

CH301 Unit 1

EXAM REVIEW: GAS LAWS AND FUNDAMENTALS

Unit One Learning Outcomes: Gases

Learning Outcomes for Gases

Students will be able to...

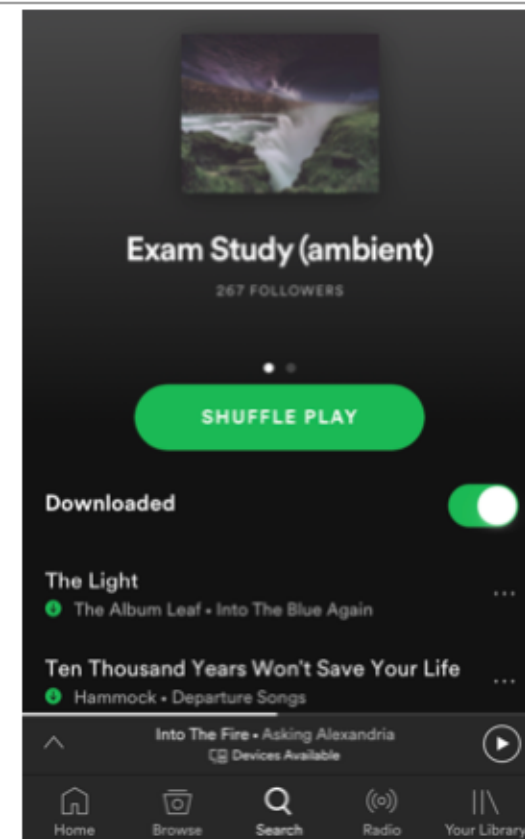
1. Describe pressure from a macroscopic and microscopic perspective.
2. Relate Boyles, Charles', and Avogadro's gas laws to observations of gas behavior.
3. Calculate the values for state functions (n , V , T , P) using the ideal gas equation.
4. Define the conditions of STP and SATP.
5. Relate the number density and mass density for a given gas, including quantitative calculations such as mass, molecular weight, and density.
6. Relate partial pressures and the total pressure as described by Dalton's Law of Partial Pressure.
7. Apply the concept of the gas laws to gas phase reactions.
8. Perform stoichiometric calculations using gas properties, masses, moles, limiting reagents, and percent yield.
9. Relate kinetic energy to the temperature of a gas.
10. Relate temperature, molar mass, and gas velocity.
11. Describe the effect of molar mass and temperature on the Maxwell-Boltzmann gas velocity distribution.
12. Apply kinetic molecular theory to a variety of gas phenomena including diffusion and effusion.
13. Calculate relative effusion and diffusion rates using Graham's Law.
14. Explain the quantitative relationship between state functions (n , T , V , and P) as described by kinetic molecular theory.
15. Describe macroscopic gas behavior using a small particle model of a gas.
16. State when the ideal gas model fails to predict the behavior of gases observed in nature and in the laboratory.
17. Explain what the breakdown of the ideal gas law reveals about the assumptions of kinetic molecular theory.
18. Explain the general principles of the hard sphere model and van der Waal's model of gas.

Don't forget your fundamentals:

- Balancing reactions
- Reaction Stoichiometry
- Composition stoichiometry (including molar mass conversions, Avogadro's number, mass percent, etc.)
- Unit Conversions
- Diatomics
- Prefix Conversion
- Etc.

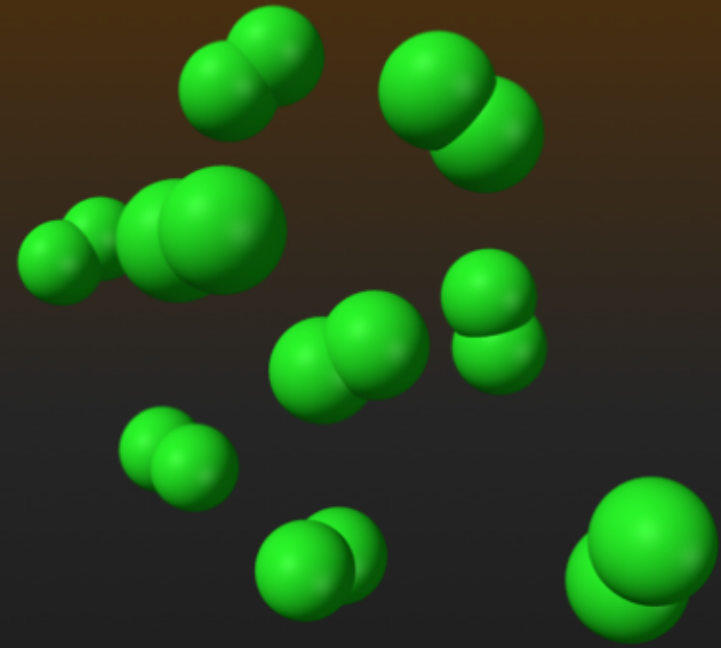
Wadman Study (Cram) Schedule

- 1. 6:00- 7:30 pm** Get a good meal, organize your study materials, take a deep breath
- 2. 7:30- 9:30 pm** Create a thorough outline of the topics covered in class
 - Use lecture notes, g-chem learning outcomes, and review slides to create a comprehensive study guide
 - The goal here is twofold: create a great study guide for the rest of the night and to create a meaningful study experience
- 3. 9:30-11:30 pm** Work through practice problems without looking through the solutions; if you need help, use your outline. If your outline doesn't help, add to it!
- 4. 11:30 pm-12:00 am** Cut yourself off from chemistry, wind down, get a good night's sleep. Your memory is much better when you get good sleep before the exam.



Study playlist: Jimmy Wadman on Spotify

Non-Ideal Gas Law



The Non-Ideal Gas Law

- Based on our ideal gas model, we made two assumptions that are tragically wrong for certain gases at certain conditions:

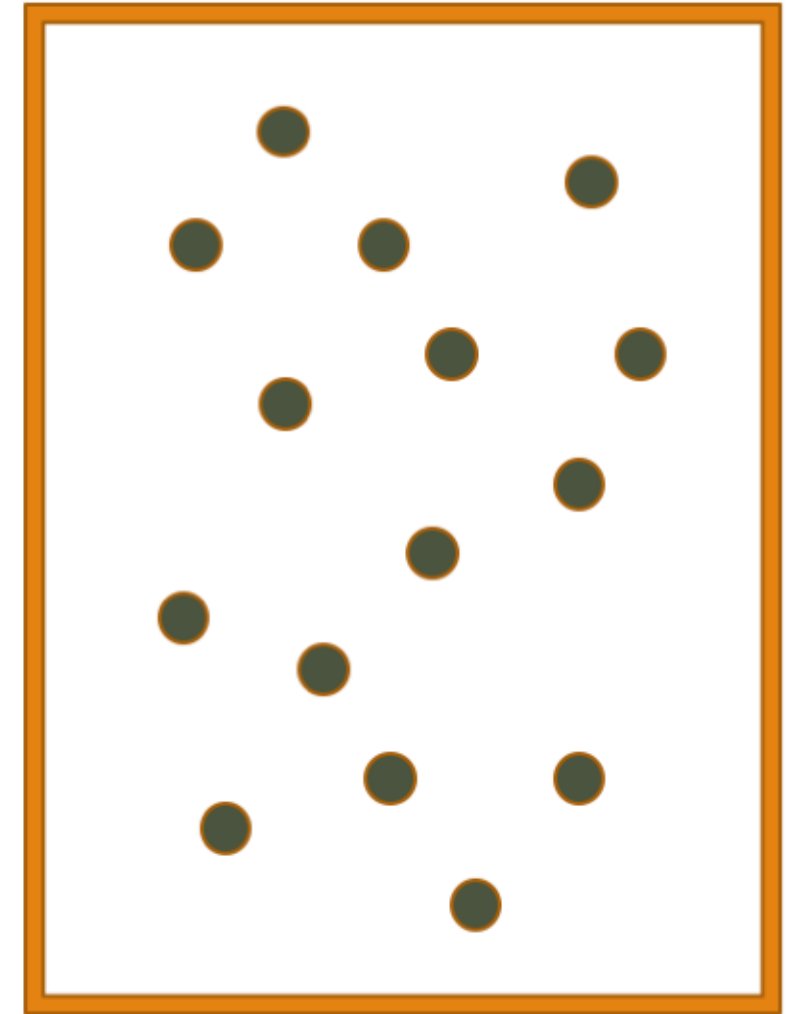
- Gas molecules have no volume**
- Gas molecules have no attractions or repulsions**

- With certain gases at certain conditions, you can get a situation in a lab where:

$$P_{measured} V_{measured} \neq nRT \quad \checkmark$$

- The way to “fix” this is by adding correction factors to account for the volume of gas molecules and the attractions between them. This equation looks like this:

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



The Non-Ideal Gas Law

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

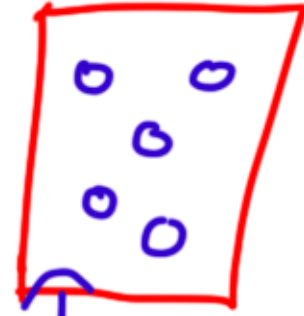
P = measured pressure
 $P < P_{\text{ideal}}$
 $P_{\text{ideal}} = P + \frac{an^2}{V^2}$
 $a \propto$ attractions

$z < 1$
 $PV < nRT$
attractions dominate
 $\frac{PV}{nRT} < 1$

V = volume of the container
 $V_{\text{available}} < V$
 $\hookrightarrow V - nb$
 nb = volume of particles

$b \propto$ repulsions
 $z > 1$
 $PV > nRT$
 $\frac{PV}{nRT} > 1$

The Non-Ideal Gas Law



- The Ideal Gas Law is adjusted for real gases in the following way:

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

- Important Notes:
 - The pressure expression $\left(P + \frac{an^2}{V^2}\right)$ is equal to the **ideal pressure**. The P term is the **measured pressure**.
 - The volume expression $(V - nb)$ is equal to the **available volume**. The V term is the **volume of the container**.
- Key points:
 - Attractive forces decrease the pressure. **Stronger attractive forces give a larger "a" value**.
 - The size of the molecules decrease the available volume. **Larger molecules give a larger "b"**.
 - A larger b value correlates with more repulsions. Repulsions are an effect of the fact that two gas molecules can't take up the same space – therefore, they repel each other

Favoring Non-Ideal Vs. Ideal Conditions

Ideal Conditions

- High Temperature
- Low Pressure
- High Volume
- Small Molecules
- No attractive forces
- Small number of molecules

→ more E_k
→ Less interactions
→ $V \gg V_{particles}$

Non-Ideal Conditions

- Low Temperature
- High Pressure
- Low Volume
- Large Molecules
- High Attractive Forces (Intermolecular forces)
- Large number of molecules

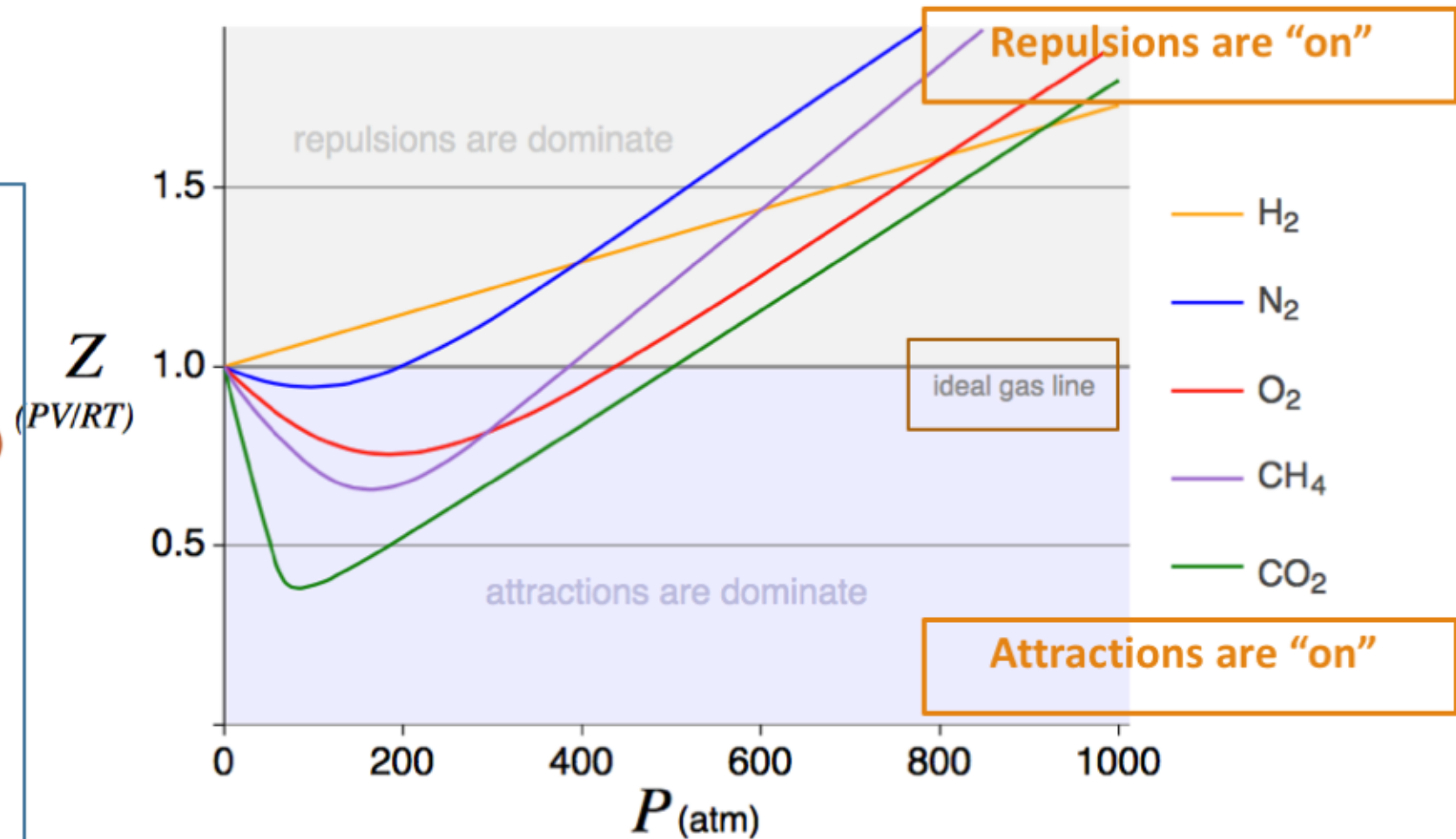
→ "sticky" molecules

Quantifying Non Ideality: Compressibility, z

$$z = \frac{P_{\text{measured}} V_{\text{measured}}}{nRT}$$

Know this:

- When $P_{\text{measured}} V_{\text{measured}} > nRT$,
 - $z > 1$
 - **Repulsions Dominate**
 - **Measured volume is larger than the ideal (available) volume.**
 - b term correlates with repulsions
- When $P_{\text{measured}} V_{\text{measured}} < nRT$,
 - $z < 1$
 - **Attractions dominate**
 - **Measured pressures is less than the ideal pressure.**
 - a term correlates with attractions



Quantifying Non Ideality: Compressibility, Z

For a non ideal gas, $P_{measured} V_{measured} \neq nRT$. Compressibility is essentially the “why” of non ideal gases.

- A gas can be non ideal either because repulsions dominate or because attractions dominate. Compressibility helps us quantify this.

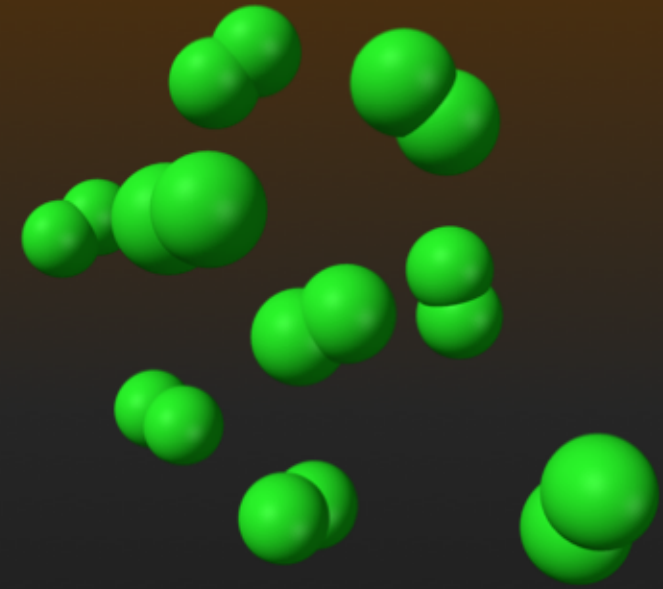
When $P_{measured} V_{measured} < nRT$, **$Z < 1$, Attractions dominate**

- **Measured pressure is less than the ideal pressure.**
- **Caused by the $\frac{an^2}{V^2}$ term**

$$Z = \frac{P_{measured} V_{measured}}{nRT}$$

When $P_{measured} V_{measured} > nRT$, **$Z > 1$, Repulsions Dominate**

- **Measured volume is larger than the ideal (available) volume.**
- **Caused by the nb term**



Quick Fundamentals Review

Chemistry Fundamentals

There is a simple relationship between mass, moles, and molar mass you must use on the exam

moles = mass ÷ MM

- If you have 50 g N₂, you have 1.8 moles
- $50gN_2 \cdot \frac{1 \text{ mol}}{28gN_2} = 1.8 \text{ moles}$

Molar Mass = mass ÷ moles

- If you have a 1.8 moles of a compound, X, with a total mass of 50g, the molar mass is 28g and the identity is N₂
- $\frac{50g X}{1.8molX} = 28 \text{ g/mol}$

mass = mol · MM

- If you have 1.8 moles of N₂, your total mass is 50g
- $1.8 \text{ moles } N_2 \cdot \frac{28 \text{ g}N_2}{1 \text{ mol}} = 50 \text{ g}$

Avogadro's Number

Avogadro's Number connects the macro world of moles and experimental chemistry with the micro world of atoms and molecules:

1. From Molecules to Moles

- Remember: Avogadro's number ($N_A = 6.022 \times 10^{23} \text{ mol}^{-1}$) represents the number of molecules in a mole.

$$\frac{\# \text{ molecules}}{N_A} = \# \text{ moles}$$

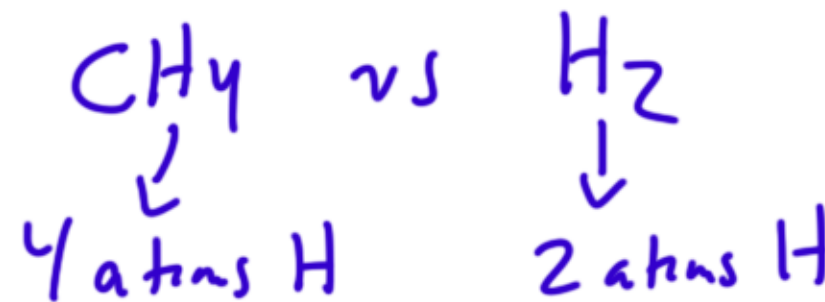
2. From Moles to Molecules

- On the exam, what if you forget if you multiply or divide by Avogadro's number? Remember, a mole is a "packet" of molecules. Converting from moles to molecules should always give you a **bigger number**

$$\# \text{ moles} \cdot N_A = \# \text{ molecules}$$

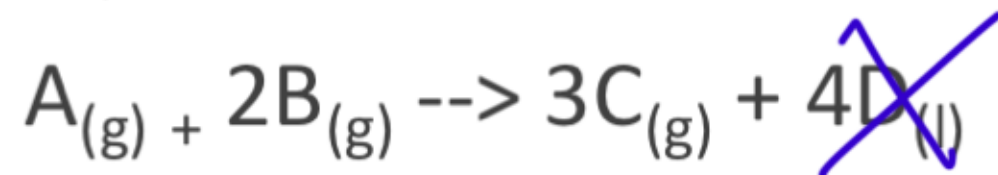
3. Molecules to atoms

A molecule is a group of atoms connected by covalent bonds. CH_4 has 4 hydrogen atoms in it. Hydrogen gas has 2 hydrogen atoms in it. 1 mole of hydrogen gas has 1.204×10^{24} hydrogen atoms in it.



Chemistry Fundamentals: Stoichiometry

- Given the following generic, balanced chemical reaction:



- Simple ratio problems:** how many moles of B are required to make 7 moles of C?
 - This problem can be solved by simply understanding the stoichiometric ratio

$$7 \text{ moles } C \times \frac{2 \text{ moles } B}{3 \text{ moles } C} = 4.67 \text{ moles } B$$

- Limiting Reagent, Gas Law Problems:** what is the total volume of your system when 10 L of A are reacted with 10 L B. Assume this reaction goes to completion.
 - This problem should be solved by working out the reaction

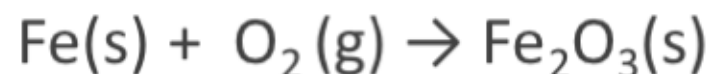
Ans: 20L (15L products formed, 5L reactant in excess)

***Do not include condensed states (liquids, solids) in the volume or pressure of the final solution**

Answer/work in review one

Common Stoichiometry Problems

Given the following unbalanced reaction:



Easy: How many moles of Fe_2O_3 are produced when:

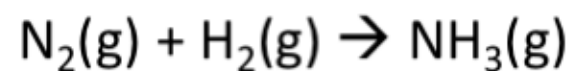
1. 4 moles of iron are reacted with 3 moles of oxygen?
2. 8 moles of iron are reacted with 6 moles of oxygen?
3. 2 moles of iron are reacted with 1.5 moles of oxygen?

Medium: How many moles of oxygen are required to form 5 moles of iron oxide?

Difficult: If you have 5 moles of Fe and 10 moles of O_2 , identify the limiting reagent and the number of moles of your iron (III) oxide product.

Exam-Style Question

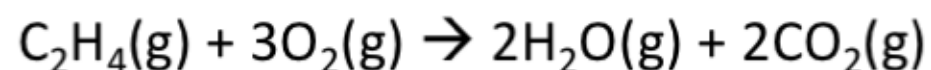
Calculate the mass of NH_3 that can be produced from 30.0g of N_2 and 5.0g H_2 in the following unbalanced chemical reaction:



1. 42.5 g
2. 18.2 g
3. 36.4 g
4. 28.3 g
5. 63.8 g

Exam-Style Question

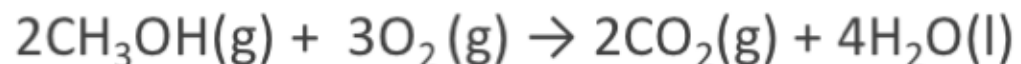
A 6.00 mole sample of C_2H_4 at 1.00 atm and 400K is treated with 6.00 moles of oxygen gas at the same temperature and pressure to form carbon dioxide gas and gaseous water. If the reaction goes to completion, what is the total number of moles of all gases at 1.00 atm and 400K?



1. 2.66 moles
2. 12.00 moles
3. 2.00 moles
4. 8.00 moles
5. 6.00 moles

Stoichiometry Problem 2.0

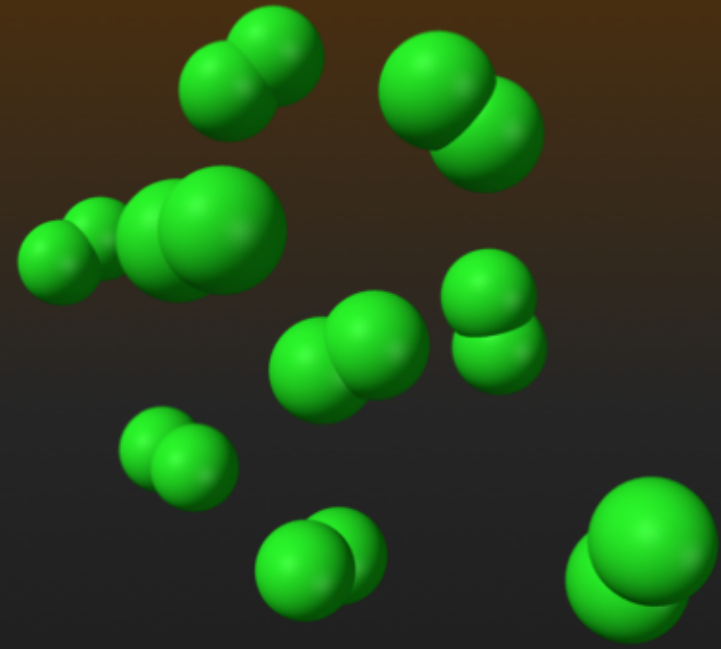
Consider the combustion of methanol at STP.



If 13 L of methanol reacts with 8 L oxygen, what is the number of moles of the reactant in excess in the final reaction mixture?

On your own: What is the total volume of all gases? What is the partial pressure of each gas?

Ideal Gases



$$PV = nRT$$

All Eponymous Gas Laws

Pretty much everything on the exam can be solved using the ideal gas law, but don't forget the individual relationships that make it possible.

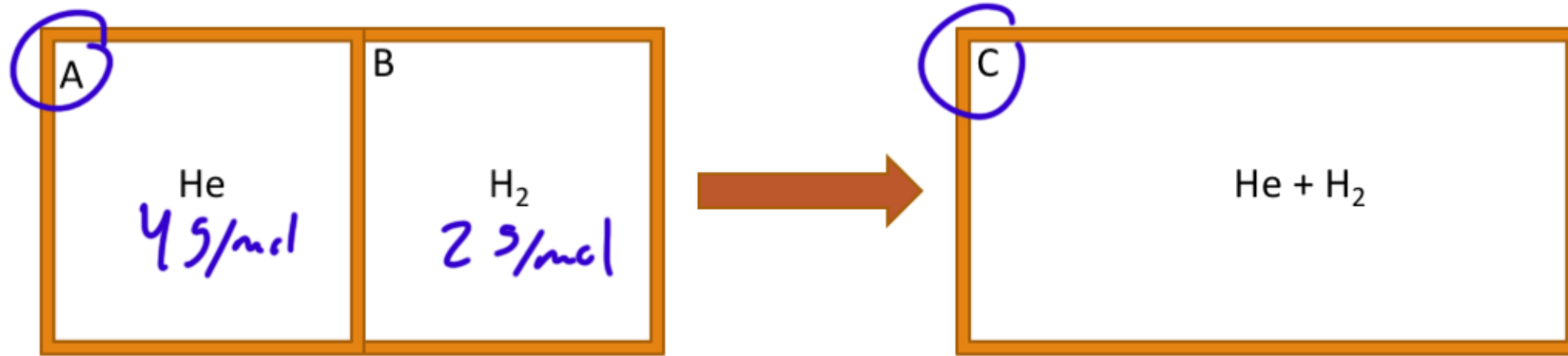
In words

- **Boyle's Law:** Pressure and volume have an inverse relationship
 $P \propto \frac{1}{V}$
- **Charles' Law:** Volume and temperature have a direct relationship
 $V \propto T$
- **Avogadro's Law:** Volume and number of moles have a direct relationship
 $V \propto n$
- **Bike Pump Law:** Pressure and number of moles have a direct relationship
 $P \propto n$
- **Amonton's Law:** Pressure and temperature have a direct relationship
 $P \propto T$
- **Hard-to-do-this Law:** number of moles and temperature have an inverse relationship
 $n \propto \frac{1}{T}$



- Or just remember that given $PV = nRT$:
 - State functions on the same side have an **inverse** relationship
 - State functions on opposite sides have a **direct** relationship

The Ideal Gas Law: Concept Questions



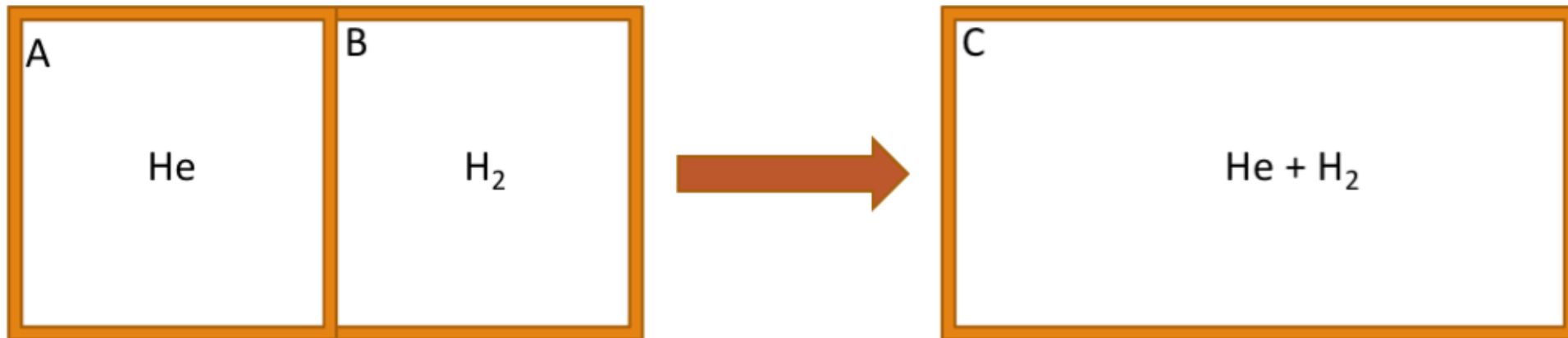
- You have a container that separates 50 L helium (A) and 50 L hydrogen (B) at 1 atm and 298K. If you remove the divider between the two and allow them to mix:
 - What is the final pressure of container C? *1 atm*
 - How does the number density change between A and C? *Same*
 - How does the mass density change between A and C? *decrease*
 - How does the partial pressure of Helium change? *0.5 atm*

$$P_{\text{total}} = P_{\text{He}} + P_{\text{H}_2}$$
$$0.5 + 0.5$$

At home: Now ask yourself some additional questions. What if you change the temperature? What if you add a third gas? Which container has the highest mass density? Which container has the highest kinetic energy? Does container A or B have faster moving particles?

The Ideal Gas Law: Concept Questions

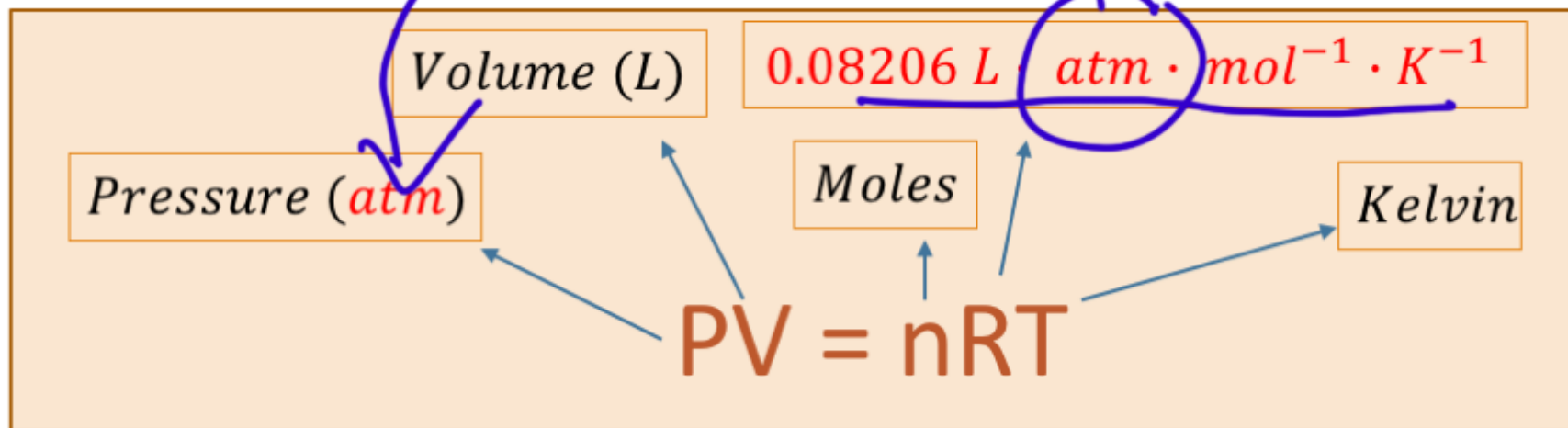
- Getting concept questions correct:
 - Put the various relationships into words – identify direct, indirect relationships
 - If you are not provided with values, you can “test” made up values to see how your numbers change when your conditions change



- You have a container that separates 50L helium (A) and 50L hydrogen (B) at 1atm and 298K. If you remove the divider between the two and all them to mix:
 - What is the final pressure of C? **Same**
 - How does the number density change between A and C? **Same**
 - How does the mass density of the gas mixture change between A and C? **Decreases**
 - **Can you prove your thinking using the ideal gas laws?**

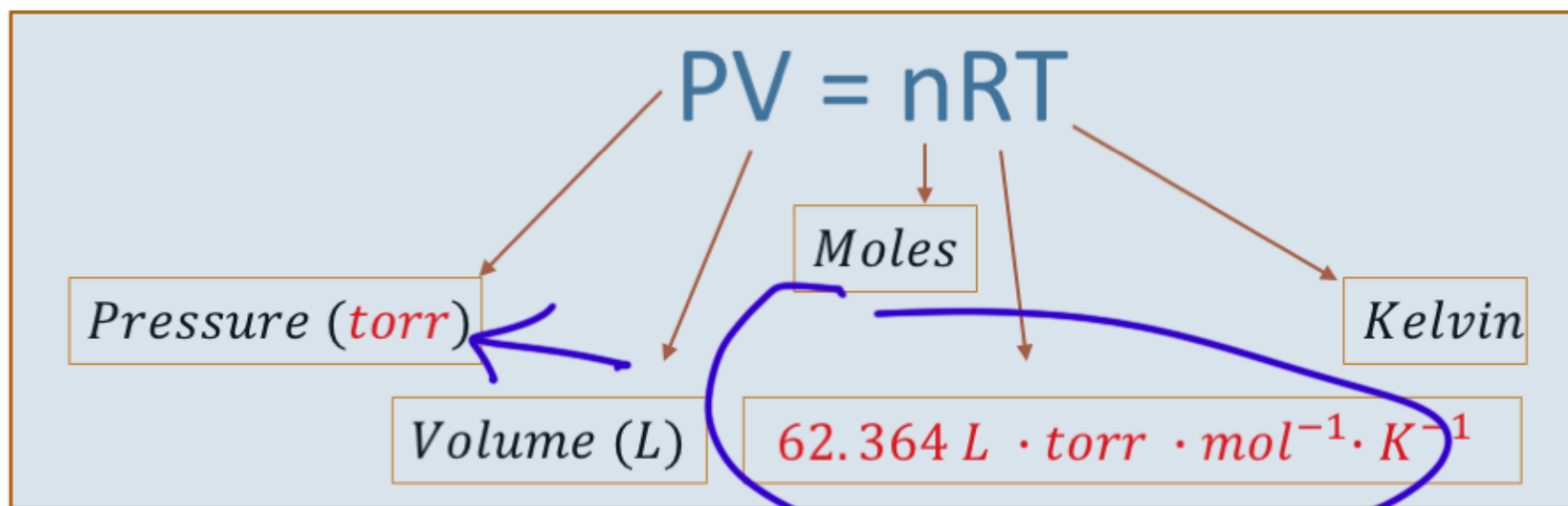
The equation that quantifies the state of a gas; solves for an unknown when the other variables are given

The Ideal Gas Law



What do we see here?

- All units are the same except for pressure and R
- **The R you use for each problem depends on your pressure units**



The two mistakes students make:

- Using the wrong R-value
- **Using Celsius instead of Kelvin**

$$\frac{PV}{nRT} = \frac{PV}{nRT}$$

Standard Conditions (Memorize)

- STP is the default standard for gases
 - 0 degrees Celsius, 1atm
 - **At STP, one mole of ideal gas occupies 22.4L**
- STAP is a “modern” default for gases (less used)
 - 25 degrees Celsius (room temperature), 1 bar

one mole
= 22.4 L

Note: P is pressure, ρ is mass density (m/V)

The Ideal Gas Law and Density

Choose one to memorize:

$$\checkmark \rho = \frac{MM \cdot P}{R \cdot T}$$

Solves directly for mass density (g/L)

$$\checkmark MM = \frac{\rho \cdot R \cdot T}{P}$$

Solves directly for molar mass (g/mol)

$$PV = nRT$$

$$MM \times P = \frac{n}{V} RT \times MM$$

$$MM \times P = \frac{m}{V} RT$$

Note: P is pressure, ρ is mass density (m/V)

The Ideal Gas Law and Density

Choose one to memorize:

$$\rho = \frac{MM \cdot P}{R \cdot T}$$

(Handwritten: ρ is circled in blue, with an arrow pointing to $\frac{g}{L}$)

Solves directly for mass density (g/L)

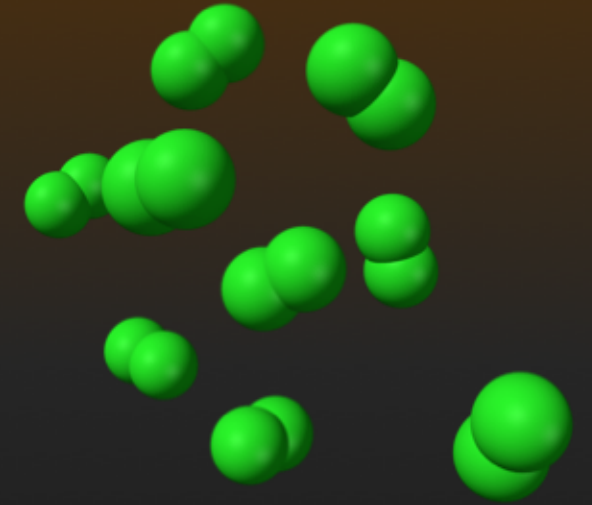
$$MM = \frac{\rho \cdot R \cdot T}{P}$$

(Handwritten: identity with an arrow pointing to the equation)

Solves directly for molecular weight (g/mol)

Two take-away conceptual points:

- All other conditions being equal in the ideal gas law equation (P, V, & T), **all ideal gases have the same number density**; number density depends on P, V, & T but **NOT** molar mass
- All other conditions being equal in the ideal gas law equation (P, V, & T), **a heavier gas will have a greater mass density**; mass density depends on the molecular weight of the gas

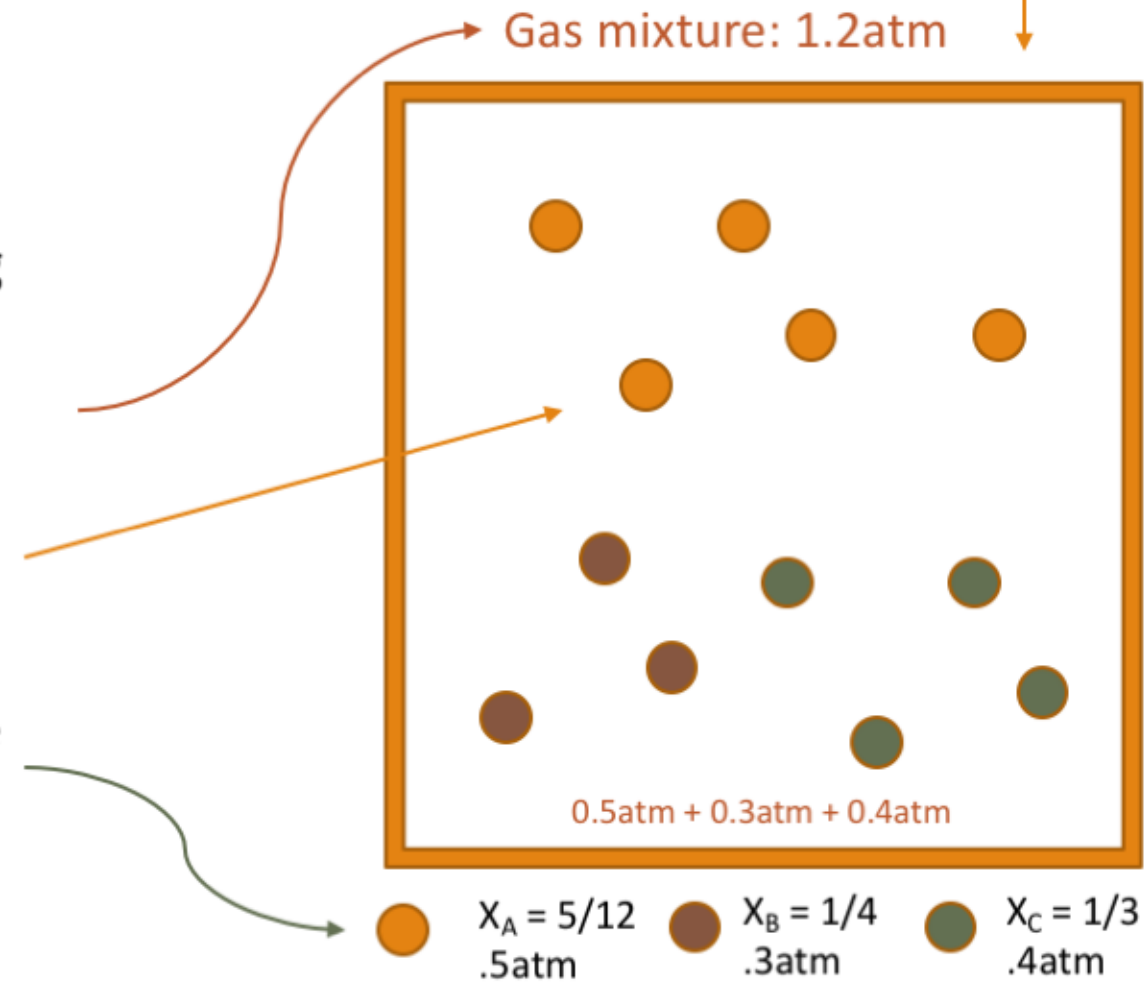
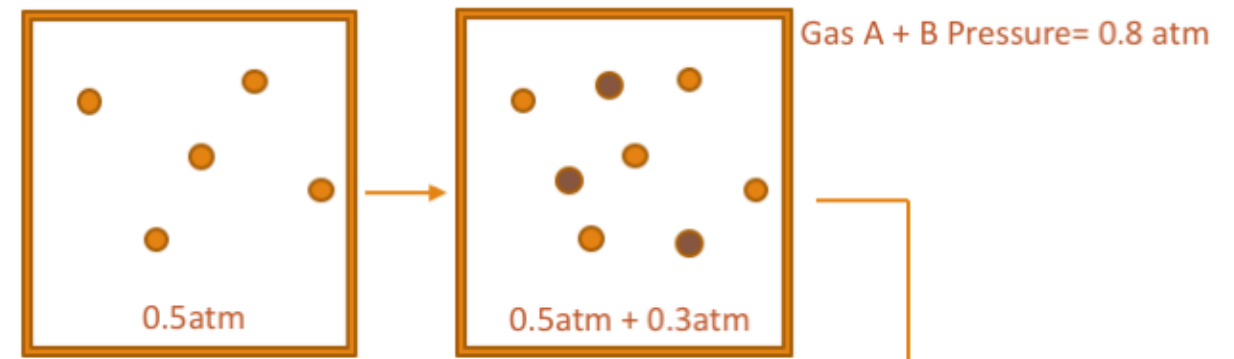


Gas Mixtures

Partial Pressure

Partial pressure is a method for quantifying the pressures exerted by individual species in a gas mixture.

- Three terms are necessary for understanding partial pressures:
 1. **Total pressure** (P_{total}): the “actual” pressure of the system
 2. **Partial pressure** (P_i): the pressure exerted by a single species (i)
 3. **Mole fraction** (X_i): the ratio between the number of moles of a single species (i) and the total number of moles in the system



Dalton's Laws of Partial Pressure

- **Dalton's Law of Partial Pressures**

- The total pressure is equal to the sum of partial pressures

Memorize: $P_{\text{total}} = P_a + P_b + P_c \dots + P_i$

- **Dalton's Law "Restated"**

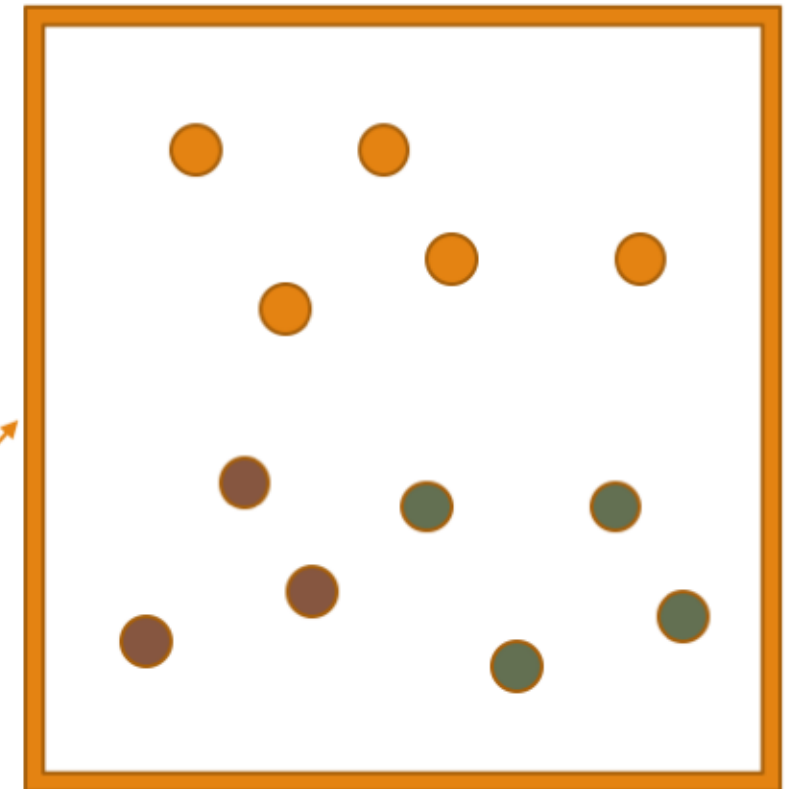
- The partial pressure of a gas species, i , is equal to the total pressure times the mole fraction of i .

Memorize: $P_i = X_i P_{\text{total}}$

$$\frac{P_i}{P_{\text{total}}} = X_i$$

$$\text{Mole fraction } (X_A) = \frac{\text{moles of A}}{\text{moles of A} + \text{moles of B} + \text{moles of C}}$$

Gas mixture: 1.2atm



● $X_A = 5/12$.5atm ● $X_B = 1/4$.3atm ● $X_C = 1/3$.4atm

Partial Pressure Stoichiometry Application

When heated, solid mercury oxide (HgO) will decompose into mercury and oxygen gas according to following equation:



Starting with a container that has only solid HgO in it, the temperature is raised to 700 K and all of the solid decomposes. The total pressure in the container is 0.75 bar. What is the partial pressure of oxygen?

0.75 bar

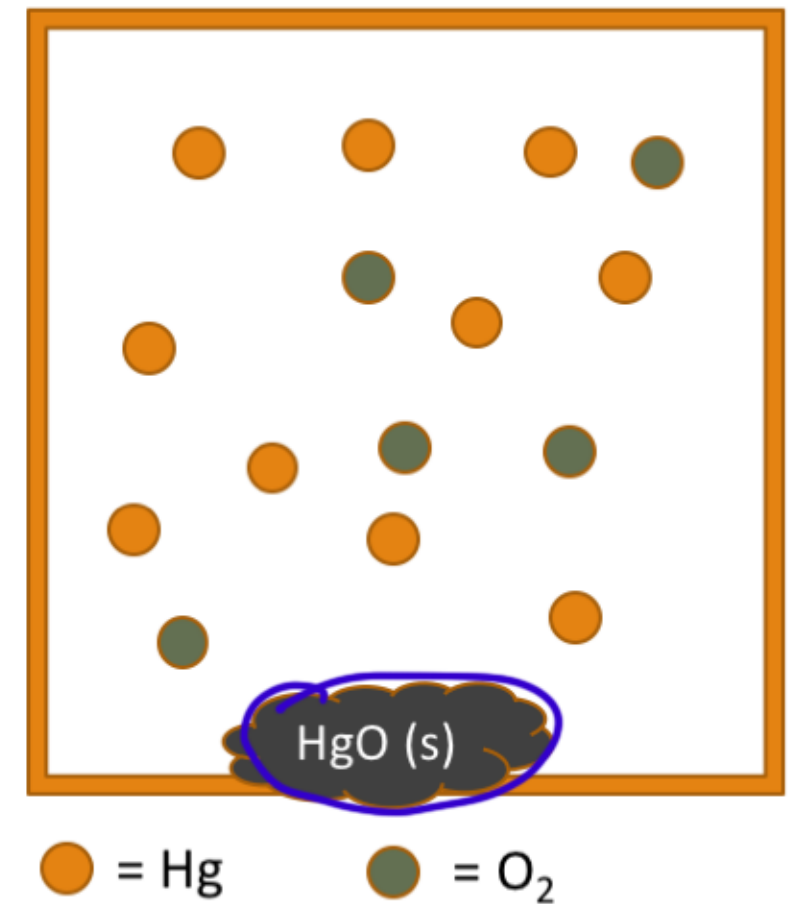
0.50 bar

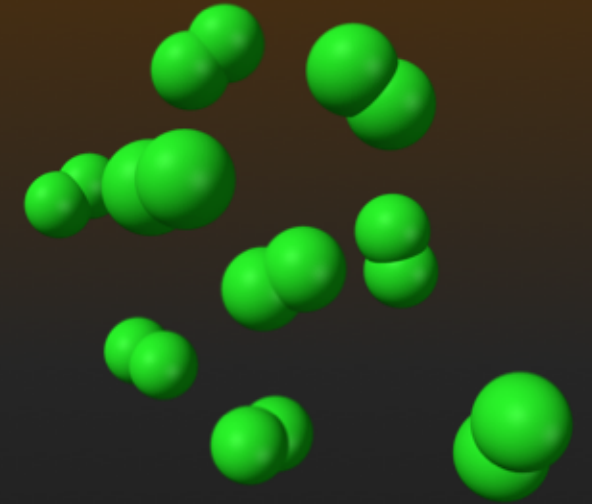
0.125 bar

0.25 bar

$$P_{\text{O}_2} = \chi_{\text{O}_2} P_{\text{total}}$$
$$0.25 = \frac{1}{3} (0.75)$$
$$2\text{Hg} : 1\text{O}_2 \quad \frac{1\text{O}_2}{1\text{O}_2 + 2\text{Hg}}$$

Gas mixture: 0.75 bar





Kinetic Molecular Theory

Kinetic Molecular Theory

Kinetic Molecular Theory explains the “energetic” component of gases, including kinetic energy, velocity, and the relationship between kinetic energy and temperature

1. Gases are constantly moving in random directions
2. The distance between particles is large compared to the particle size
 - True ideal gases have relatively no volume
3. All particles have perfectly elastic collisions
 - There is no energy loss in the system to collisions; energy cannot be created or destroyed based on Newtonian Physics
4. No other forces act upon ideal gases
 - There are no attractive or repulsive forces that act upon ideal gas particles
5. **Kinetic energy depends only on temperature**

We now know that the ideal gas law fails us when these two assumptions do not hold true.

$v_{\text{container}} \gg v_{\text{particle}}$

Kinetic Molecular Theory: Relationships

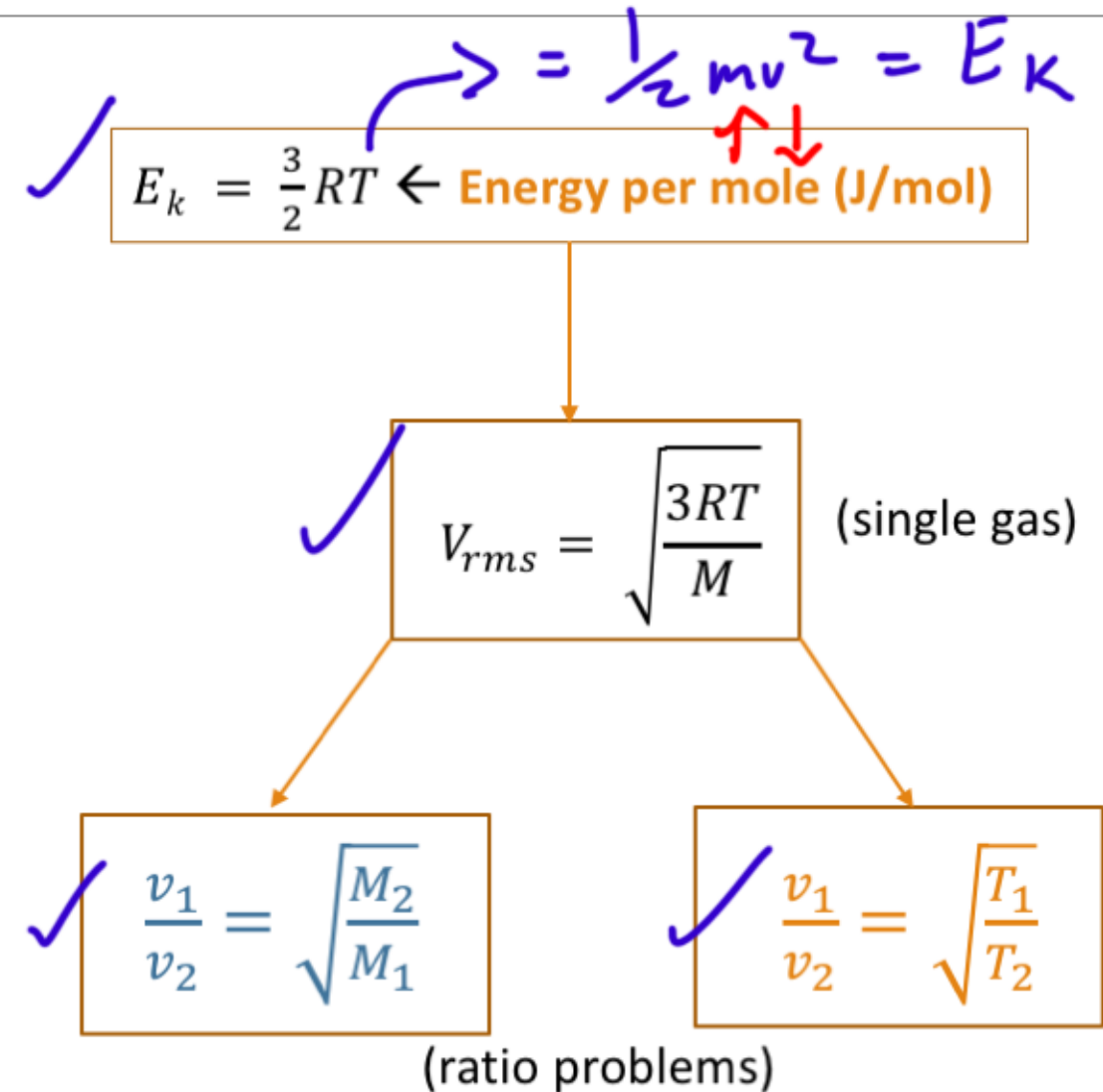
- Kinetic Molecules Theory gives us three key relationships that you should know as equations and by definition (in words)

1. Kinetic Energy vs. Temperature

- Kinetic energy is dependent solely on the temperature of a gaseous system (direct relationship)
- $R = 8.314 \text{ J / mol K}$
- Temperature in Kelvin

2. Mass vs. Velocity

3. Temperature vs. Velocity



KMT: Mass vs. Velocity; Velocity vs. Temperature

Based on the equation:

$$V_{rms} = \sqrt{\frac{3RT}{M}}$$

$R = 8.314 \text{ J/mol K}$

Molar Mass in Kg/mol

We can determine that velocity is proportional to:

1. The **square root of temperature**
 - Gas particles move faster at higher temperatures
2. The **inverse square root** of mass.
 - Gas particles move slower at higher molecular weights

$$V_{rms} \propto \sqrt{T}$$

$$V_{rms} \propto \sqrt{\frac{1}{M}}$$

Pay very close attention to how the 1 and 2 match up in these relationships

KMT: Ratio Problems

When considering two gases, you can use the KMT equation to create a ratio:

$$v_i \propto \sqrt{\frac{1}{m_i}}$$

1. Mass vs. Velocity (V_{rms}) (same temperature, different molecule)

- Velocity is proportional to the **inverse** square root of mass.
- When temperature is constant, lighter particles move faster**

$$\frac{v_1}{v_2} = \sqrt{\frac{M_2}{M_1}} \quad \checkmark$$

2. Velocity (V_{rms}) vs. Temperature (same molecule, different T)

- Velocity is proportional to the square of temperature
- When dealing with the same species gas, particles move faster at higher temperatures**

$$\frac{v_1}{v_2} = \sqrt{\frac{T_1}{T_2}} \quad \checkmark$$

Practice: What is the ratio of the effusion rates of SO₂ to chlorine gas?

Note: v_{rms} is typically measured through effusion
Effusion: a gas particle moves through a small hole

Practice Problem



What is the ratio of the effusion rates of SO_2 to chlorine gas?

$$\frac{v_{SO_2}}{v_{Cl_2}} = \sqrt{\frac{M_{Cl_2}}{M_{SO_2}}}$$

The position of these is important because it is asking you for the ratio of $SO_2 : Cl_2$

Work on your own: If a one mole sample of Cl_2 has a velocity of 500m/s, what is the temperature?

$$V_{Cl_2} = \sqrt{\frac{3RT}{M}}$$

$R = 8.314 \text{ J/mol K}$
 $M = \text{molar mass in kg/mol} = 0.071 \text{ kg/mol}$

Same gas @ different T
different gases @ same T

Maxwell-Boltzmann Distribution

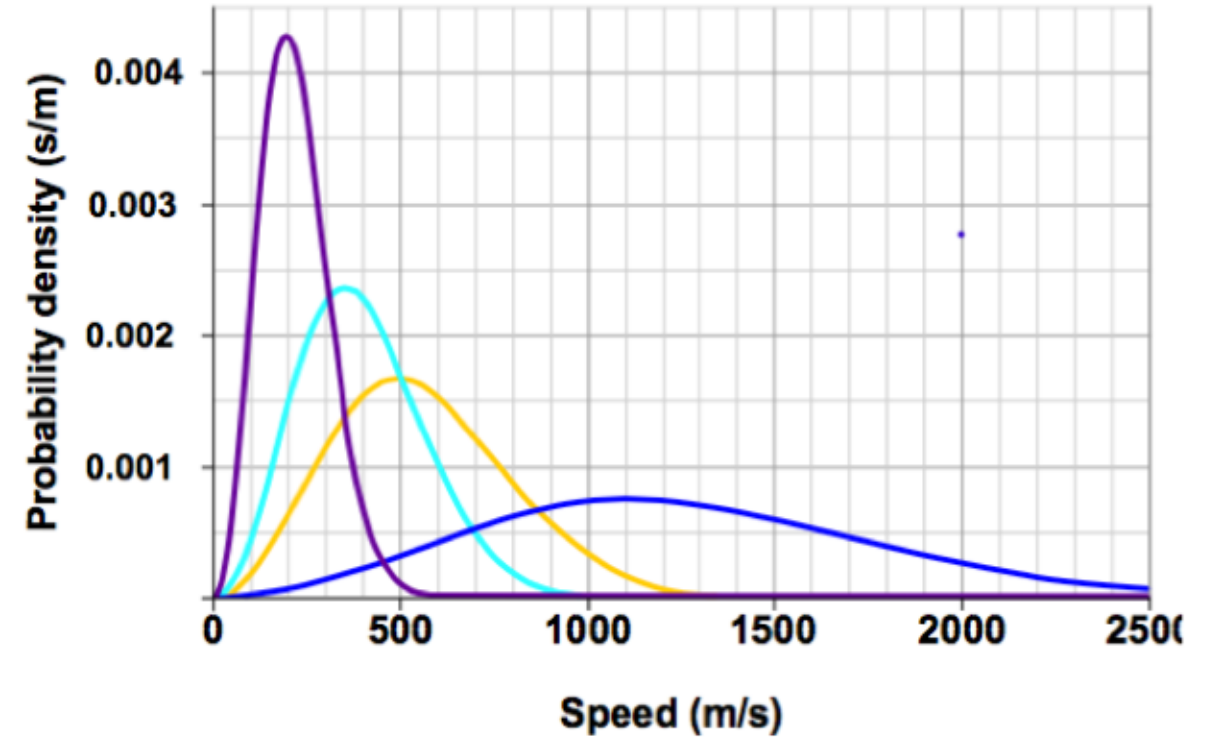
- When you have different gases at a given temperature:

$$v_{rms} \propto \sqrt{\frac{1}{M}}$$

- The heavy gases move slower on average
- The lighter gases move faster on average**

- Checklist:

- Heavier gases have a slower v_{rms} and a narrow distribution
- Lighter gases have a faster v_{rms} and a wider distribution

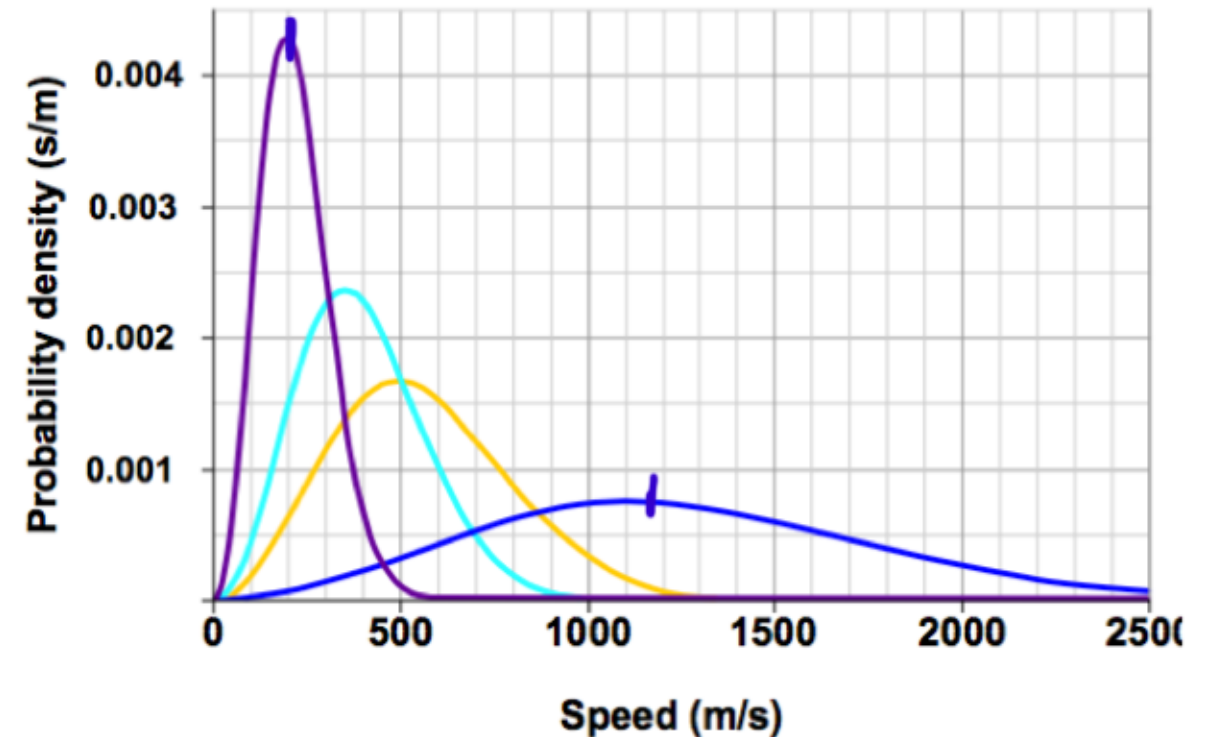


Maxwell-Boltzmann Distribution

When you have the same gas at different temperatures:

$$v_{rms} \propto \sqrt{T}$$

- **The hotter gas sample moves faster on average**
- The colder gas sample moves slower on average
- Checklist:
 - The colder samples will have a slower v_{rms} and a narrow distribution
 - The hotter samples have a faster v_{rms} and a wider distribution



One Last Exam-Style Question

Consider three closed 10L containers at room temperature. Container 1 contains neon gas. Container 2 contains argon gas. Container 3 contains hydrogen gas. Each container has a constant pressure equal to 1.2 atm. Which of the following properties, if any, are the same between the containers?

I. Average velocity of gas particles

II. Number density

III. Kinetic energy $\rightarrow \propto T$ ONLY

IV. Mass density

1. II and III only

2. IV only

3. I and IV only

4. None of the above

5. I, II, III and IV

6. I and III only

7. II and IV only

8. III only

different

$$E_k = \frac{3}{2} RT = \frac{1}{2} mv^2$$

$\uparrow \downarrow$