

## A Molecular Model of Gases

### observation

Gases are easy to expand

Gases are easy to compress

Gases have densities that are about 1/1000 of solid or liquid densities

Gases completely fill their containers

Hot gases leak through holes faster than cold gases

### hypothesis

gas molecules don't strongly attract each other

gas molecules don't strongly repel each other

molecules are much farther apart in gases than in liquids and solids

gas molecules are in constant motion

the hotter the gas, the faster the molecules are moving

### A Few Gas Laws

#### Boyle's Law

This law says: *The pressure and the volume of a gas are inversely proportional to each other, temperature kept constant.* What this means, is that if you keep the temperature constant, then if you increase the pressure on the gas, the volume will decrease, and if you lessen the pressure on a gas, the volume will increase.

Boyle's Law can be expressed like this:  $k = P / V$ .

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#### Charles's Law

This law says: *The temperature and the volume of a gas are directly proportional, assuming pressure is kept constant.* That means as the temperature goes up, the volume will go up as well.

This is easy to picture. If the temperature goes up. The gas expands, and the space it takes up increases. The opposite also happens; if temperature is lowered, volume will be less.

Charles Law is  $V = kT$ , pressure being held constant.

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#### Avogadro's Law

Equal moles of gases at the same temperature and pressure will occupy equal volumes. STP = 0 °C, and at 1 atm. At STP, one mole of gas, any gas, will occupy exactly 22.4 liters.

Units for pressure are tremendously annoying. You have to be good at converting because ideal gas law problems demand unit consistency. Also remember to always use absolute temperature.

$$1 \text{ atm} = 760 \text{ torr}$$

$$1 \text{ atm} = 760 \text{ mmHg} \quad 1 \text{ mmHg} = 1 \text{ torr}$$

$$1 \text{ atm} = 101.3 \text{ Kpa}$$

$$1 \text{ atm} = 14 \text{ psi}$$

$$K = C^{\circ} + 273 \quad F = 1.8C^{\circ} + 32$$

$$760 \text{ mmHg or torr} = 1 \text{ atm}$$

- most behaviors of a pure gas sample can be related to just 4 physical properties:

convenient			
property	symbol	units	
pressure	P	atm, torr, Pa	
volume	V	L, cm <sup>3</sup>	
temperature	T	K	
moles	n	mol	

## • SUMMARY OF GAS LAWS

changing these variables in an experiment...	while holding these variables constant...	...reveals this relationship:	
P, V	n, T	$P_1 V_1 = P_2 V_2$	Boyle's Law
V, T	n, P	$V_1 / T_1 = V_2 / T_2$	Charles' Law
n, V	P, T	$V_1 / n_1 = V_2 / n_2$	Avogadro's Law
		1 mole at STP = 22.4 L	
P, V, T	n	$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$	Combined Gas Law
P, V, T, n		Ideal Gas Law	
		$PV = nRT$	

- standard temperature and pressure (STP)
  - convenient for reporting gas volumes
  - standard temperature is 0°C (273.15 K).
  - standard pressure is 1 atm

$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$

**COMBINE GAS LAW ;** USE WHEN CONDITIONS CHANGE

**Standard Condition: STP = 1 atm and 0°C at STP 1 mole = 22.4L**

**IDEAL GAS LAW**

**PV = nRT** USE WHEN CONDITIOND *DO NOT* CHANGE

- P - pressure (in atms).
- V - volume (in liters).
- n - quantity of gas (in moles).
- R - the gas constant .0821 atm l / mol K.
- T - temperature, IN KELVINS

**DALTON'S LAW OF PARTIAL PRESSURES**

$P_T = P_1 + P_2 + P_3 + \dots$  total pressure = sum of all partial pressures

- Therefore, the pressure of a mixture of gases is equal to the sum of the partial pressures of the individual gases.
- **partial pressure:** pressure exerted by one gaseous component in a mixture

**THE KINETIC MOLECULAR THEORY AND GRAHAM'S LAW**

Since  $KE_{avg}$  is dependent only upon T, two different gases at the same temperature must have the same  $KE_{avg}$ .

$$\frac{1}{2} m_H v_H^2 = \frac{1}{2} m_O v_O^2$$

Simplify the equation by multiplying both sides by two:

$$m_H v_H^2 = m_O v_O^2$$

Rearrange to give the following:

$$\frac{v_H^2}{v_O^2} = \frac{m_O}{m_H}$$

Take the square root of both sides to obtain the following relationship between the ratio of the velocities of the gases and the square root of the ratio of their molar masses:

$$\frac{v_H}{v_O} = \frac{\sqrt{m_O}}{\sqrt{m_H}} \quad \text{OR} \quad \frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$$

This equation states that the velocity (rate) at which gas molecules move is inversely proportional to the square root of their molar masses.

**THE KINETIC MOLECULAR THEORY** is used to explain the behavior of gases and is based upon the following postulates:

1. Gases are composed of a many particles in constant, random motion.
2. Collisions between gas particles or collisions with the walls of the container are elastic. That is, none of the energy of the gas particle is lost in a collision.
3. These particles are much smaller than the distance between particles, therefore the volume of a gas is mostly empty space and the volume of the gas molecule is considered to be zero.
4. There is no force of attraction between gas particles or between the particles and the walls of the container.
5. The average kinetic energy of a collection of gas particles is dependent only upon the temperature of the gas.

In short, the key parts are:

1. **constant random motion**
2. **elastic collisions**
3. **negligible volume**
4. **no forces of attraction**
5.  $KE = \frac{1}{2}m \cdot v^2$

### **Real vs Ideal**

**Pressure.** The Ideal Gas Law assumes that there are no interactions between molecules. They just bounce around, with some kinetic energy. But in reality, the molecules are attracted to each other. Therefore, they wouldn't be hitting the walls of the container as hard as are expected. Therefore the actual pressure will be greater than the observed pressure.

**Volume.** The Ideal Gas Law assumes that the molecules are points that don't take any actual space themselves. But in reality...molecules take up space!!! (Don't act surprised.) So the actual volume that is there is less than what is observed, since the molecules take some space up which the gas should be taking up according to the Ideal Gas Law.

**Practice problems on the following page:**

### Problems and Answers:

1. A sample of 0.50 moles of gas is placed in a container of volume of 2.5 L. What is the pressure of the gas in torr if the gas is at 25°C?  $P = 4.89 \text{ atm} = 3719 \text{ torr}$
2. A sample of gas is placed in a container at 25°C and 2 atm of pressure. If the temperature is raised to 50°C, what is the new pressure?  $P = 2.17 \text{ atm}$
3. At 1 atm of pressure water boils at 100°C, if the sample was placed under 2 atm of pressure, what would be the temperature? (This would be like a pressure cooker).  
 $T = 746 \text{ K} = 473^\circ\text{C} = 883^\circ\text{F}$
4. At what temperature would water boil if the pressure is 600 torr? (Use information from problem 3: this shows why food doesn't cook well at higher elevations)  
 $T = 294 \text{ K} = 21.5^\circ\text{C} = 70.7^\circ\text{F}$
5. Calculate the volume of 40.6 g of  $\text{F}_2$  at STP.  $V = 23.9 \text{ L}$
6. A sample of 2.0 moles of hydrogen gas is placed in a container with a volume of 10.4 L. What is the pressure of the gas in torr if the gas is at 25°C?  $P = 4.70 \text{ atm} = 3576 \text{ torr}$
7. The tire pressure is 32 psi. What is the pressure in torr if 1 atm = 14.7 psi?  
 $P = 1654 \text{ torr}$
8. A gas is placed in a balloon with a volume of 3.0 L at 28°C and 900 torr. What would be the new volume for the gas if placed under STP?  $V = 3.2 \text{ L}$
9. How many moles of gas would occupy a volume of 14 L at a pressure of 700 torr and a temperature of 30°C?  $n = 0.52 \text{ mol}$
10. Calculate the volume of 24.0 g of HCl at STP.  $V = 14.8 \text{ L}$
11. What is the volume of one mole of acetylene gas at STP?  $V = 22.414 \text{ L}$
12. What is the volume of 0.75 mol of gas at 72°C and 2 atm?  $V = 10.6 \text{ L}$
13. A student collects a sample of gas at 0.97 atm and 26°C which occupies a volume of 3.5 L, calculate its volume at STP.  $V = 3.1 \text{ L}$
14. Ammonia ( $\text{NH}_3$ ) is placed in 1.5 L flask at 25°C. If the pressure of the gas is 0.899 atm, what is the density?  $d = 0.626 \text{ g/L}$
15. A mixture of Ar and CO gases is collected over water at 28°C and an atmospheric pressure of 1.05 atm. If the partial pressure of Ar is 600 torr, what is the partial pressure of CO? (vapor pressure of water at 28°C is 28.3 mmHg)  $P_{\text{CO}} = 0.223 \text{ atm}$
16. Determine the partial pressures of each of the gases in the following mixture: 17.04 g  $\text{NH}_3$ , 40.36 g Ne and 19.00 g  $\text{F}_2$ . The gases are at 1.5 atm of pressure.  
 $P_{\text{NH}_3} = 0.428 \text{ atm}; P_{\text{Ne}} = 0.857 \text{ atm}; P_{\text{F}_2} = 0.2124 \text{ atm}$
17. Potassium chlorate decomposes under heat as follows:  
 $2 \text{KClO}_3 (\text{s}) \rightarrow 2 \text{KCl} (\text{s}) + 3 \text{O}_2 (\text{g})$   
The oxygen gas is collected over water at 25°C. The volume of gas is 560 mL measured at 1 atm. Calculate the number of grams of  $\text{KClO}_3$  used in the reaction. (vapor pressure of water = 0.0313 atm)  $n_{\text{O}_2} = 0.022 \text{ mol}; 1.81 \text{ g KClO}_3$