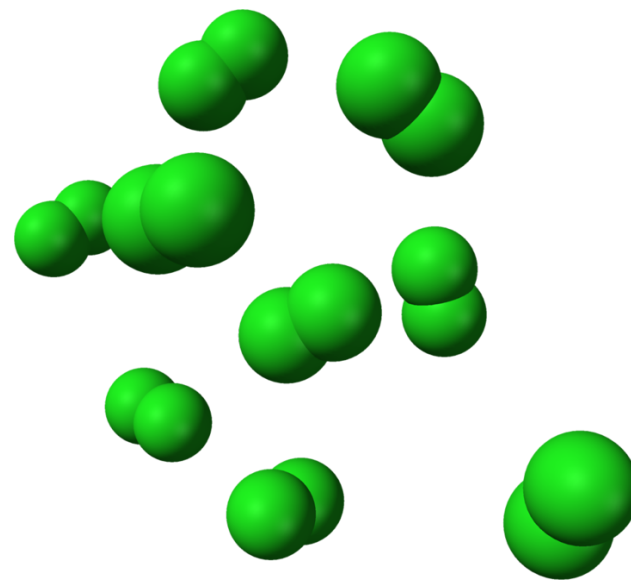


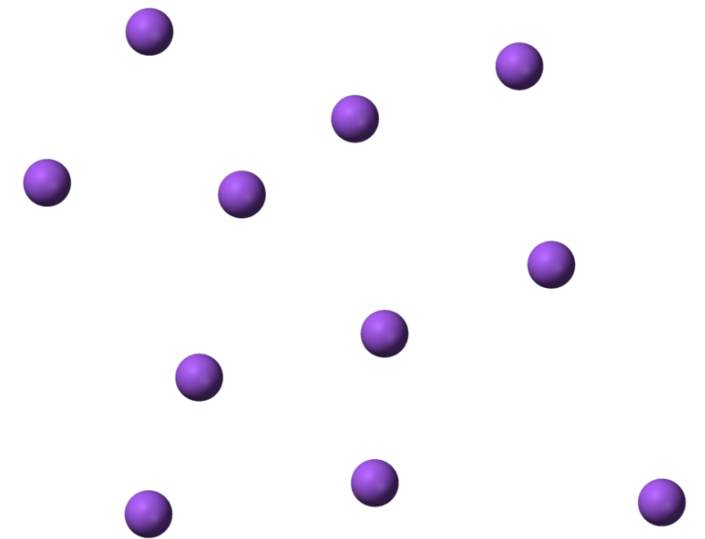
CH301 Unit 1

EXAM REVIEW JIMMY WADMAN



Goals for the Exam Review

- Sit back and try to absorb what you can – everything from today will be posted online
- STUDY TIPS!!!
- Non-Ideal Gases, Compressibility
- Focus on the big picture of the units
- Discuss specific topics tested
- Highlight important question types for the topics tested
- Predict particularly confusing / challenging exam topics
- The **theme** of today: commit yourself to understanding what is actually going on, and you will be much more likely to succeed on the exam



Learning Objectives - Gases

- Describe pressure from a macroscopic and microscopic perspective.
- Relate Boyles, Charles', and Avogadro's gas laws to observations of gas behavior.
- Calculate the values for state functions (n , V , T , P) using the ideal gas equation.
- Define the conditions of STP and SATP.
- Relate the number density and mass density for a given gas, including quantitative calculations such as mass, molecular weight, and density.
- Relate partial pressures and the total pressure as described by Dalton's Law of Partial Pressure.
- Apply the concept of the gas laws to gas phase reactions.
- Perform stoichiometric calculations using gas properties, masses, moles, limiting reagents, and percent yield.
- Relate kinetic energy to the temperature of a gas.
- Relate temperature, molar mass, and gas velocity.
- Describe the effect of molar mass and temperature on the Maxwell-Boltzmann gas velocity distribution.
- Apply kinetic molecular theory to a variety of gas phenomena including diffusion and effusion.
- Calculate relative effusion and diffusion rates using Graham's Law.
- Explain the quantitative relationship between state functions (n , T , V , and P) as described by kinetic molecular theory.
- Describe macroscopic gas behavior using a small particle model of a gas.
- State when the ideal gas model fails to predict the behavior of gases observed in nature and in the laboratory.
- Explain what the breakdown of the ideal gas law reveals about the assumptions of kinetic molecular theory.
- Explain the general principles of the hard sphere model and van der Waal's model of gas.
- Define the compressibility factor (Z) and how its value gives insight into the types of forces in between gas molecules.
- Don't forget your fundamentals!
 - Stoichiometry, mole conversions, etc.

Wadman Study (Cram) Schedule

1. When you get home: create a thorough outline of the topics covered in class
 - If you kept up with the G-chem site and took good notes, make another one with an added dimension (add color coding, highlighting, etc.)
 - Adding **color** increasing the effective time of engagement with the material
2. Work through the blank homework assignments provided and g-chem problem sets **without looking at the solutions** until you feel like you can go to sleep with confidence in the material
3. Designate thirty minutes to an hour before going to sleep (and getting at least 8 hours of sleep!) for doing something relaxing / unrelated to chemistry
4. The day of the exam, continue working through problem sets (old exams?) and referring to your outline for help.
 - If your outline is not sufficient for answering a questions, add to it!
5. Eat well throughout the day of the exam
6. If you do well on the exam, save that outline for the final!

Neuroscience Study Tips

1. Memory consolidation requires sleep. Your studying is only as good as the rest you get before the exam.
2. Green Tea is better than coffee
 - One of the only chemicals actually demonstrated to promote learning is caffeine. However, it is only effective in a range lower than normal consumption (coffee is too much, 1-3 cups of green tea throughout the day is just right).
3. Music can help the learning process
 - In fact, good music is one of the most amazing things in the world for the brain 😊

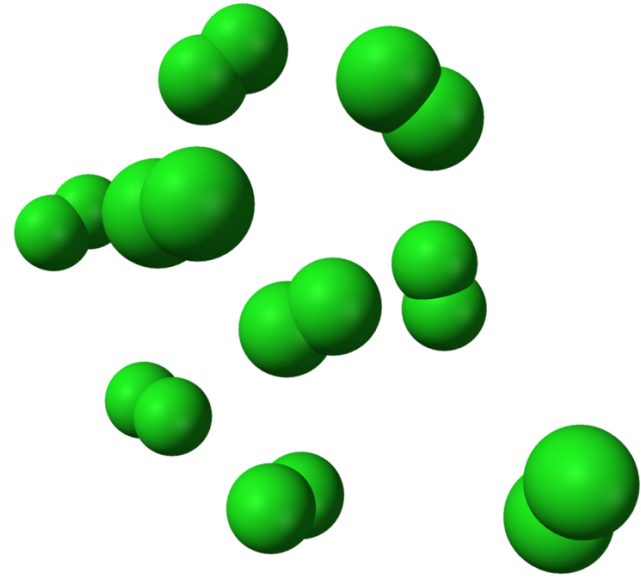


Overview of the Exam

- 25 multiple choice questions (+ free response for MWF classes)
- Four main chapters:
 1. Fundamentals
 2. Ideal Gases
 - Ideal gas law
 - Gas mixtures
 3. Kinetic Molecular Theory
 4. Non-ideal Gases
- No surprises: all questions are straightforward applications of the information provided on the homework, lectures, and the g-chem

Non-Ideal Gases

EXAM REVIEW



The Non-Ideal Gas Law

- Based on our ideal gas model, we made two assumptions that are tragically wrong in certain cases:
 1. Gas molecules have no volume
 2. Gas molecules have no attractions or repulsions

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

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

- 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

- 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
 - More concisely, gases are modeled best **at high temperatures and low pressures.**
 - Attractions increase at lower temperatures (think of water at low temperatures vs. high temperatures)
 - Repulsions increase at high pressures

Favoring Non-Ideal Vs. Ideal Conditions

Ideal Conditions

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

Non-Ideal Conditions

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

Non-Ideal Vs. Ideal: Compressibility

- Suppose you have an ideal gas where:

$$P_{measured} V_{measured} = nRT$$

- You can restate this relationship to say that:

$$z = 1 = \frac{P_{measured} V_{measured}}{nRT}$$

- The value of this ratio, Z, is known as the **Compressibility Factor** and deviates from 1 when your gas has attractive or repulsive forces (when $P_{measured} V_{measured} \neq nRT$)

- Key points:

- When $P_{measured} V_{measured} > P_{ideal} V_{ideal}$, **Z will be greater than 1**
 - Repulsions Dominate
 - Always increases with pressure (b dominates)
- When $P_{measured} V_{measured} < P_{ideal} V_{ideal}$, **Z will be less than 1**
 - Attractions dominate
 - Decreases with pressure (a dominates)

Compressibility Notes

if $PV = nRT$

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

if $PV \neq nRT$

$$Z > \frac{PV}{nRT}$$

Pressure
is lower
than "ideal"
attractions
dominate

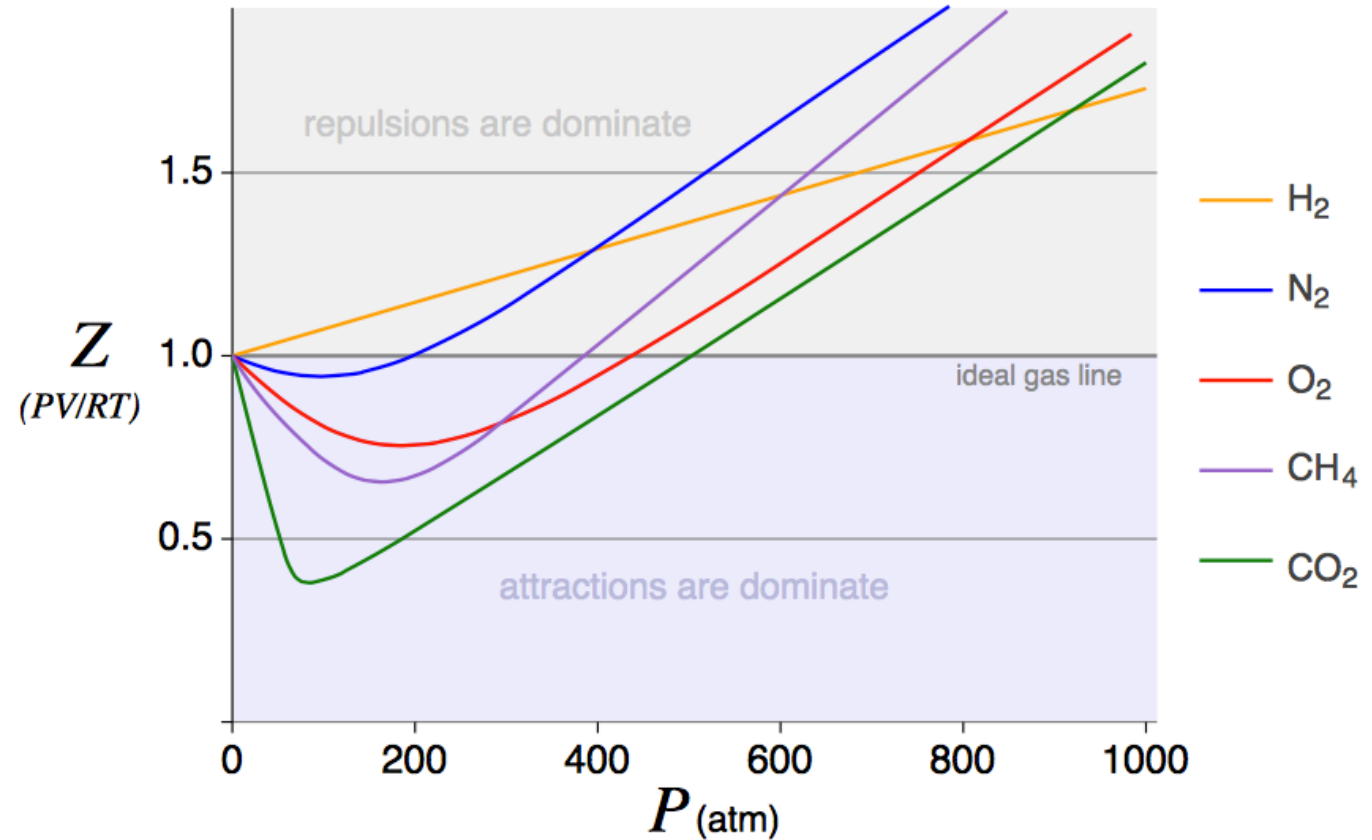
$$Z < \frac{PV}{nRT}$$

Volume is larger
than the "ideal"
repulsions
dominate

Compressibility for Visual Learners

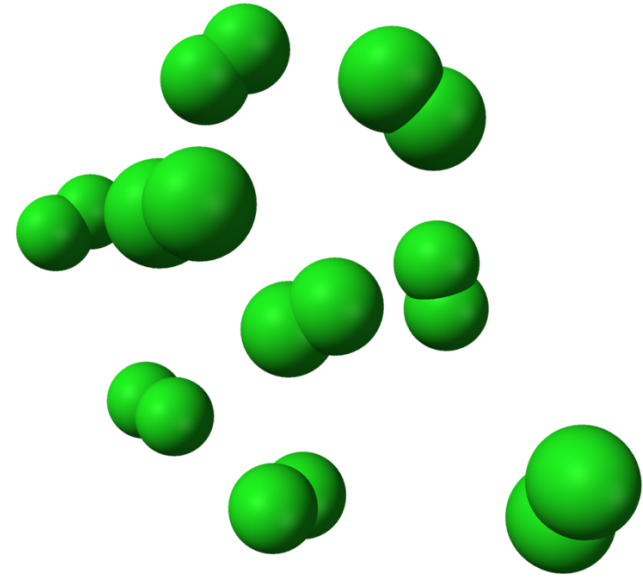
Repulsions Dominate

Attractions Dominate



Fundamentals

EXAM REVIEW



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
- Limiting Reagent, Gas Law Problems:** what is the total volume of your system when 10L of A are reacted with 10L B. Assume this reaction goes to completion.
 - This problem should be solved by working out the reaction
 - *Do not include condensed states (liquids, solids) in the volume or pressure of the final solution**

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 10L of A are reacted with 10L B. Assume this reaction goes to completion.

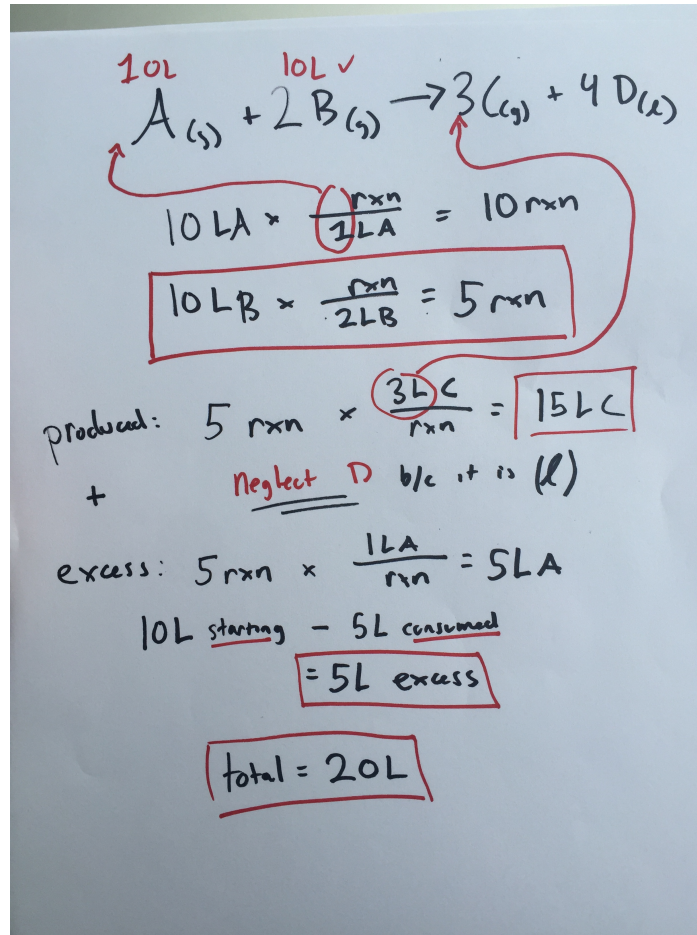
- This problem should be solved by working out the reaction

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**

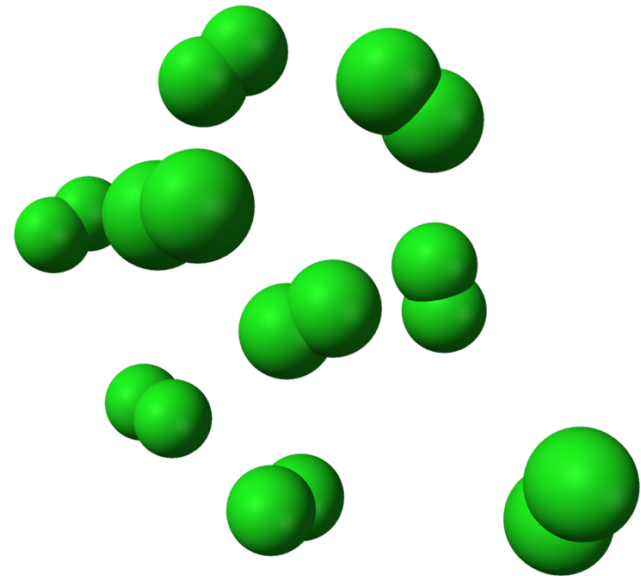
For the exam, this you should be able to do these steps plus applying the information. For example, you might have to apply the final volume to an ideal gas law problem or you might have to do a mass to mole conversion, etc.

Stoichiometry Problem



Ideal Gases

EXAM REVIEW



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

The Ideal Gas Law

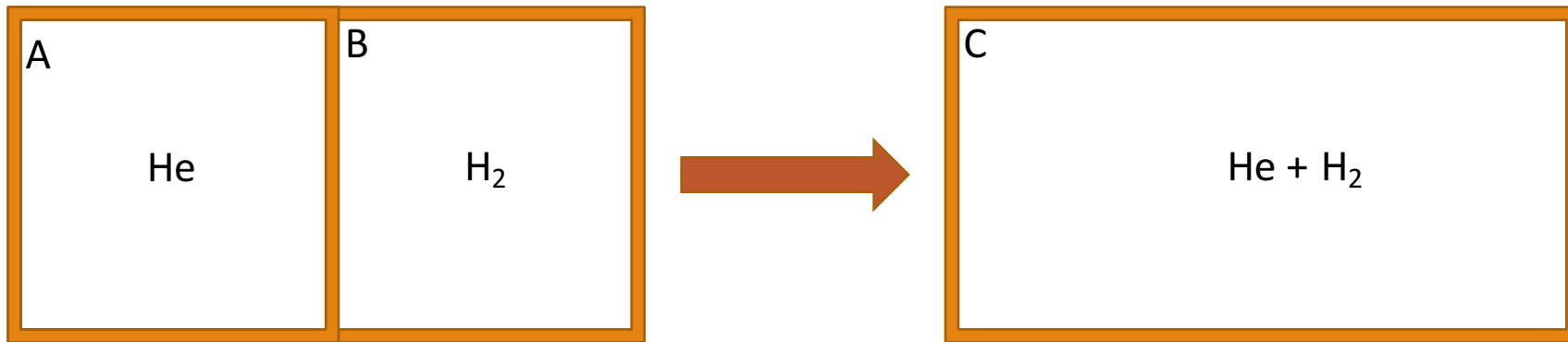
- This exam will test both your ability to “**plug-and-chug**” with the ideal gas law and your ability to apply the law to **conceptual** questions
- The relationship between these state functions is presented in the following equation:

$$PV = nRT$$

- Key points:
 - State functions on the same side have an **inverse** relationship
 - State functions on opposite sides have a **direct** relationship
 - R is a **constant**, not a state function
 - Temperature is always in Kelvin
 - Why? Kelvin can only be expressed as a value 0 or greater. If you used Celsius, you would get negative volumes, pressures, and numbers of moles at certain temperatures

The Ideal Gas Law: Concept Questions

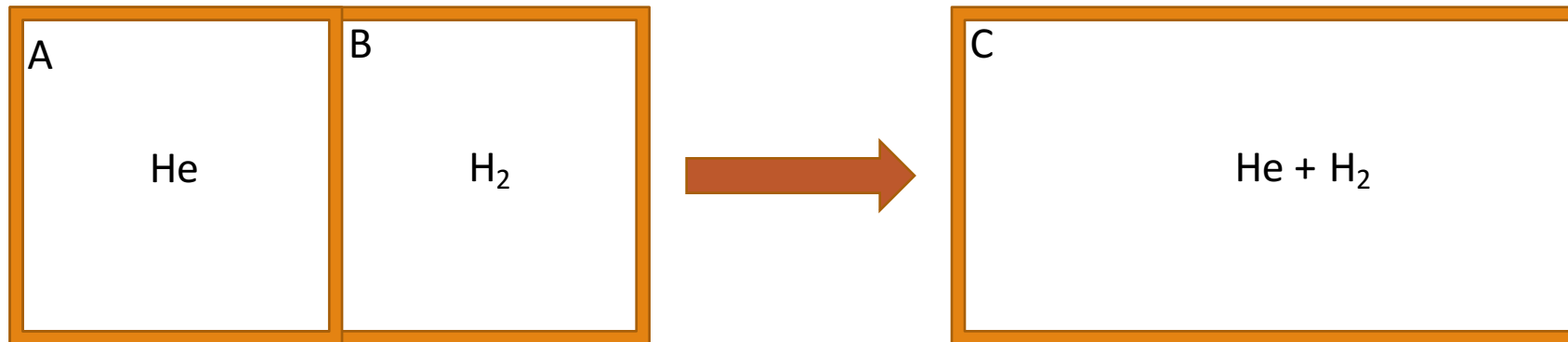
- Getting concept questions correct:
 - Put the various relationships into words – identify direct, indirect relationships
 - If you are not provided with values, you can “test” made up values to see how your numbers change when your conditions change



- 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 allow them to mix:
 - What is the final pressure of C?
 - How does the number density change between A and C?
 - How does the mass density change between A and C?

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 Ideal Gas Law

- Getting quantitative questions correct:
 - Use the correct R value, temperature (in K), and convert any other units as necessary
 - Example: you have 15 moles of Nitrogen gas at STP. What is the change in volume if you increase the temperature by 20 degrees Celsius and double the pressure?
 - Solve for the initial volume
 - Solve for the final volume and take the difference

Ideal Gas Law Answer

$$\begin{aligned}PV &= nRT \rightarrow \text{initial} \\PV &= nRT \rightarrow \text{final} \\(1 \text{ atm})V &= 15 \cdot 0.082 \cdot 273 && \text{or } \underline{\underline{22.4 \text{ L/atm} \times 15}} \\V &= 335 \text{ L} \\(\underline{2 \text{ atm}})V &= \frac{15 \cdot 0.082 \cdot 293}{\underline{2 \text{ atm}}} \\&= 180 \text{ L} \\335 - 180 \text{ L} &= 155 \text{ L}\end{aligned}$$

Individual 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.

- **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

$$P \propto n$$

- **Gay-Lussac 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 and Density

$$\rho = \frac{MW \times P}{RT} \text{ or } MW = \frac{\rho \times RT}{P}$$

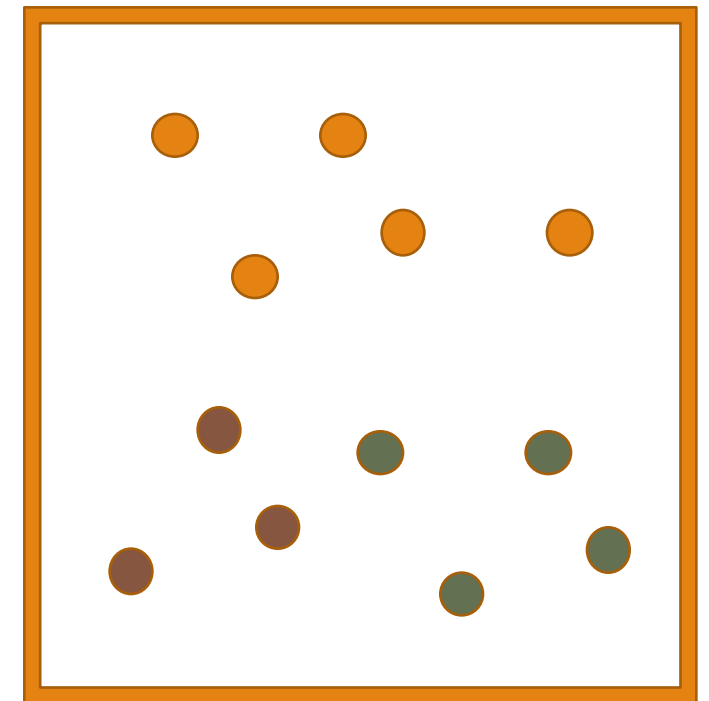
- You can solve for the density of a gas using the molecular weight (or vice versa) with a simple modification of the ideal gas law.
- You can be asked to solve for these values or asked to compare relative densities of gases with differing molecular weights.
- Furthermore, you can be asked conceptual questions that compare number densities and mass densities
 - Number density = n/V
 - Mass density (or just density) = mass/V

Note: Partial Pressures should be easy – the challenge on the exam will most likely be applying partial pressure to a more complex problem (stoichiometry, gas laws, etc.)

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 or 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 Law of Partial Pressures:
 - The total pressure is equal to the sum of partial pressures
 - $P_{\text{total}} = P_a + P_b \dots + P_i$
- Dalton’s Law “Restated:”
 - The partial pressure of a gas species, i , is equal to the total pressure times the mole fraction of i .
 - $P_i = X_i P_{\text{total}}$

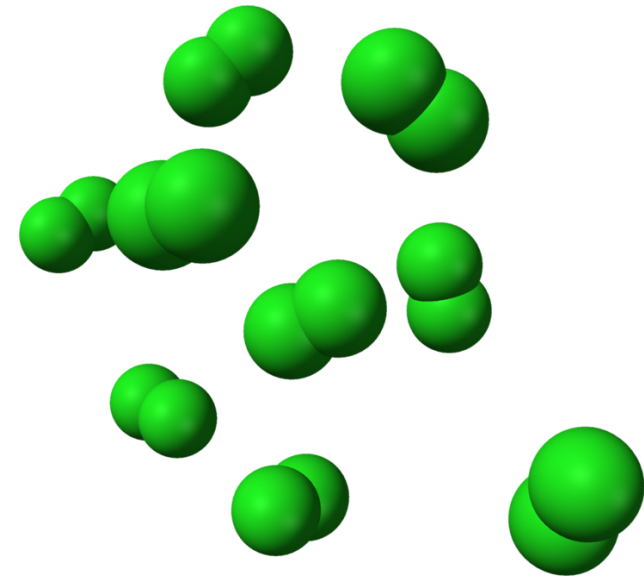
Gas mixture: 1.2atm



● $X = 5/12$.5atm ● $X = 1/4$.3atm ● $X = 1/3$.4atm

Kinetic Molecular Theory

EXAM REVIEW



Kinetic Molecular Theory

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

The ideal gas law fails us when these two assumptions do not hold true. Repulsions are caused by a low **available volume** (b); attractions are caused by significant **intermolecular forces** (a)

Main conclusions: the ideal gas law works because when these pillars of KMT hold true in a system. The ideal gas law is modeled best at **High Temperature** and **Low Pressure**

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

$$E_k = \frac{3}{2}RT \leftarrow \text{Energy per mole (J/mol)}$$

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

$$E_k = \frac{3}{2}kT \leftarrow \text{Energy per molecule (J)}$$

- $k = \frac{R}{N_a}$

2. Mass vs. Velocity

3. Temperature vs. Velocity

Remember that in physics we describe kinetic energy with the equation:

$$\frac{1}{2}mv^2$$

KMT: Mass vs. Velocity; Velocity vs. Temperature

1. Kinetic Energy vs. Temperature

Based on the equation:

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

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

Molecular Weight in Kg

We can determine that velocity is proportional to the **square root of temperature** and the **inverse square root** of mass.

1. Mass vs. Velocity (V_{rms})

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

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

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

2. Velocity (V_{rms}) vs. Temperature

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

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

Practice Problem

What is the ratio of the effusion rates of SO_2 to Cl_2 ?

$$\frac{v_{\text{SO}_2}}{v_{\text{Cl}_2}} = \sqrt{\frac{M_{\text{Cl}_2}}{M_{\text{SO}_2}}}$$

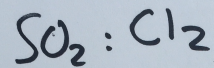
The position of these is important because it is asking you for the ratio of $\text{SO}_2 : \text{Cl}_2$

If the velocity of Cl_2 is 500m/s, what is the temperature?

$$V_{\text{SO}_2} = \sqrt{\frac{3RT}{M}}$$

KMT Answer

What is the ratio of effusion rate of SO_2 to Cl_2 ?



$$\frac{v_{\text{SO}_2}}{v_{\text{Cl}_2}} = \sqrt{\frac{M_{\text{Cl}_2}}{M_{\text{SO}_2}}}$$

↓

$$\sqrt{\frac{71 \text{ g/mol}}{64 \text{ g/mol}}} = 1.05$$

↓
Velocity of a gas escaping a little hole

$$500 \text{ m/s} = \sqrt{\frac{3RT}{M}}$$

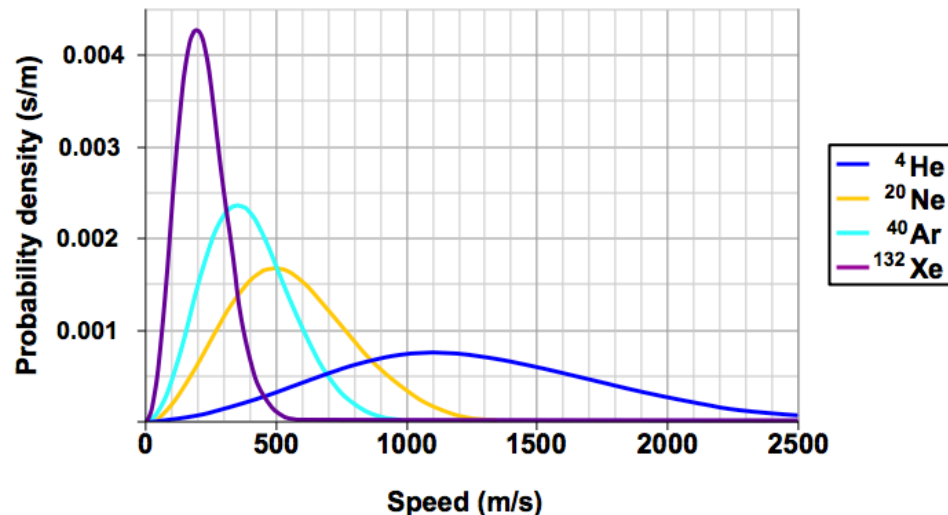
$$250,000 = \frac{3RT}{M}$$

$$250,000 \frac{\text{m}}{\text{s}} \cdot \left(71 \frac{\text{g}}{\text{mol}} \cdot \frac{\text{kg}}{1000 \text{g}} \right) = T$$
$$3 (8.314 \text{ J/mol K})$$

$$= 712 \text{ K}$$

KMT: Mass vs. Velocity; Velocity vs. Temperature

Maxwell-Boltzmann Molecular Speed Distribution for Noble Gases



- Some key features of this graph include:
 - Each curve looks like a unimodal distribution with a “tail” that approaches the limit infinite velocity (0 probability density)
 - Molecules are traveling at a variety of speeds but there is a clear average
 - The actual V_{rms} is slightly to the right of the peak

1. Mass vs. Velocity (V_{rms})

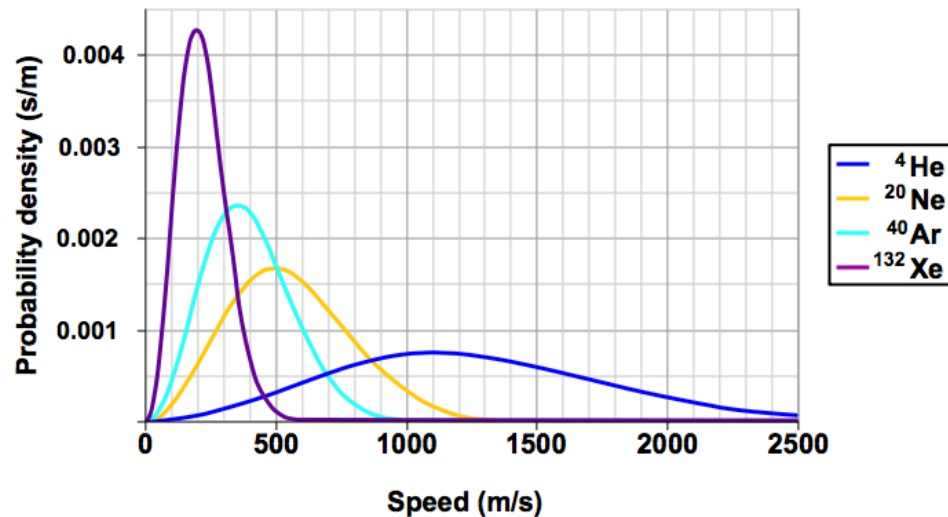
- We can see that the **heavier gases** move **slower** and the **lighter gases** move **faster**
- The **faster** the gas, the **wider** the distribution

2. Velocity (V_{rms}) vs. Temperature

- If you were working with the same gas, a similar graph could be created by modifying temperature instead.
- In this case, higher temperatures of the same gas result in faster speeds.

Maxwell-Boltzmann Distribution

Maxwell-Boltzmann Molecular Speed Distribution for Noble Gases

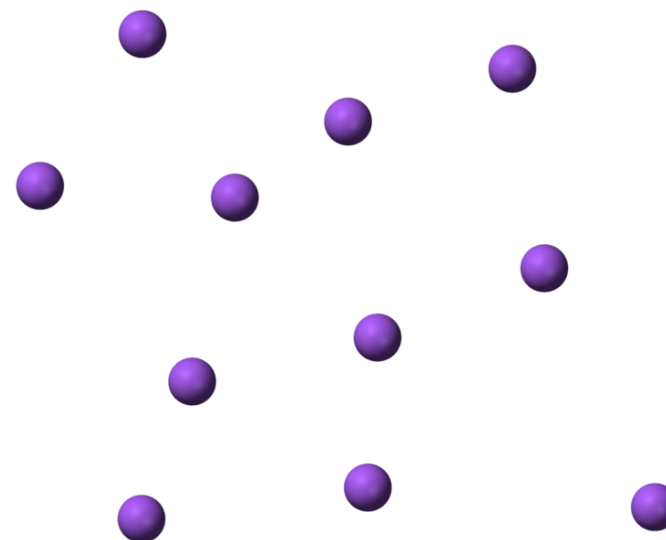


- Given a Maxwell distribution, you should know:
 - Find the approximate V_{rms}
 - Be able to label **which gas is which** (or which temperature is which)
 - Understand the **relationships** between mass, velocity, and temperature
 - Understand how these relationships impact the **shape** of the curve

Extra Slides

Pre-CH301 Knowledge

- You are expected to come into CH301 with a little bit of prior chemistry knowledge:
 - You will not be provided with prefixes (milli, kilo, micro, nano, etc.)
 - You should understand the relationship between molecular weight and weight; moles and atoms
 - You should be able to do stoichiometry as second nature by the first exam (or be okay with getting 8-12 points off to start)
 - Apply dimensional analysis to convert units
 - Understand the relationship between Celsius and Kelvin. **Kelvin = Celsius + 273.15**



Basic Terminology in Reaction Stoichiometry

- Mole: a mole is basically a packet of atoms (6.02×10^{23} atoms to be exact)
 - We use the term “mole” because it is easier to work with in a lab.
 - The mass of each element is presented as its **molar mass** on the periodic table (g/mol)
 - Based on the phase of matter, moles will look different (condensed phases vs. gas phase)
- Limiting reagent: the reactant that **runs out first**, thereby forcing the reaction to stop
- Excess: a reactant that is added in high quantities so that another reactant runs out first
 - You will have a certain amount of this “excess reagent” left over once the limiting reagent runs out
- NOTE: to determine the limiting reagent or the reactant(s) in excess, you must consider the ratio between **the amount present** and the **moles required** to run the reaction.
 - In other words, the limiting reagent is *not always* the reactant with the least number of moles present in the beginning (we will see an example of this later)

Chemistry Fundamentals

There is a simple relationship between mass, moles, and molecular weight.

$$\text{moles} = \text{mass} \div \text{MW}$$

- If you have 50 g N₂, you have 1.8 moles
- $50\text{g N}_2 \times \frac{1\text{ mol}}{28\text{g N}_2} = 1.8\text{ moles}$

$$\text{MW} = \text{mass} \div \text{moles}$$

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

$$\text{mass} = \text{mol} \times \text{MW}$$

- If you have 1.8 moles of N₂, your total mass is 50g
- $1.8\text{ moles N}_2 \times \frac{28\text{ gN}_2}{1\text{ mol}} = 50\text{g}$

Common Stoichiometry Problem

Given the following unbalanced reaction:

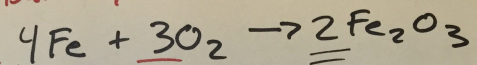


If you have 5 moles of Fe and 10 moles of O₂, identify the limiting reagent and mass of your iron (III) oxide product.

1. Balance the reaction: $4\text{Fe(s)} + 3\text{O}_2(\text{g}) \rightarrow 2\text{Fe}_2\text{O}_3(\text{s})$
2. Solve for the number of reactions you can run with each reactant by comparing the coefficient and the moles you have.
 - You have 5 moles of iron and it takes 4 moles to run one “complete” reaction. Therefore, you can run 1.25 reactions.
 - You have 10 moles of oxygen and it takes 3 moles to run one “complete” reaction. Therefore, you can run 3.33 reactions.
3. Identify the limiting reagent based on which reactant can run the *least number of reactions* based on the amount present.
 - It should be clear that you will run out of iron first. Therefore, iron is the limiting reagent and you have excess oxygen.
4. Using the number of reactions you can run with the limiting reagent, solve for how many moles of iron (III) oxide are present.
 - We decided we can only run 1.25 reactions. Therefore, we have multiply this times the 2 iron (III) oxide per reaction. This means we end up with 2.5 moles iron(III) oxide.
5. Using the number of moles of iron (III) oxide present, multiply by the molar mass to solve for the total mass of product.
 - The molar mass is 160g/mol. You will end up with 400 g iron oxide.

Common Stoichiometry Problem

1) Balance your Rxn



2) How many rxn's can we run?

$$\rightarrow \left[\frac{4 \text{ moles Fe}}{\text{rxn}} \right]$$

$$5 \text{ mol Fe} \times \frac{\text{rxn}}{4 \text{ mol Fe}} = 1.25 \text{ rxn}$$

$$10 \text{ mol O}_2 \times \frac{\text{rxn}}{3 \text{ mol O}_2} = 3.33 \text{ rxn}$$

3) ID Limiting Reagent



4) How much Product?

$$\frac{2 \text{ moles Fe}_2\text{O}_3}{\text{rxn}} \times 1.25 \text{ rxns}$$

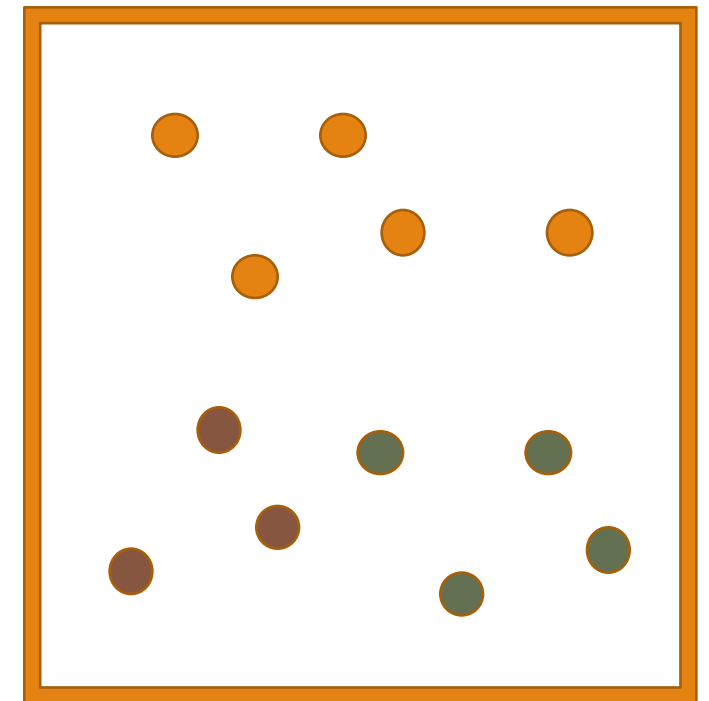
$$= 2.5 \text{ mole Fe}_2\text{O}_3$$

$$\rightarrow 2.5 \text{ moles Fe}_2\text{O}_3 \times \frac{160 \text{ g}}{1 \text{ mol}} = 400 \text{ g}$$

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 or 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 Law of Partial Pressures:
 - The total pressure is equal to the sum of partial pressures
 - $P_{\text{total}} = P_a + P_b \dots + P_i$
- Dalton’s Law “Restated:”
 - The partial pressure of a gas species, i , is equal to the total pressure times the mole fraction of i .
 - $P_i = X_i P_{\text{total}}$

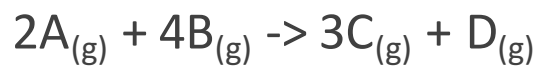
Gas mixture: 1.2atm



● $X = 5/12$.5atm ● $X = 1/4$.3atm ● $X = 1/3$.4atm

Partial Pressure Stoichiometry

Suppose you run the gas phase reaction shown below:



If you start with 8atm of A and 8atm of B, what is the mole fraction of D in the final system (assuming your reaction goes to completion)?

$P_j = X_j P_{total}$

$2A_{(g)} + 4B_{(g)} \rightarrow 3C_{(g)} + D_{(g)}$
 $\xrightarrow{L.R.}$

$8atm \frac{rxn}{4atm B} \Rightarrow 2rxn$

$8atm A \frac{rxn}{2atm A} \Rightarrow 4rxn$

$2rxns \times \frac{3atm C}{rxn} = 6atm C$

$2rxns \times \frac{1atm D}{rxn} = 2atm D$

$8atm A - \left(2rxns \times \frac{2atm A}{rxn} \right)$
 $= 4atm A$ excess

$2atm = X_D | 12atm$

$X_D = \frac{2atm}{12atm} = \frac{1}{6}$

The Ideal Gas Law - R

- The relationship between these state functions is presented in the following equation:

$$PV = nRT$$

- Common error: choosing the correct R value
 - The role of R is to cancel your units. If it does not match the units of pressure, volume, moles, and temperature, you will end up with the wrong value.
 - The units of R stem from the following relationship:

$$\frac{PV}{nT} = R$$