chemistry. Although Dalton had an erroneous conceptual framework and was misled by his own sloppy experiments, he proposed the idea first that there were to be atoms of different kind and size for each pure substance. Dalton also admitted the combination of two or more atoms to form compound atoms (molecules), which were spherical structures with the centers of the constituent atoms in close proximity and surrounded by a sphere of caloric. Atoms (or molecules) of a given substance were taken to be identical.

It wasn’t until the generalized Law of Conservation of Energy and the equivalence of mechanical work and heat had been established in the period following 1842 that people returned to Boyle’s second model (as it had later been elaborated by Daniel Bernoulli). With the notion of caloric in disrepute, James Joule turned to the kinetic (motion) theory of heat and began to give some of the qualitative explanations for phenomena in terms of moving molecules, much as we have done in this chapter. Starting in 1857, Rudolf Clausius (Germany), James Clerk Maxwell (Britain), and Ludwig Boltzmann (Austria) provided the quantitative tests of the theory. In this work the molecules of gases were finally put into rapid and chaotic motion, moving in straight lines at constant speed until they struck one another or the walls of the container.
Despite the success of this work, the molecules were still invisible, and some chemists (notably early Nobel Laureate Wilhelm Ostwald) remained unconvinced of their reality until Einstein’s quantitative theory of Brownian Motion in 1905.


STUDY GUIDE
Chapter 11: The Molecular Model of Matter

A. FUNDAMENTAL PRINCIPLES: No new fundamental principles. Chapters 3-8 discuss the laws of force, motion, and conservation that govern the behavior of molecules.

B. MODELS, IDEAS, QUESTIONS, OR APPLICATIONS
1. The Molecular Model of Matter: The essential defining characteristics of the Molecular Model are: (a) Matter consists of tiny particles called molecules. (b) Each different kind of matter consists of a different kind of molecule. (c) The molecules in matter are in constant motion. (d) Molecules move and interact in accord with the laws of motion, the laws of force and the laws of conservation applicable to motion.
2. How is Brownian Motion explained?
3. How are the different states of matter and changes between states explained?
4. What is temperature and how is it related to internal energy?
5. How can heat conduction be explained?
6. How is the pressure in a fluid explained?

C. GLOSSARY
1. Brownian Motion: The constant, irregular motion of very fine particles (such as fine dust or smoke) suspended in a fluid and observed with a microscope. Brownian Motion is taken as evidence for molecules, which collide with the observed particles and cause the jittery motion.
2. Evaporation: The process by which a liquid loses molecules from its surface and thus changes (at least, in part) from the liquid physical state to the gaseous physical state.
3. Heat Conduction: The transfer by physical contact of molecular kinetic energy from a material consisting of molecules with more average energy to a material consisting of molecules with less average molecular kinetic energy.
4. Internal Energy: The kinetic and electrical potential energy of the molecules within a sample of material.
5. Latent Heat: Internal electrical potential energy that must be added to or removed from matter as molecules undergo systematic reorganization in a change of physical state at constant temperature. Latent heat must be removed to change matter from gas to liquid or from liquid to solid. An amount of internal energy equivalent to the corresponding latent heat must be added to change from solid to liquid or liquid to gas.
6. Molecules: The tiny constituent particles of which matter is composed. An atom may be a molecule, but not all molecules are atoms. Within the Molecular Model of Matter, however, the size, shape, and structure of molecules are not specified.
7. Pressure: Force per unit area. The pressure exerted by fluids, in particular, arises from the force exerted by the collision of their molecules with a surface or with each other.
8. Temperature: A measure of the average kinetic energy of molecules in a material.
9. Thermal Energy (or heat): That portion of internal energy which is associated with the kinetic energy of molecules.

D. FOCUS QUESTIONS
1. For each of the following:
   a. Outline the main elements of the Molecular Model of Matter.
   b. Describe the experiment or the phenomenon in terms of the Molecular Model. If an experiment is involved, describe it and the observed results.
      (1) Brownian Motion.
      (2) A wet finger is held up into the wind. The finger feels cold.
      (3) A glass thermometer is placed into a cold pan of water. The water is at a lower temperature than the thermometer.
      (4) A cold object is placed in contact with a hot object.
      (5) Pressure exerted by a gas.

E. EXERCISES
11.1. What is actually “seen” when Brownian Motion is observed?
11.2. How would you expect Brownian Motion to change if the temperature were increased?
11.3. What is meant by “Brownian Motion”?
11.4. How is Brownian Motion explained by the Molecular Model?
11.5. Using the Molecular Model, explain why gases readily change volume when pressure is applied while liquids and solids do not change volume appre-
11.6. Using the Molecular Model, explain why solid materials resist changes in shape, while liquids readily assume the shape of their container.

11.7. Using the Molecular Model, explain why evaporation takes place more rapidly at high temperature.

11.8. What is meant by the Molecular Model of Matter?

11.9. What is meant by the Kinetic Theory of Matter?

11.10. Explain how the different properties of solids, liquids, and gases are accounted for by the Molecular Model of Matter.

11.11. Using the Molecular Model of Matter, describe what happens when a solid melts.

11.12. In terms of the Molecular Model, describe how evaporation takes place.

11.13. Two gases have the same temperature, but the molecules of one have more mass than the molecules of the other. In which gas are the molecules traveling faster?

11.14. The temperature of boiling water does not rise above 100 °C until all the water has evaporated, even though considerable energy is added to the water during this time. Where does the energy go?

11.15. The temperature outdoors often rises just before it starts to snow. Why?

11.16. Describe the energy changes that occur when an ice cube melts in a glass of warm water.

11.17. How is the internal energy of water vapor different from that of ice?

11.18. Using the Molecular Model, describe how the temperature of the mercury in a mercury-in-glass thermometer is increased when the thermometer is placed in a hot liquid.

11.19. A common practice for insulating windows is to use two pieces of glass with a layer of air between them. Using the Molecular Model, explain why less heat conduction would occur for this arrangement than if the two pieces of glass were touching.

11.20. Describe how heat conduction occurs as explained by the Molecular Model.

11.21. Exactly what is the “internal energy” of a gas? What is different about the motion of the molecules in a gas if the internal energy of the gas is increased?

11.22. The pressure of all gases rises when their temperature is increased. For example, the air pressure in the tires of a car is higher when the tires are hot than when they are cold. Using the Molecular Model, explain why this is so.

11.23. What is meant by the “temperature” of a gas? What is different about the motion of the molecules in a gas if the temperature of the gas is increased?

11.24. Using the Molecular Model of gases, explain how a gas exerts pressure upon its surroundings.

11.25. What is meant by the “absolute zero of temperature”?

11.26. Two different gas samples, A and B, are contained in identical containers at the same temperature. Each has the same number of molecules, but each molecule of A has 64 times as much mass as each molecule of B.

(a) In which sample are the molecules moving faster?
(b) In which sample do the individual molecules exert more force as they strike the walls of the container?
(c) In which sample are there more collisions between molecules and walls?
(d) In which sample is the pressure larger?