Introduction

We've talked about the gas laws and how they were derived from experiment. As scientists, we would like to figure out **why** gases obey the empirical gas laws and the ideal gas equation. If we can form a mental picture of a gas at the molecular level, we would better understand things like why the pressure of a gas increases if it's held in a fixed volume container and the temperature is raised.

The kinetic theory: Molecules in motion

Several theories have been proposed in the past about gas behavior. The theory currently accepted today is called the **kinetic theory**. The word **kinetic** refers to **motion**, and the kinetic theory of gases explains gas properties using the motion of gas molecules.

According to kinetic theory, **pressure** is a result of moving gas molecules hitting the sides of the container they're in. How is this pressure? Imagine yourself holding up a sheet of cardboard. Now imagine that your friends are throwing tennis balls as hard as they can at the cardboard you're holding.

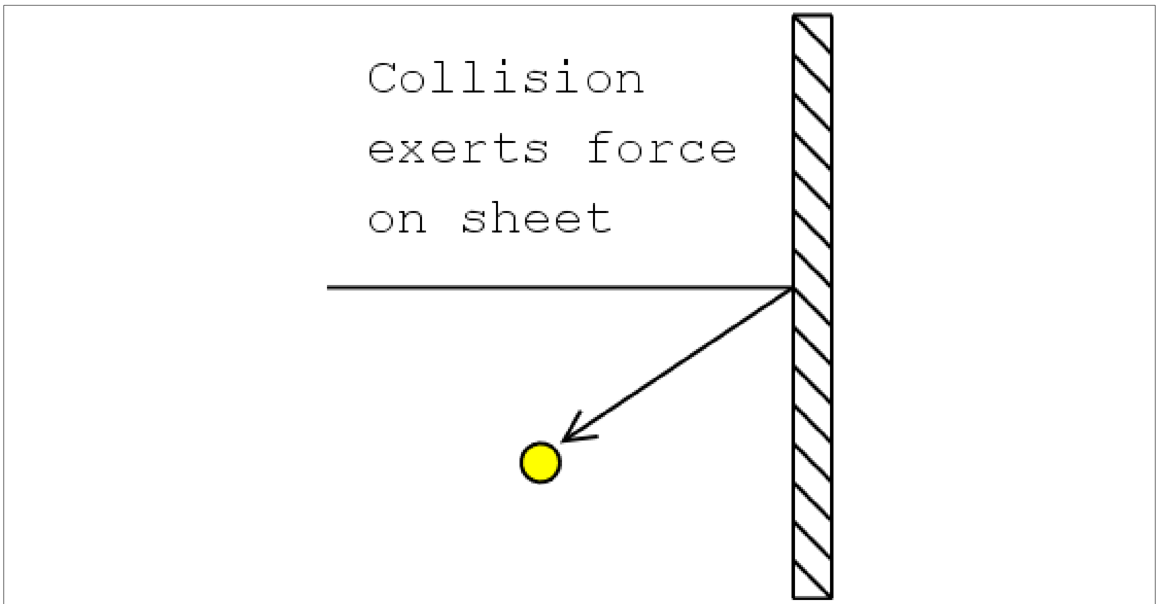


Illustration 1 - Pressure is caused by collisions

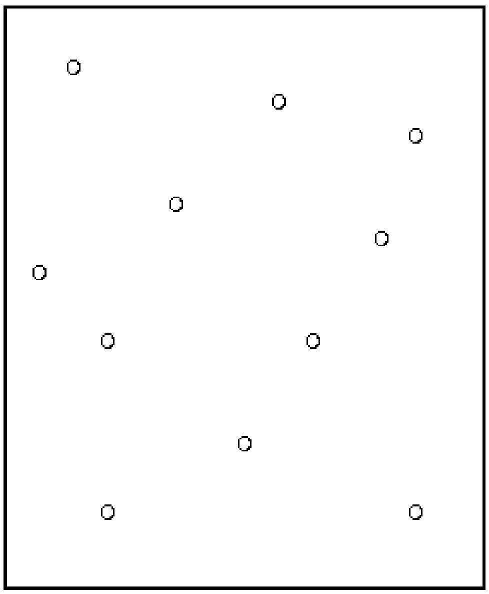
You'd feel pressure against the cardboard - that is, you'd feel a **force** from the tennis balls acting over the **area** of the sheet of cardboard.

Postulates of kinetic theory

The kinetic theory (like Dalton's atomic theory) has several **postulates**. You'll need to be familiar with these postulates, but you might find it easier to think of them in terms of a **picture** of a gas.

Postulate 1: Gas molecules are small compared to the average distance between them.

Remember that a gas is **not** a dense phase of matter, so there's a lot of empty space around a gas molecule. That's exactly what the first postulate says. Here is a picture that represents this postulate.

	• Gases have a lot of empty space in them! • Gas molecules are very small compared to this amount of empty space (so gases are light)!
Illustration 2 - Gas molecules are far apart.	

So the picture we have of gases so far is of small gas molecules in a large space.

Postulate 2: Molecules move randomly in straight lines in all directions and with various speeds.

Here's where we get into the **kinetic** part of the theory. These molecules are randomly moving around in different directions. Since the molecules are moving through essentially **nothing**, they will move in straight lines until they hit something. Think of balls on a pool table. This adds some detail to our picture.

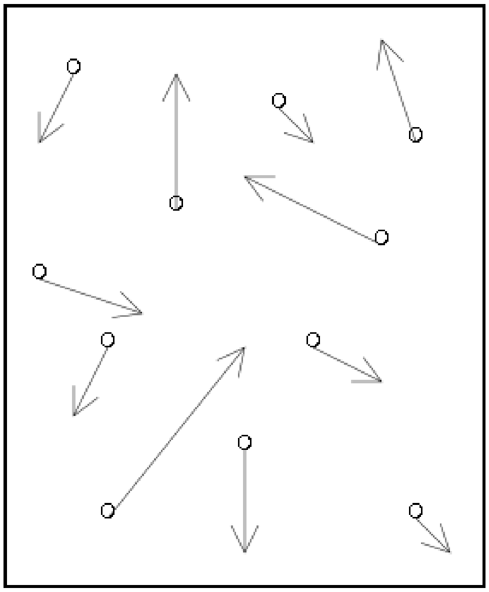


Illustration 3: Gas molecules in motion

- The gas molecules are in motion through the empty space!
- Molecules move in straight lines since unless they hit something there's no reason to change direction!

But what about the other molecules and the wall? Do they have an influence on the gas molecules? This is covered by the third and fourth postulates of kinetic theory.

Postulate 3: Attractive and repulsive forces between gas molecules are negligible except at a collision.
Postulate 4: Collisions between molecules and between molecules and the wall are elastic.

Let's look at the third postulate. What are attractive and repulsive forces? The forces that are being discussed here are the forces between molecules that hold solids and liquids together. We'll discuss these in detail in CHM 111, but you can compare these right now to forces between charged species. **If gas molecules are far apart** (as they are in a gas), these forces are so small that the molecules aren't affected by them.

The fourth postulate says that the collisions between gas molecules are **elastic**. Effectively, this means that kinetic energy is transferred from one molecule to another, but no kinetic energy is **lost** during the collision as heat, etc. Visually, you can think of these collisions as billiard balls on a pool table.

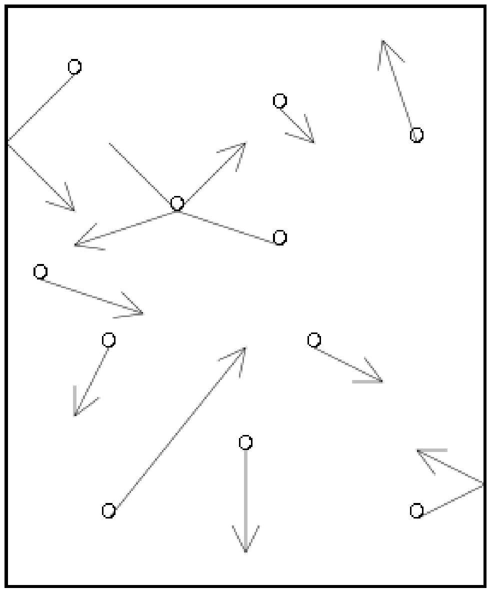


Illustration 4: Collisions are elastic

- Gas molecules don't have curvy paths because they aren't influenced by anything except collisions!
- Gas molecules collide with each other and the wall much like billiard balls do!

The last postulate talks about temperature (Remember that temperature also appears in the ideal gas law!)

Postulate 5: The average kinetic energy of the gas molecules is proportional to the absolute temperature.

The **average kinetic energy** of the molecules is the average "energy of motion" the molecules have, and the **absolute temperature** we will express in units of **Kelvins**. The energy of motion depends on **how fast** the molecules are moving. This postulate tells us that **the speed of the gas molecules is related to the average temperature.** The hotter it is, in other words, the faster the gas molecules move. We can add this to our picture.

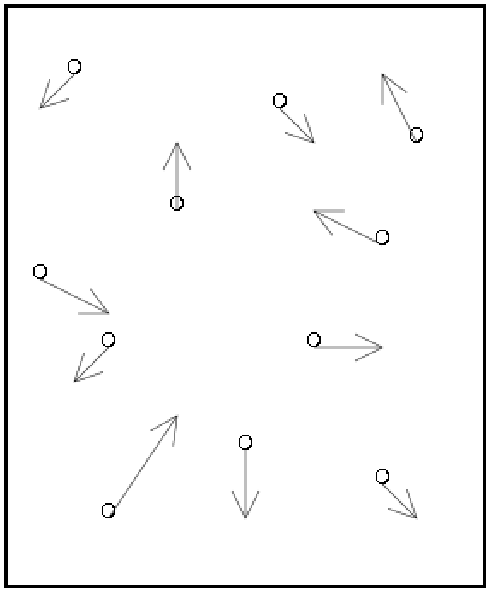


Illustration 5: Gas at low temperature

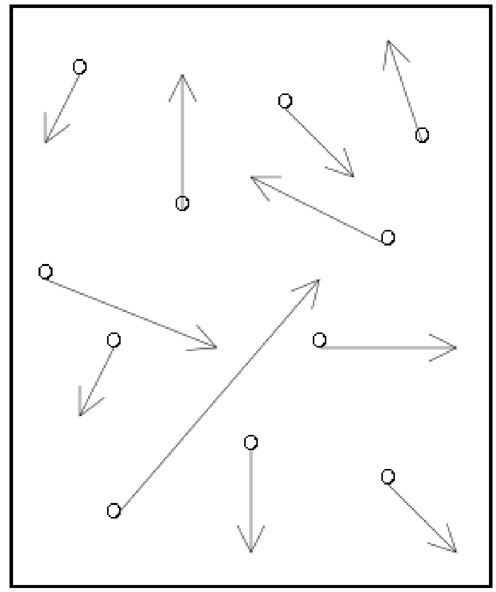


Illustration 6: Gas at higher temperature

The length of the arrows in front of the gas molecules indicates the speed of the gas molecules. But what conclusions can we draw from all this? How can we relate this to the empirical gas laws?

Conclusions: Relating theory to experiment

The most important things to get from kinetic theory are that

- Pressure is created by collisions of gas molecules with each other and the container walls.
- Temperature controls how fast the molecules move.

Pressure can be increased by either increasing the **number of collisions** or increasing **how hard each wall is hit**.

Let's look at changing temperature while holding pressure constant - in other words, **Charles's law.** Charles's law ($V_1/T_1 = V_2/T_2$) says that if you increase the temperature of a gas, it will take up more space at the same pressure. **Is this predicted by kinetic theory**?

Kinetic theory says that if you increase the temperature, the molecules move faster. If the amount of space the molecules had to move in was fixed, they would hit the walls **harder and more often**. This would mean **pressure would increase**. But we said the pressure was constant! How can we hold the pressure constant? We can give the molecules more room to move in, or **increase the volume**! So, kinetic theory agrees with Charles's law - as the temperature increases, so does volume.

What about **Boyle's law**? ($P_1V_1 = P_2V_2$) Boyle's law states that if the volume is decreased at constant temperature, pressure increases. **Is this predicted by kinetic theory**?

At constant temperature, the average speed of the molecules will be the same. If we decrease the volume, we will have **more molecules per unit of space** traveling at the same speed they were before. The **number of collisions** will increase. Since pressure is created by collisions, the **pressure increases**. So, kinetic theory also agrees with Boyle's law.

Try on your own to test if these conclusions from the gas laws are predicted by kinetic theory:

- As temperature increases, pressure increases (at constant volume).
- As volume increases, pressure decreases (at constant temperature).
- As the number of moles of gas increases, the pressure increases (at constant volume and temperature).

Summary

This note pack describes the kinetic theory of gases. The kinetic theory attempts to make sense of the gas laws - why are pressure, volume, and temperature related for gases the way they are? Gases are treated as being made of small particles (atoms and molecules) in a large volume of space - and we developed a picture of how a gas looks at the molecular level. This picture isn't perfect (what is?), but we can use it to explain every trend we've observed from the gas laws.