9.13: Kinetic Theory of Gases- Postulates of the Kinetic Theory

In other sections it was mentioned that many of the properties of solids, liquids, and gases could be accounted for if we assumed that substances are made of atoms or molecules which are constantly in motion. Boyle's law and the other gas laws have now given us much more quantitative information about gases, and it is worth asking whether with the previous model we can make quantitative predictions in agreement with these laws. In answering this question, we will also gain important insights into the nature of temperature and of heat energy.

The microscopic theory of gas behavior based on molecular motion is called the **kinetic theory of gases**. Its basic postulates are listed in below:

> ⚙ Postulates of the Kinetic Theory of Gases.
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> 1. The molecules in a gas are small and very far apart. Most of the volume which a gas occupies is empty space.
> 2. Gas molecules are in constant random motion. Just as many molecules are moving in one direction as in any other.
> 3. Molecules can collide with each other and with the walls of the container. Collisions with the walls account for the pressure of the gas.
> 4. When collisions occur, the molecules lose no kinetic energy; that is, the collisions are said to be **perfectly elastic**. The total kinetic energy of all the molecules remains constant unless there is some outside interference.
> 5. The molecules exert no attractive or repulsive forces on one another except during the process of collision. Between collisions, they move in straight lines.
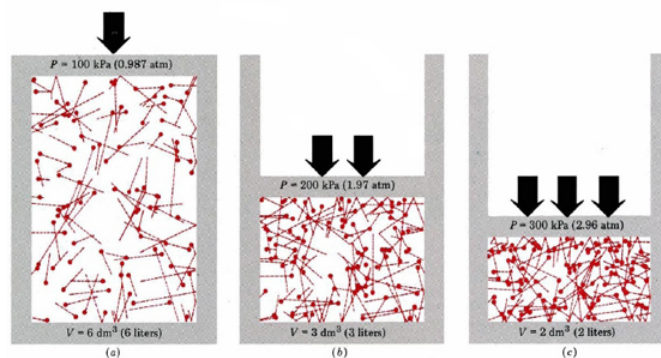
From them it is possible to derive the following expression for the pressure of a gas in terms of the properties of its molecules:

$$P = \frac{1}{3}\frac{N}{V}m(u^2)_{\text{ave}} \tag{9.13.1}$$

where P, V = pressure and volume of the gas, N is the number of molecules, m s mass of each molecule and $(u^2)_{\text{ave}}$ = average (or mean) of the squares of all individual molecular velocities. This mean square velocity must be used because pressure is proportional to the square of molecular velocity, and molecular collisions cause different molecules to have quite different velocities.

Rather than concerning ourselves with the procedure for deriving Eq. 9.13.1, let us inspect the equation and see that its general features are much as we would expect. In some ways, the ability to do this with a formula is more useful than the ability to derive it. Figure 9.13.1

First of all, the equation tells us that the pressure of a gas is proportional to the number of molecules divided by the volume. This is shown graphically in Figure 9.13.1, where a computer has drawn the same number of gas molecules occupying each of three different volumes. The "tail" on each molecule shows the exact path followed by that molecule in the previous microsecond—the longer the tail, the faster the molecule was going. The average of the squares of the tail lengths is proportional to $(u^2)_{\text{ave}}$ and is the same in all three diagrams. It is also assumed that all the molecules have equal masses.

As you can see, reducing the volume of the gas increases the number of collisions per unit area on the walls of the container. Each collision exerts force on the wall; force per unit area is pressure, and so the number of collisions per unit area is proportional to pressure. Halving the volume doubles the pressure, a prediction which agrees with the experimental facts summarized in Boyle's law. Equation 9.13.1 also says that the pressure is proportional to the mass of each gas molecule. Again, this is what we would expect. Heavy molecules give a bigger "push" (the technical term for this is **impulse**) against the wall than do light ones with the same velocity.

Finally, the equation tells us that pressure is proportional to the average of the squares of the molecular velocities. This dependence on the square of velocity is reasonable if we realize that doubling the velocity of a molecule has *two* effects.

First, the molecule can move farther in a given length of time, doubling the number of collisions with the walls. This would double the pressure. Second, doubling the velocity of a molecule doubles the push or impulse of each collision. This doubles the pressure again. Therefore doubling a molecule's velocity quadruples the pressure, and for a large number of molecules, *P* is proportional to the mean *square* velocity.