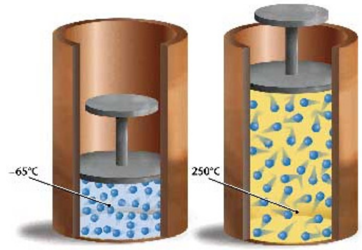
Unit IX: Gas Laws Review Sheet (Chapter 12)



1. Describe the motion of particles of a gas according to the kinetic theory.

2. Explain gas pressure and volume in terms of kinetic theory.

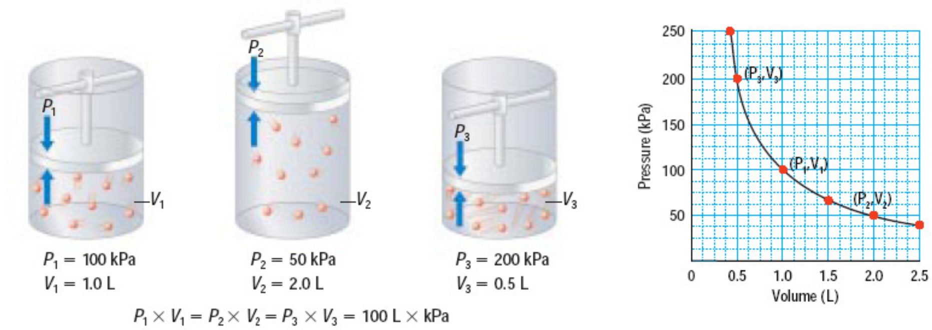
3. Be able to read, and/or describe the design/function of a barometer, and open-ended/closed-ended manometers, to determine the pressure of a gas sample.

4. Define the temperature of a substance as a measure of the kinetic energy of the particles in the substance. Temperature is simply an average of the kinetic energy of the gas particles. The temperature represents the KE of the majority of the gas molecules.

5. Be able to REASON and USE the four Gas Laws, that are simply derivations of the Combined Gas Law:

    1. **Boyle's Law:** describes the relationship of pressure and volume of a gas at constant temperature. As the pressure increases, the volume will decrease, at constant temperature.

$$P_1 \times V_1 = P_2 \times V_2$$



    2. **Charles's Law:** describes the relationship of temperature and volume of a gas at constant pressure. As the temperature of a gas increases, so too does the volume, while pressure is held constant:

$$V_1/T_1 = V_2/T_2$$

What is the significance of **absolute-zero temperature**? This is the temperature at 0 degrees Kelvin, where gas molecules have no motion (no Kinetic Energy), and therefore, the gas has no volume.

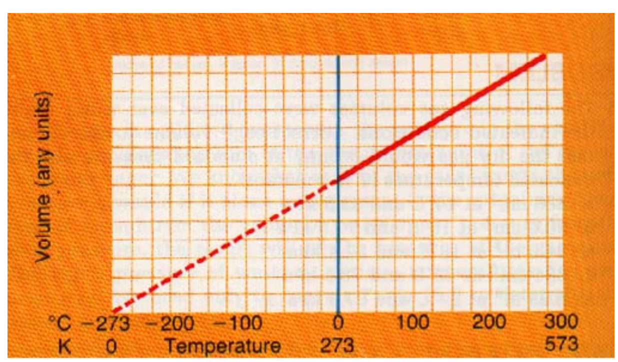


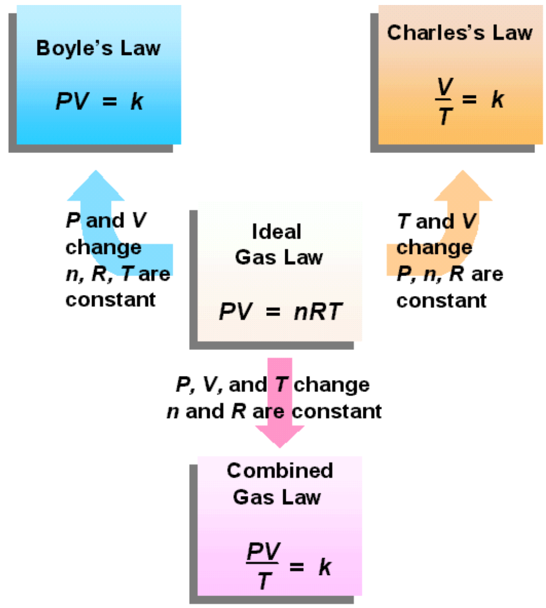
FIGURE 18-5. The volume of a gas at zero kelvin is theoretically zero.

3. **The Gay-Lussac Law:** describes the relationship of temperature vs. pressure of a gas at constant volume

$$P_1/T_1 = P_2/T_2$$

A) **The Combined Gas Law:** describes the relationship between pressure, volume and temperature of a contained gas. Simply put, pressure, volume, and temperature will vary so that P x V / T equals a constant that we could call "k." From this relationship, we can write:

$$V_1 P_1/T_1 = V_2 P_2/T_2$$



6. Following the four laws that collectively describe pressure, volume, and temperature, we finally have the **Ideal Gas Law: PV = nRT,** which allows scientists to now calculate the number of moles of a gas, along with pressure, temperature, and volume. In addition to moles of a gas, we can now extend the calculations to include grams, or molecules.

Rearranging the Ideal Gas Law, we can:

- Calculate the molar mass or density of a gas.
- Calculate the amount of gas (moles, grams, molecules, etc) at any specified conditions of pressure, volume and temperature.

- $PV = nRT$
- Molar mass $= \dfrac{m R T}{P V}$
- Density $= \dfrac{P \times Mmass}{R T}$

7. The next Gas Law that we studied was **Dalton's Law of Partial Pressures**, which states:

$$\mathbf{P_{Total} = P_1 + P_2 + P_3 + \ldots}$$

From this relationship, we can use Mole Fraction of each gas in the mixture to determine the Partial Pressure of each gas.

$$\mathbf{P_A = (Mole_A/Moles_{Total}) * P_{Total}}$$

Included in this discussion were the problems where gases are collected over water. In those cases, the vapor pressure of water is obtained from a Water Vapor Pressure Table at the given temperature, and the pressure of the "dry gas" is derived from the following:

$$\mathbf{P_{Dry\ Gas} = P_{Total} - P_{water}}$$

Once the pressure of the dry gas is determined, fully expect to solve an Ideal Gas Law problem to determine moles, molecules, grams, etc of that dry gas.

8. We quickly recapped **Avogadro's Hypothesis** using the kinetic theory. The significance of this is that we can directly relate the volumes of gases in a balanced chemical equation, without having to calculate, and compare moles across a reaction.

9. Kinetic Energy is calculated as:

K.E. = ½ mv² *where m = mass & v = velocity of the gas molecules*

Since we know that temperature is simply a measure of the average Kinetic Energy of molecules within a sample, then it follows that two gases at the same temperature have the same K.E. It makes sense, then, that the heavier gas (larger mass) will travel slower (lower v), while the lighter gas (lower m) will have a higher velocity (larger v).

If we rearrange the KE equation: velocity of the gas $\mathbf{v = (2K.E./Mass)^{1/2}}$ or, the velocity of a gas is inversely proportional to the square root of the mass.

10. This is the basis for **Graham's Law:**

$$\frac{Rate_A}{Rate_B} = \frac{\sqrt{Mass_B}}{\sqrt{Mass_A}}$$

When using Graham's Law, be careful that you are comparing rates, and not simply times. We went over a few examples in class that show how to manage times.

11. *What differentiates a real gas from an ideal gas?* The Ideal Gas Law assumes,
    - The molecules of an ideal gas have no volume
    - And, there are no attractive forces between the molecules within a gas.
    - The gas molecules move in a random manner, and their collisions with each other, and the container, are perfectly elastic.

This, of course, is not true. There are attractive forces between the molecules, and the gas molecules do have volume, although small. With these in mind, Real gases behave like Ideal Gases at the following conditions:
- At low pressures: the gas molecules spread apart so that the small volume of each gas molecule is negligible relative to the entire volume of the gas.
- At high temperatures: the gas molecules are moving so quickly; the small attractive forces between the molecules of a gas are overcome by their rapid speed.

What types of problems will you see on your test? *Bank on the following:*

1. Boyle's Law
2. Charles's Law
3. Gay-Lussac's Law
4. Combined Gas Law
5. The Ideal Gas Law (any of the 3 flavors)
6. Ideal Gas Law Stoichiometry
7. Dalton's Law of Partial Pressure
8. Gas Over Water
9. Graham's Rate of Effusion