Lesson Objectives

I can:

☐ describe the properties of gases in terms of the kinetic molecular theory.
☐ convert between different units of pressure (atm, mmHg, torr, kPa, psi), temperature (K and C), and volume (L and mL).
☐ describe the relationship between the pressure, volume, temperature, and the amount of a gas
☐ describe the properties of a mixture of gases using Dalton's Law of Partial Pressures.
☐ compare the average speed of gas molecules at a particular temperature.

Class Notes

Kinetic Molecular Theory

We can describe gas samples by:

• The kind of gases in a sample
• The amount of gases in a sample (i.e. number of molecules, moles, or mass)
• The conditions of the gas sample (i.e. pressure, volume, temperature)

Kinetic Molecular Theory (KMT) makes some assumptions that help us to imagine how molecules behave:

• Gas particles are in constant, rapid, random, straight-line motion.
• There are no attractive forces between gas molecules.
• Collisions between gas particles are completely elastic.
• Gas particles occupy negligible volume.

We can use these assumptions to describe gas samples:

• The Kelvin Temperature is directly related to the average kinetic energy of the gas sample.
• The Pressure of a gas sample is related to the number of collisions against the walls of the container.

Gas samples that behave according to these assumptions are called ideal gases. While no real gases are completely ideal, these assumptions allow us to make very good approximations when describing gases.
Example 1: Use the ideas of Kinetic Molecular Theory to make and explain predictions about gas samples.

- In a rigid container, how does the pressure change when the temperature increases?

- In a rigid container, how does the pressure change when the sample contains more moles of gas (at constant temperature)?

- In a piston, how does the pressure change when the volume is decreased?

- In a balloon, how does the volume change when the temperature is increased at constant pressure?

Volume (V)
The volume is the space occupied by a gas sample, which is the volume of the container, usually measured in units of liters (L).

\[ 1 \text{ L} = 1000 \text{ mL} \]

Pressure (P)
In a physical sense, the pressure is the force exerted over an area, and has units of pascal (Pa). In a gas sample, we can imagine this force arising from gas molecules hitting the walls of its container: the more frequently and the harder gas molecules hit the walls, the greater the pressure.

Evangelista Torricelli (1608-1677), an Italian physicist, is credited for creating the first barometer by placing an evacuated tube in a basin of mercury. He found that the mercury rose 760 mm because of the pressure on the mercury exerted by the air.

We define this pressure as **Standard Pressure** (in either units of mmHg or torr, named after Torricelli). We also use Standard Pressure to define the atmospheric pressure (atm):

\[ 1 \text{ atm} = 760 \text{ mmHg} = 760 \text{ torr} = 101.325 \text{ Pa} = 101.325 \text{ kPa} = 29.92 \text{ inHg} = 14.7 \text{ psi} \]

(psi is “pounds per square inch” and commonly used to measure tire pressure. inHg is commonly used in meteorological reports.)

We can use manometers to determine the pressure of a gas sample.

Example 2:

- A sample of Ar gas is introduced in an evacuated flask of a closed-end manometer. If the change in height is 25 mm, find the pressure of the Ar gas in atm.
• In an open-end manometer, an unknown gas is introduced into the flask. Calculate the pressure of the gas in atm if

○ The level of Hg on the side of the flask is higher than the open end by 8 mm. (Picture D)

○ The level of Hg on the open end is higher by 10 mm.

**Dalton’s Law of Partial Pressures**

When there is a mixture of gases, we can describe the pressure by each gas separately – the partial pressure. The partial pressure of a gas is the pressure of that gas if it were the only gas in the sample. The total pressure of all gases in a gas sample is therefore the sum of the partial pressures of all gases:

\[ P_{\text{total}} = P_A + P_B + \ldots \]

In addition, the partial pressure of a gas does not depend on what kind of gas it is, but rather how many molecules of that gas are in the sample.

\[ \frac{P_A}{P_{\text{total}}} = \frac{n_A}{n_{\text{total}}} \quad \text{or} \quad P_A = \left( \frac{n_A}{n_{\text{total}}} \right) (P_{\text{total}}) \]

The fraction \( \frac{n_A}{n_{\text{total}}} \) is called the mole fraction of A.

**Example 3:** A container with 2.5 mol H\(_2\) and 5.0 mol O\(_2\) has a total pressure of 3.6 atm. What is the partial pressure of each gas?

**Temperature (T)**

Often we measure the temperature in Celsius (°C), since those are the units indicated by our thermometers. However, in KMT, we want to relate the temperature to the molecules’ kinetic energy, or their motion: that is, the molecules are moving faster at a higher temperature. The kelvin (K) temperature scale has the same increments as Celsius, but it has a different zero-point.

\[ T_K = T_{\circ C} + 273 \]

Because the kelvin temperature is related to a sample’s kinetic energy, when we double the kelvin temperature, we also double its kinetic energy. Absolute Zero, or 0 K, is when a sample has no kinetic energy; molecules have no motion.

**Standard Temperature** is defined as 0°C, or 273 K, while **Room Temperature** is 25°C.
Gas Laws

Boyle’s Law: Pressure vs. Volume

Charles’ Law: Volume vs. Temperature

Gay-Lussac’s Law: Pressure vs. Temperature

Combined Gas Law
Example 1: A sample of gas at 1.4 atm increases in volume from 250 mL to 400 mL at a constant temperature of 30°C. What is the final pressure of the sample?

Example 2: A sample of neon gas in a rigid 1.0 L container at 15°C has a pressure of 600 torr. What is the final pressure if the sample is heated to a temperature of 100°C?

Example 3: A sample of carbon dioxide at 10°C occupies 600 mL and has a pressure of 450 mmHg. What is the volume of the sample if the sample is brought to STP?

**Ideal Gas Law**

Example 1: What is the pressure in atm of a sample of 0.20 mol of oxygen gas that occupies a volume of 4.1 L at 40°C?
Example 2: What is the mass of a sample of hydrogen gas that occupies a volume of 2.0 L at 20°C and 720 mmHg?

What is the density of this gas?

**Graham’s Law of Effusion**

In physics, kinetic energy is directly related to an object’s mass and to the square of its velocity:

\[ KE = \frac{1}{2} mv^2 \]

If we compare two objects with the same kinetic energy, the mass is inversely related to the square of its velocity: the greater the mass, the slower the object moves.

**Diffusion** is the mixing of samples of gases, and is pretty difficult to model; **effusion** is the leaking of a gas through a small opening. **Graham’s Law of Effusion** compares the speed that different gases travel:

\[ \frac{v_A}{v_B} = \sqrt{\frac{MM_B}{MM_A}} \]

where \( v_A \) and \( v_B \) are the average speeds of gases A and B in m/s, and \( MM_A \) and \( MM_B \) are the molar masses of gases A and B. This tells us that at a given temperature, the heavier the gas (larger molar mass), the slower it travels (smaller speed), while the lighter the gas (smaller molar mass), the faster it travels (greater speed).