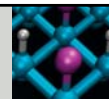


## Homework Assignments:



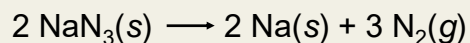
- Chapter 4 HW due **today** at 11:55 pm
- Increased number of attempts to 4 in OWL homework
- Chapter 5 HW due October 11 at 11:55 pm
- My office hours:
  - M** (10:00 – 10:50 am)
  - W, F** (12:00 – 1:00 pm)

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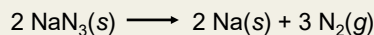
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## Stoichiometric Relationships with Gases

(1) How many liters of  $\text{N}_2$  at 1.15 atm and  $30.0^\circ\text{C}$  are produced by decomposition of 45.0 g  $\text{NaN}_3$  (sodium azide)?  
(Molar mass of  $\text{NaN}_3 = 65.010\text{ g/mol}$ )



## Stoichiometric Relationships with Gases



**Moles of N<sub>2</sub> produced:**

$$\frac{45.0 \text{ g}}{65.010 \text{ g/mol}} \times \frac{3 \text{ mol N}_2}{2 \text{ mol NaN}_3} = 1.04 \text{ mol}$$

**Volume of N<sub>2</sub> produced:**

$$V = \frac{nRT}{P} = \frac{(1.04 \text{ mol}) \left( 0.08206 \frac{\text{L atm}}{\text{K mol}} \right) (303 \text{ K})}{(1.15 \text{ atm})} = 22.5 \text{ L}$$

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## Partial Pressure and Dalton's Law

- **Dalton's Law of Partial Pressures:** The total pressure exerted by a mixture of gases in a container at **constant V and T** is equal to the sum of the pressures of each individual gas in the container.
- Partial pressure of gas 1 ( $P_1$ ) – The pressure that gas 1 would exert if it were alone in the container

$$P_{\text{total}} = P_1 + P_2 + \dots + P_N$$

$$\text{Mole Fraction (X)} = \frac{\text{Moles of component}}{\text{Total moles in mixture}}$$

Mole fraction of *i* gas,

$$X_i = \frac{n_i}{n_{\text{total}}} \quad \Rightarrow \quad P_i = X_i P \quad X_i = \frac{P_i}{P_{\text{total}}}$$

## Example Problem

- A scientist tries to generate a mixture of gases similar to a volcano by introducing 15.0 g of water vapor, 3.5 g of SO<sub>2</sub>, and 1.0 g of CO<sub>2</sub> into a 40.0-L vessel held at 120 °C
  - Calculate the partial pressure of each gas and the total pressure

For all of the gases,  $T = 120^\circ\text{C} = 393\text{ K}$ , and  $V = 40.0\text{ L}$ .

$$\text{H}_2\text{O}: 15.0\text{ g H}_2\text{O} \times \frac{1\text{ mol H}_2\text{O}}{18.0\text{ g H}_2\text{O}} = 0.833\text{ mol}$$

$$P_{\text{H}_2\text{O}} = \frac{(0.833\text{ mol})(0.08206\text{ L atm mol}^{-1}\text{ K}^{-1})(393\text{ K})}{40.0\text{ L}} = 0.672\text{ atm}$$

$$\text{SO}_2: 3.5\text{ g SO}_2 \times \frac{1\text{ mol SO}_2}{64.1\text{ g SO}_2} = 0.055\text{ mol}$$

$$P_{\text{SO}_2} = \frac{(0.055\text{ mol})(0.08206\text{ L atm mol}^{-1}\text{ K}^{-1})(393\text{ K})}{40.0\text{ L}} = 0.044\text{ atm}$$

$$\text{CO}_2: 1.0\text{ g CO}_2 \times \frac{1\text{ mol CO}_2}{44.0\text{ g CO}_2} = 0.023\text{ mol}$$

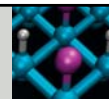
$$P_{\text{CO}_2} = \frac{(0.023\text{ mol})(0.08206\text{ L atm mol}^{-1}\text{ K}^{-1})(393\text{ K})}{40.0\text{ L}} = 0.018\text{ atm}$$

## Continuation from last page



Total pressure is the sum of these:

$$P_{\text{total}} = P_{\text{H}_2\text{O}} + P_{\text{SO}_2} + P_{\text{CO}_2} = 0.672 + 0.044 + 0.018 = 0.734\text{ atm}$$



34.2 g of  $\text{N}_2(\text{g})$ , 44.1 g  $\text{O}_2(\text{g})$ , and 63.7 g of  $\text{CO}_2(\text{g})$  are placed in a container. Which gas has the highest partial pressure?

1.  $\text{CO}_2$
2.  $\text{N}_2$
3.  $\text{O}_2$

Answer:  $\text{CO}_2$

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## Kinetic–Molecular Theory



- In many important practical settings, gases do not always behave ideally, especially at **very high pressure** and or **very low temperatures**
  - Nonideal gas behavior can be explained using **kinetic–molecular theory**

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## Postulates of the kinetic-molecular theory

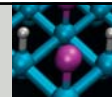


1. Gases are made up of **large collections of particles**, which are in constant random motion
2. Gas particles are infinitely small and occupy negligible volume
3. **Gas particles move in straight lines** except when they collide with other particles or with the container walls
  - ✓ These collisions are **elastic**, so kinetic energy of particles is conserved
4. Particles interact with each other only when collisions occur

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## Postulates of kinetic theory



- 5) The **average kinetic energy** of a gas is proportional to the **absolute temperature** of the gas but does not depend upon the identity of the gas

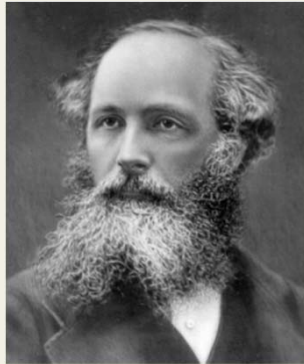
$$KE_{\text{avg}} = \frac{1}{2} m v_{\text{rms}}^2 \quad \text{rms - root-mean-square speed}$$

- As temperature increases, average molecular speed increases

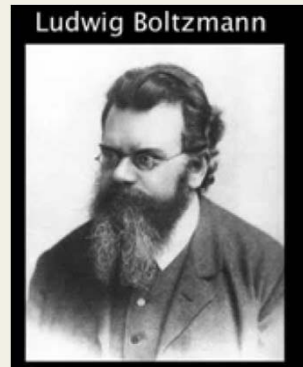
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- At a given temperature, gas molecules in a sample can be characterized by an **average speed**
  - Some gas molecules move faster than average, some move slower than average
  - The distribution function that describes the speeds of a collection of gas particles is known as the **Maxwell-Boltzmann distribution** of speeds

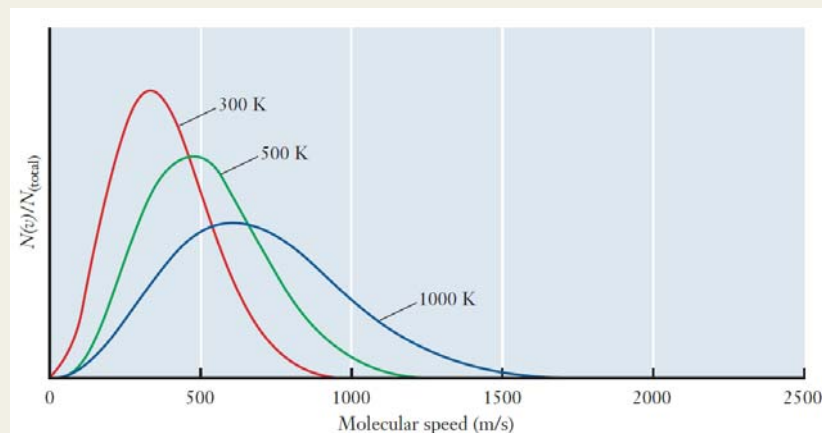


James Clerk Maxwell

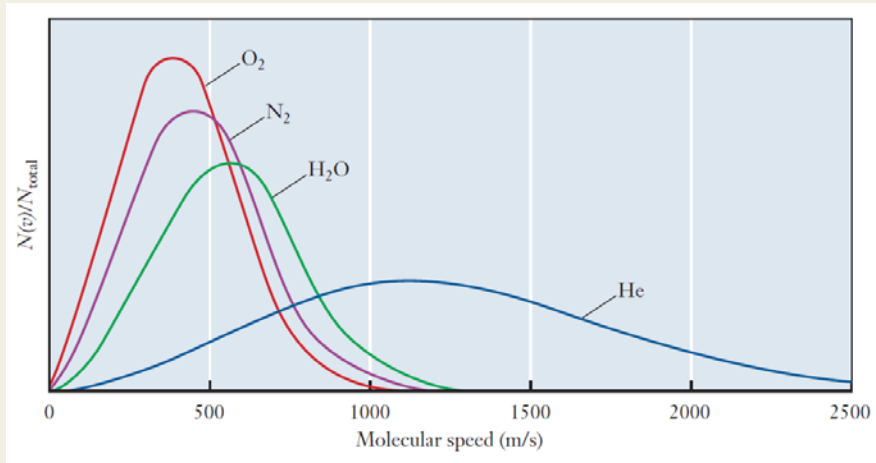


Ludwig Boltzmann

- **As temperature increases**, the fraction of molecules moving at higher speeds increases



- As molecular mass decreases, the fraction of the molecules moving at higher speed increases



- Most gas molecules move at the **most probable speed ( $v_{\text{mp}}$ )**, which is the peak of the curve in the Maxwell-Boltzmann plot

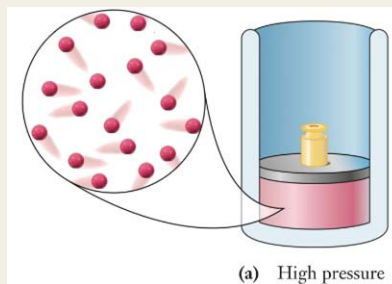
$$v_{\text{mp}} = \sqrt{\frac{2RT}{M}}$$

- 1) M - molar mass of the gas
- 2) R – Gas constant
- 3) T – Absolute temperature

## Limitations of the Kinetic Theory



- 1) **According to kinetic theory**, particles/molecules in a gas are infinitely small and they occupy no space
  - Under high pressure conditions,
    - ✓ Gas molecules are squeezed very close together
    - ✓ The average distance between molecules decreases
    - ✓ Therefore, the volume of the gas particles becomes significant



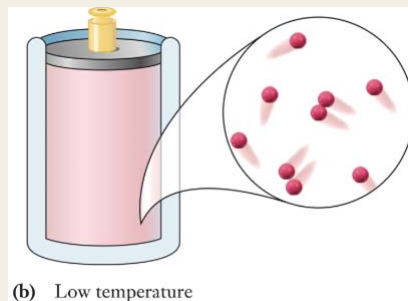
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## Limitations of the Kinetic Theory



- 2) **According to Kinetic molecular theory**, gas molecules move in straight lines and interact only through perfectly elastic collisions
  - At **low temperature**
    - Kinetic energy decreases (with temperature)
    - Gas molecules experience sticky collisions
    - Therefore, attractive and repulsive forces are significant at low temperature

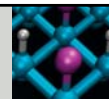


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## Correcting the Ideal Gas Equation



- **Van der Waals equation** is commonly used to describe the behavior of real gases

$$\left( P + \frac{an^2}{V^2} \right) (V - nb) = nRT$$

$a$ ,  $b$  – van der Waals constants

- $a$  corrects for **attractive forces** (attraction between gas particles)
  - Molecules with stronger attractive forces have larger  $a$  values
- $b$  corrects for the **volume occupied by gas molecules**
  - Large molecules have larger  $b$  values

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## Correcting the Ideal Gas Equation



**Table 5.2** Van der Waals constants

Gas	$a$ (atm L <sup>2</sup> mol <sup>-2</sup> )	$b$ (L mol <sup>-1</sup> )
Ammonia, NH <sub>3</sub>	4.170	0.03707
Argon, Ar	1.345	0.03219
Carbon dioxide, CO <sub>2</sub>	3.592	0.04267
Helium, He	0.034	0.0237
Hydrogen, H <sub>2</sub>	0.2444	0.02661
Hydrogen fluoride, HF	9.433	0.0739
Methane, CH <sub>4</sub>	2.253	0.04278
Nitrogen, N <sub>2</sub>	1.390	0.03913
Oxygen, O <sub>2</sub>	1.360	0.03183
Sulfur dioxide, SO <sub>2</sub>	6.714	0.05636
Water, H <sub>2</sub> O	5.464	0.03049

- The van der Waals constants  $a$  and  $b$  are **gas specific**
  - Both are zero in gases behaving ideally

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