

Chapter 6

Gases

Opening Essay

Perhaps one of the most spectacular chemical reactions involving a gas occurred on May 6, 1937, when the German airship *Hindenburg* exploded on approach to the Naval Air Station in Lakehurst, New Jersey. The actual cause of the explosion is still unknown, but the entire volume of hydrogen gas used to float the airship, about 200,000 m³, burned in less than a minute. Thirty-six people, including one on the ground, were killed.

Hydrogen is the lightest known gas. Any balloon filled with hydrogen gas will float in air if its mass is not too great. This makes hydrogen an obvious choice for flying machines based on balloons—airships, dirigibles, and blimps. However, hydrogen also has one obvious drawback: it burns in air according to the well-known chemical equation

$$2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\ell)$$

So although hydrogen is an obvious choice, it is also a dangerous choice. Helium gas is also lighter than air and has 92% of the lifting power of hydrogen. Why, then, was helium not used in the *Hindenburg*? In the 1930s, helium was much more expensive. In addition, the best source of helium at the time was the United States, which banned helium exports to pre–World War II Germany. Today all airships use helium, a legacy of the *Hindenburg* disaster.

Of the three basic phases of matter—solids, liquids, and gases—only one of them has predictable physical properties: gases. In fact, the study of the properties of gases was the beginning of the development of modern chemistry from its alchemical roots. The interesting thing about some of these properties is that they are independent of the identity of the gas. That is, it doesn't matter if the gas is helium

gas, oxygen gas, or sulfur vapors; some of their behavior is predictable and, as we will find, very similar. In this chapter, we will review some of the common behaviors of gases.

Let us start by reviewing some properties of gases. Gases have no definite shape or volume; they tend to fill whatever container they are in. They can compress and expand, sometimes to a great extent. Gases have extremely low densities, one-thousandth or less the density of a liquid or solid. Combinations of gases tend to mix together spontaneously; that is, they form solutions. Air, for example, is a solution of mostly nitrogen and oxygen. Any understanding of the properties of gases must be able to explain these characteristics.

6.1 Kinetic Theory of Gases

LEARNING OBJECTIVES

State the major concepts behind the kinetic theory of gases.

Relate the general properties of gases to the kinetic theory.

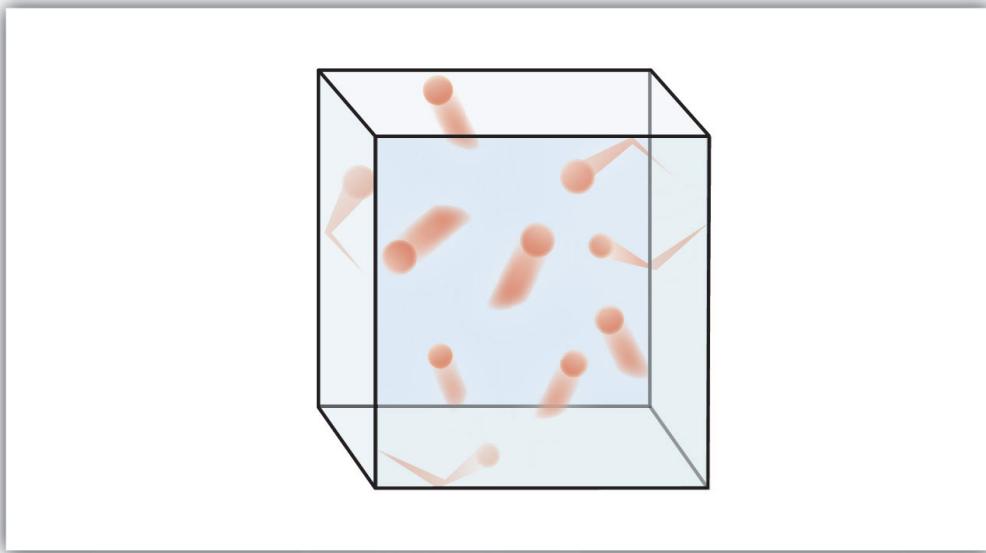
Gases were among the first substances studied in terms of the modern scientific method, which was developed in the 1600s. It did not take long to recognize that gases all shared certain physical behaviors, suggesting that all gases could be described by one all-encompassing theory. Today, that theory is the **kinetic theory of gases**. It is based on the following statements:

1. Gases consist of tiny particles of matter that are in constant motion.
2. Gas particles are constantly colliding with each other and the walls of a container. These collisions are elastic; that is, there is no net loss of energy from the collisions.
3. Gas particles are separated by large distances, with the size of a gas particle tiny compared to the distances that separate them.
4. There are no interactive forces (i.e., attraction or repulsion) between the particles of a gas.

5. The average speed of gas particles is dependent on the temperature of the gas.

[Figure 6.1 "The Kinetic Theory of Gases"](#) shows a representation of how we mentally picture the gas phase.

Figure 6.1 *The Kinetic Theory of Gases*



The kinetic theory of gases describes this state of matter as composed of tiny particles in constant motion with a lot of distance between the particles.

This model of gases explains some of the physical properties of gases. Because most of a gas is empty space, a gas has a low density and can expand or contract under the appropriate influence. The fact that gas particles are in constant motion means that two or more gases will always mix, as the particles from the individual gases move and collide with each other.

An **ideal gas** is a gas that exactly follows the statements of the kinetic theory.

Unfortunately, *real gases* are not ideal. Many gases deviate slightly from agreeing perfectly with the kinetic theory of gases. However, most gases adhere to the statements so well that the kinetic theory of gases is well accepted by the scientific community.

KEY TAKEAWAYS

- The physical behavior of gases is explained by the kinetic theory of gases.

- An ideal gas adheres exactly to the kinetic theory of gases.

EXERCISES

1. State the ideas behind the kinetic theory of gases.
2. The average speed of gas particles depends on what single variable?
3. Define *ideal gas*. Does an ideal gas exist?
4. What is a gas called that is not an ideal gas? Do such gases exist?

ANSWERS

1. Gases consist of tiny particles of matter that are in constant motion. Gas particles are constantly colliding with each other and the walls of a container. These collisions are elastic; that is, there is no net loss of energy from the collisions. Gas particles are separated by large distances, with the size of a gas particle tiny compared to the distances that separate them. There are no interactive forces (i.e., attraction or repulsion) between the particles of a gas. The average speed of gas particles is dependent on the temperature of the gas.
3. An ideal gas is a gas that exactly follows the statements of the kinetic theory of gases. Ideal gases do not exist, but the kinetic theory allows us to model them.