

UWO **CHEM 1302**

Fall 2024, Chapter 2 Notes



Table of Contents

Chapter 2. Gases & The Ideal Gas Law

2.1. Phases of Matter

2.1.1. Phases of Matter

2.2. Gas Laws

2.2.1. Gas Laws

2.2.2. Example

2.2.3. Practice

2.2.4. Practice

2.3. Ideal Gas Law

2.3.1. The Ideal Gas Law

2.3.2. Example

2.3.3. Example

2.3.4. Practice

2.4. Other Applications of the Ideal Gas Law

2.4.1. Applications of the Ideal Gas Law

2.4.2. Standard Molar Volume

2.4.3. Example

2.4.4. Practice

2.4.5. Practice

2.4.6. Practice

2.5. Gas Mixtures & Partial Pressures

2.5.1. Gas Mixtures & Partial Pressures

2.5.2. Example

2.5.3. Practice

2.5.4. Practice

2.6. Kinetic Molecular Theory of Gases

2.6.1. Kinetic Molecular Theory

2.7. Kinetic Energy, Temperature, and Speed of Gas Particles

2.7.1. Relationships Between Kinetic Energy, Temperature, and Speeds!

2.7.2. Root Mean Squared Speed Equation

2.7.3. Example

2.7.4. Example

2.7.5. Practice

2.7.6. Practice

2.8. Graham's Law of Effusion

2.8.1. Diffusion & Effusion

2.8.2. Practice

2.8.3. Practice

2. Gases & The Ideal Gas Law

2.1 Phases of Matter

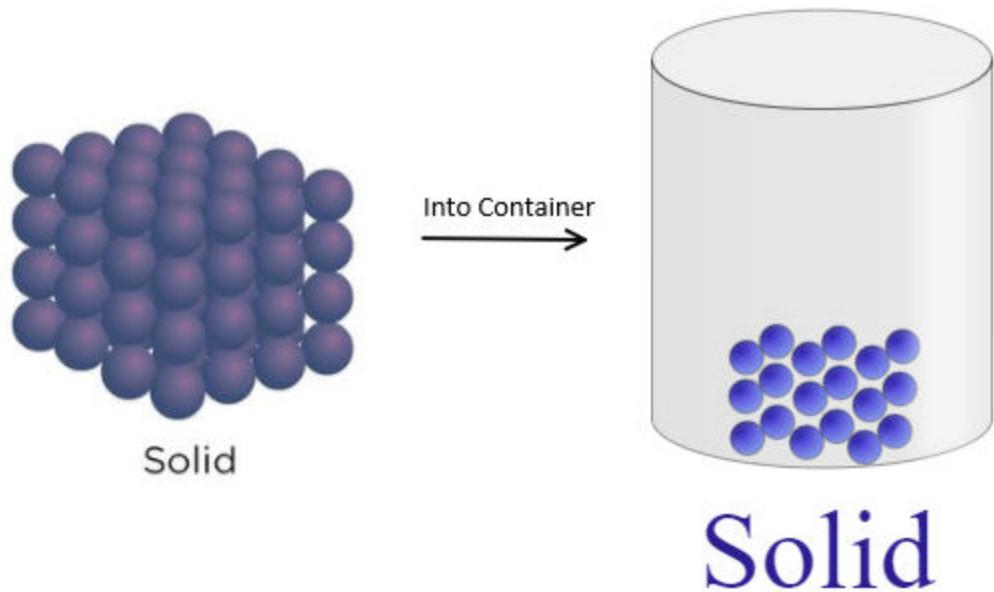
2.1.1

Phases of Matter

There are **3 phases of matter**: solid, liquid, and gas. We need to have a general understanding of each one to understand how a phase change occurs. For example, for the phase change of a solid → liquid, energy is required to break apart some of the strong intermolecular forces that solids have!

Solid

- Don't fill the volume or take the shape of their container (have a fixed volume and shape)
- Hard
- Are not fluid like liquids and gases (don't flow unless extreme shearing/stretching forces)
- Has the **strongest intermolecular forces**

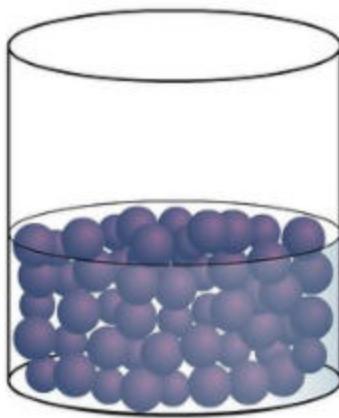


Left photo by Siyavula Education / CC BY

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Liquid

- Don't fill the volume of the container, but take the shape of their container
 - Fixed volume, not shape since if you transfer the liquid to a different container the volume of liquid will remain the same but the shape of the liquid will change depending on the shape of the container
- Fluid
- Has the **weaker intermolecular forces**

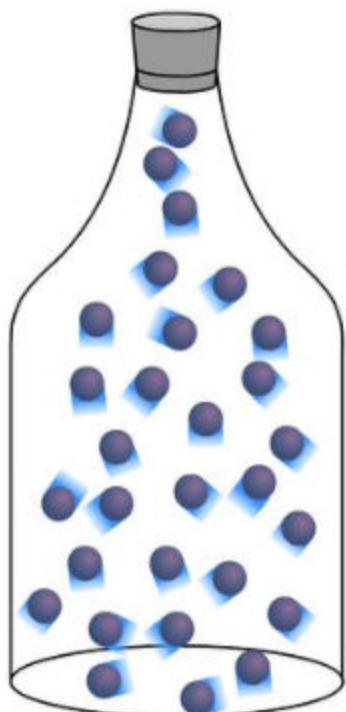


Liquid

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Gas

- Fill the volume and take the shape of their container
 - Volume and shape is not fixed, both can change when you change the container that is holding the gas!
- Fluid
- Has the **weakest intermolecular forces**
 - Note how the molecules are very spread apart from each other. This is because the intermolecular forces between them are very weak!



Gas

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2.2

Gas Laws

2.2.1

Gas Laws

Contained within the ideal gas law are a series of simpler relationships describing how pairs of gas properties change when others are held constant. We will go through each equation and understand how to get each equation without having to memorize them! 🤓

Name	Conditions	Show Your Work to Find Equation	Equation	Are the variables (inversely/directly) proportional
Boyle's Law	Constant T & n			
Charles' Law	Constant P & n			
Gay-Lussac's (aka Amonton's) Law	Constant V & n			
Avogadro's Law	Constant P & T			
Combined Gas Law	Constant n			

 **WATCH OUT!**

Remember that we must plug in a temperature value in K not °C!

$$T(K) = ^\circ C + 273$$

 **WIZE TIP**

You could see two types of gas law questions:

- 1) You're given 4/5 variables for the ideal gas law and asked to solve for the 5th ($PV=nRT$)
- 2) You're told about an **initial scenario and a final scenario** and asked to solve for a missing variable. For these types of problems you'll need to use one of the above **gas laws** depending on what is being kept constant.

Note: You do not have to memorize the names of the equations (ex. Charles' Law) but you should be able to know which equation to use when told certain variables are constant!

Example: Gas Laws

A sample of argon is loaded into a 3 L sealed vessel and at room temperature the initial pressure is 2.6 atm. The vessel is then loaded into an autoclave and heated to 580 °C. What is the final pressure inside the vessel?

2.2.3

Practice: Gas Laws

If the pressure inside a container is kept constant and the container is heated, what would happen to the volume of the gas?

volume of the gas expands

volume of the gas contracts

volume of the gas expands then contracts

more information would be required to answer this question

2.2.4

Practice: Gas Laws

A 15 L sample of N₂ gas held at 328 K is compressed such that the pressure rises from 1.0 atm to 7.8 atm. What is the final volume of the gas?

0.02L



2L



8L



13L



2.3 Ideal Gas Law

2.3.1

The Ideal Gas Law

The equation below **describes ideal gases**. Real gases act differently, but act most similarly to ideal gases at **high temperatures and low pressures!**

$$PV = nRT$$

P is **pressure** measured in kPa (or maybe atm Torr, mmHg...)

V is **volume** measured in **L** or **m^3**

n is **number of moles** of the gas

R is the **ideal gas constant** (will be either **8.314 J/mol K** OR **0.082 L atm/mol K**)

T is **temperature** measured in **K**

 **WIZE TIP**

Pressure=Force/unit area=N/m²=1 Pa

How to convert different units of pressure:

$$101\ 300\ Pa = 101.3\ kPa = 1\ bar = 1\ atm = 760\ mmHg = 760\ torr$$

We mentioned that R (the ideal gas constant) could be either 8.314J/mol K or 0.082 L atm/mol K. **How do we know which value to use?**

! WATCH OUT!

Using the correct value for R is more important than you might think! Students often make the mistake of using the wrong value of R on an exam and they get the question wrong. Let's see how we can easily prevent that :)

The ideal gas constant appears frequently in chemistry, for this reason, it comes in many forms depending on the context of its use and the **units involved**:

Common Values for R
8.3145 kPa · L / mol · K
8.3145 N · m / mol · K
8.3145 J / mol · K
8.3145 kg · m ² /s ² mol · K
0.08206 atm · L / mol · K

Example:

Suppose the variables in the ideal gas law equation had the following units:

$$P(kPa) \times V(L) = n(mol) \times R(?) \times T(K)$$

What should R be in this case? (These terms need to combine correctly to ensure that this equation makes sense!)

2.3.2

Example: Ideal Gas Law

1.1 mols of Argon are stored in a 2.0 L container kept at 10.0 °C. What is the pressure of this gas?
(ans. in kPa)

Example: Ideal Gas Law

- a) A 4 L cylinder containing 3 moles of He(g) arrives from a chemical supply company. Given that the lab you are working in has a room temperature of 298 K, calculate the pressure inside the flask.

- b) At what temperature will the cylinder have an internal pressure of 9 atm?

- c) Your lab mate steals some of the gas from your cylinder for a reaction he is doing. He fills a 1 L bulb with 1 atm of your gas at 298 K. If the cylinder is at room temperature, what is the internal pressure of the cylinder once the gas has been removed?

2.3.4

Practice: Stoichiometry and the Ideal Gas Law

The combustion of methanol (CH_3OH) is shown below. If 16g of methanol is burned in oxygen, what volume of CO_2 is produced if the pressure is 101.3kPa, $T=25^\circ\text{C}$, and $R=8.314 \text{ kPa L/mol K}$? Round to the nearest whole number.



6L



12L



24L



27L



2.4

Other Applications of the Ideal Gas Law

2.4.1

Applications of the Ideal Gas Law

Determining the Molecular Weight (M) of an Unknown Gas

$$PV = nRT \text{ plug in } n = \frac{m}{M}$$

$$PV = \frac{m}{M} RT$$

where, m=mass (g) and M=molar mass (g/mol)

Rearrange the equation to solve for M (molecular weight) of the unknown gas:

$$M = \frac{mRT}{PV}$$

Finding the Density of a Gas

The equation for density (δ) is:

$$\delta = \frac{m}{V}$$

We can rearrange the ideal gas law to find density:

$$PV = nRT \text{ plug in } n = \frac{m}{M}$$

$$PV = \frac{mRT}{M}$$

Solve for density:

$$\frac{m}{V} = \delta = \frac{MP}{RT}$$

Standard Molar Volume

Recall: we talked about Avogadro's law where P and T are kept constant.

$$PV=nRT$$

$$V/n=RT/P \text{ where } RT/P=\text{constant}$$

$$V/n=\text{constant}$$

Based on this, imagine if we were looking at two different gases in two different containers, each with the same P, T, and number of moles of gas particles.

We have made n (moles) a constant too. If that is the case, then the volume for all the gases should be the same, and they are!



WIZE CONCEPT

Specifically at STP ($P=1\text{ atm}$ and $T=0^\circ\text{C}$ or 273 K), 1 mol of any gas occupies 22.4L! **Memorize this!**

STP=Standard Temperature and Pressure

@ STP:

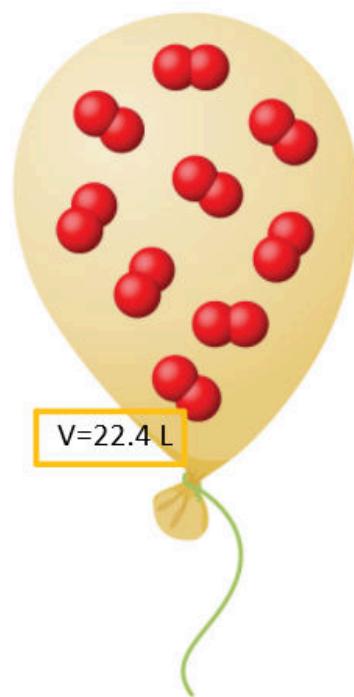
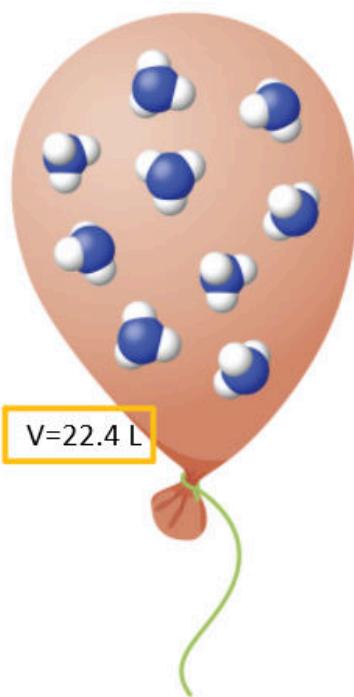
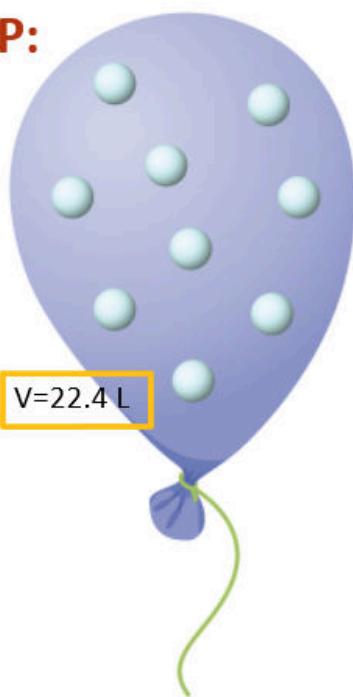
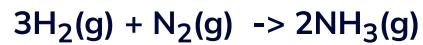


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2.4.3

Example: Standard Molar Volume

In the following reaction, if you start with 4L H₂(g) and 1.5L N₂(g) at STP, what would the volume of each of the 3 gases be when the reaction is complete?



2.4.4

Practice: Solving for Molar Mass of a Gas

The density of an unknown gas at 100°C and 746 torr is 1.994 g/L. What is the molecular mass of this unknown gas?

8.3 g/mol



31.0 g/mol



62.1 g/mol



124.2 g/mol



1689 g/mol



2.4.5

Practice: Determine the Gas

0.238 g of a gas were placed in a 250 mL vessel. If the gas is heated to 85 °C, the gas has a pressure of 7 atm. Which gas is in the vessel?

H₂ (g)

He (g)

N₂ (g)

F₂ (g)

2.4.6

Practice: Solve for Molecular Weight

A 6.00 L flask is evacuated and weighed, 21.64g. The vessel is then filled with a gas at STP, after filling the vessel weighs 27.31g. What is the molecular weight of the gas?

0.2446 g/mol

1.0 g/mol

2.446 g/mol

21.46 g/mol

227.2 g/mol

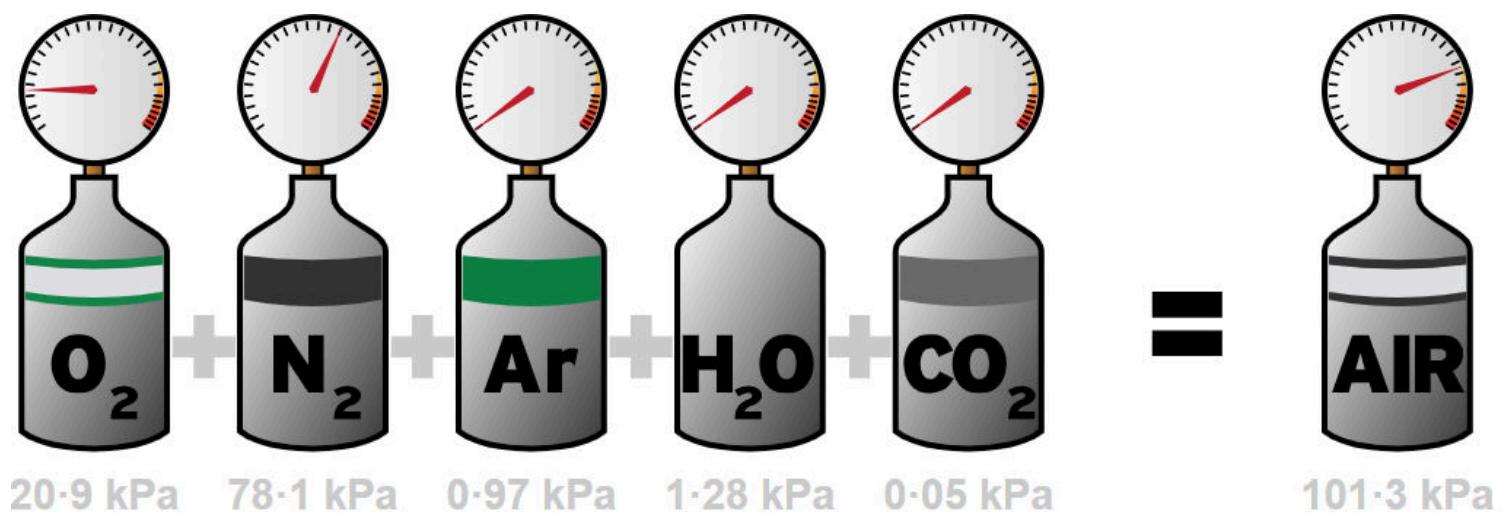
2.5

Gas Mixtures & Partial Pressures

2.5.1

Gas Mixtures & Partial Pressures

When dealing with mixtures of gases, the total pressure of the gases is equal to the sum of the **partial pressure** of each component.



Dalton's Law of Partial Pressures:

$$P_{Total} = P_1 + P_2 + P_3 \dots = \sum P_i$$

P_{Total} is the single **measured pressure of the mixture**

P_1 is the **partial pressure of gas 1**, and so on

Partial pressure: is the pressure that would be exerted by one of the gases in the mixture if it occupied the same volume on its own.

Note: The total pressure depends on the total number of moles (not on the identity of the gases)

$$PV = nRT$$

$$P_{total} = \frac{n_{total}RT}{V}$$

We can also relate the partial pressure of an individual gas (ex. Gas "A") to the total pressure by the **mole fraction (X)** of that gas.

The **mole fraction (X)** is a **proportion** of that gas.

$$P_A = \frac{n_A}{n_{Total}} P_{Total} = X_A P_{Total}$$

2.5.2

Example: Partial Pressures

If we have a container with 75% nitrogen, 20% oxygen, and 0.04% carbon dioxide, the total pressure in the container is measured to be 760 Torr (recall this is equal to 1 atm!). What is the partial pressure of oxygen gas? (Ans in torr)

2.5.3

Practice: Partial Pressures

A vessel contains He, Ne, and Ar gas. There's twice as much He as Ne and half as much Ar as He. The total pressure in the vessel is 400torr. What is the partial pressure of Ar(g)?

75 torr

100 torr

175 torr

225 torr

2.5.4

Practice: Partial Pressures

Two gases were mixed together in a 5 L vessel at 243 K to create a total pressure of 9.3 atm. If gas 1 has 2 moles, what is the partial pressure of gas 2? ($R = 0.08206 \text{ L}^*\text{atm}/\text{mole}^*\text{K}$)

0.7 atm



1.3 atm



3.1 atm



5.4 atm



2.6

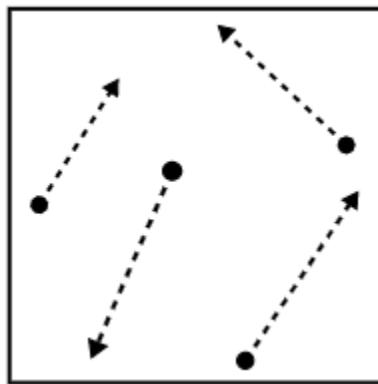
Kinetic Molecular Theory of Gases

2.6.1

Kinetic Molecular Theory (KMT)

This theory describes the behaviour of **ideal gases**.

In normal conditions we can apply this to real gases as well...(more on the conditions required for this later)



The KMT operates under the following assumptions/simplifications:

 WIZE CONCEPT

1. Ideal gases are in **constant, random motion** and move in straight lines
2. The size of gas particles is negligible compared to the distance between particles. As a result of this, the **volume of gas particles is considered to be 0**.
3. There is constant motion of particles and **collisions** of gas particles (with other gas particles and with the walls of the container) **are completely elastic** so that the total kinetic energy is conserved before and after a collision.
4. There are **no intermolecular forces** between gas particles (they don't attract or repel each other).
5. **Temperature is directly proportional to the average kinetic energy** of the gas particles.

- If we increase temperature, do you think the gas molecules would collide with container walls more/less often? _____ often.
- And with an increase in temperature would the gas particles collide at slower or greater speeds?
_____ speeds

! WATCH OUT!

T should always be in Kelvins (K) for chemistry calculations :)

$$T(\text{in K}) = T(\text{in } ^\circ\text{C}) + 273$$

i WIZE TIP

You can just assume we are dealing with an ideal gas unless stated otherwise in a particular exam question!

Note: to measure pressure($P=F/A$), we need to take into consideration the **number of collisions and force of those collisions between gas particles and the walls of the container**

This can help us understand WHY the KMT says what it does for IDEAL gases:

- If gas particles had volume they would be quite large and would have more/less: _____ intermolecular forces/attractions to one another. This would mean that they would interact with the walls of the container more/less: _____ and result in a decrease/increase: _____ in pressure.
- If collisions weren't elastic, then over time gas particles would hit the walls of the container with more/less: _____ force and increase/decrease: _____ the pressure.
- We will explain this concept more when we look at real vs ideal gases!

The KMT for ideal gases helps to keep conditions in the container stable!

2.7

Kinetic Energy, Temperature, and Speed of Gas Particles

2.7.1

Relationships Between Kinetic Energy, Temperature, and Speeds!

Just focus on **understanding**  the following lesson.

Scenario #1:

We have two different gases at the **same temperature** in two different containers. What can we say about the **average kinetic energy** of the two gases? (are they the same, different, if different how?)

Hint: This was the part of the kinetic molecular theory!

We can make sense of this using the following equation:

$$KE_{avg} = \frac{3}{2}RT$$

KE_{avg} is the average kinetic energy of a mole of gas particles

R here is 8.314 J/mol K

How are KE_{avg} and T related?

! WATCH OUT!

The molar masses of the gases are not important when it comes to the average kinetic energy so if you're ever asked about the average kinetic energy, you just need to think about how it relates to temperature!

Scenario #2:

We have two different gases at the same temperature in two different containers. What can we say about the speeds of the two gases? (are they the same, different, if different how?)

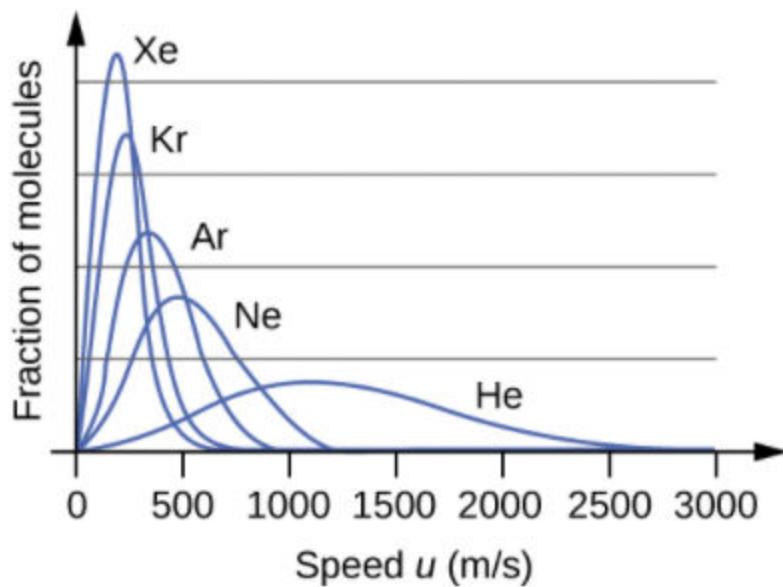


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Therefore, a heavier gas will move at a (faster/slower) _____ average speeds than a lighter gas at the same temperature.

Note: The lighter the gas molecule, the more (broad/narrow) _____ the speed distribution

Scenario #3:

If we just consider one gas in a container, **all of the gas particles in the container don't move at the exact same speed!**

There is actually a distribution of speeds.

If temperature gets increased, then the *average* speed (increases/decreases) _____ and the average kinetic energy (increases/decreases) _____ and the spread in the graph below (increases/decreases) _____.

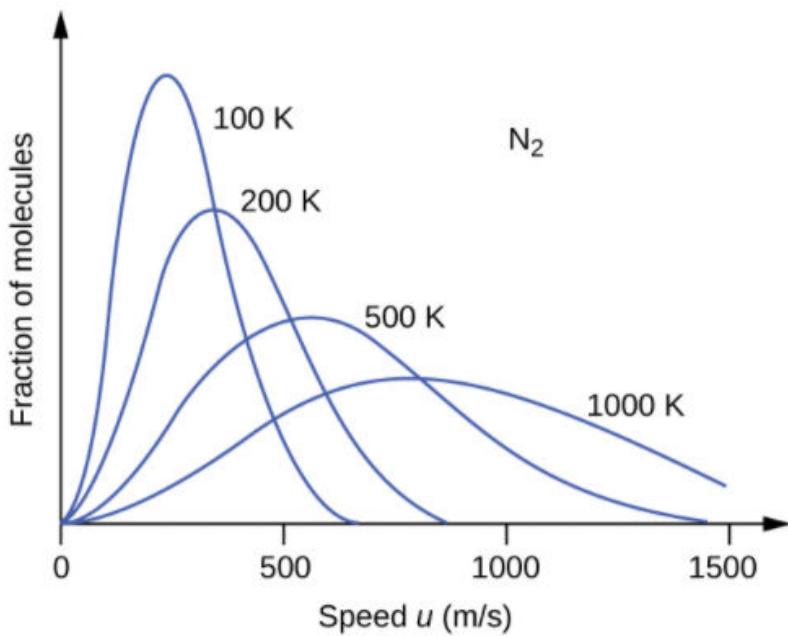


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Root Mean Square Speed (u_{rms})

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

R is the **ideal gas constant (8.314 J/mol K)**

T is **temperature in K**

M is **molar mass in kg/mol**

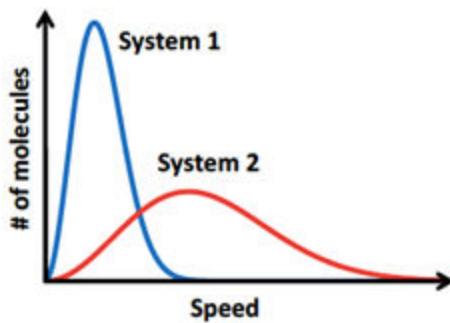
! WATCH OUT!

To get the correct units for u_{rms} (m/s) we need to **use units kg/mol for molar mass in this equation.**

The good thing is that for all other equations, we use the units g/mol for molar mass.

Example: Applications of Kinetic Molecular Theory

Consider the following two curves showing distributions of the number of molecules vs. their molecular speeds for two different gas samples, System 1 and System 2. If the area under both curves is equal, provide two possible scenarios which would result in these observed differences. Complete the table below to describe each of your two scenarios.



	Scenario 1	Scenario 2
What single property could be different between System 1 and System 2?		
For the property you chose, which system has a higher value?		
Sketch what the kinetic energy distribution would look like for System 2.		

Example: Root Mean Squared Speed

Gas particles in an unknown sample are measured to have a root mean square speed of 324.2 m/s at 80 °C. If this sample is known to be a pure element, what is the most likely identity of this gas?

2.7.5

Practice: Root Mean Squared Speed

Which of the following would have the lowest u_{rms} ?

A) NO₂ at 273 K

B) NO₂ at 453 K

C) O₂ at 273 K

Practice: Kinetic Energy of Gases

Which of the following have the same average kinetic energy?

- I He at 297 K
- II He at 351 K
- III H₂ at 297 K
- IV O₂ at 212 K
- V SO₂ at 297 K

I and II

II and IV

I, III, and V

all of the above

2.8

Graham's Law of Effusion

2.8.1

Diffusion and Effusion

Diffusion: mixing molecules of one gas with another as a result of random molecular movement. (This term is seen more in biology!)

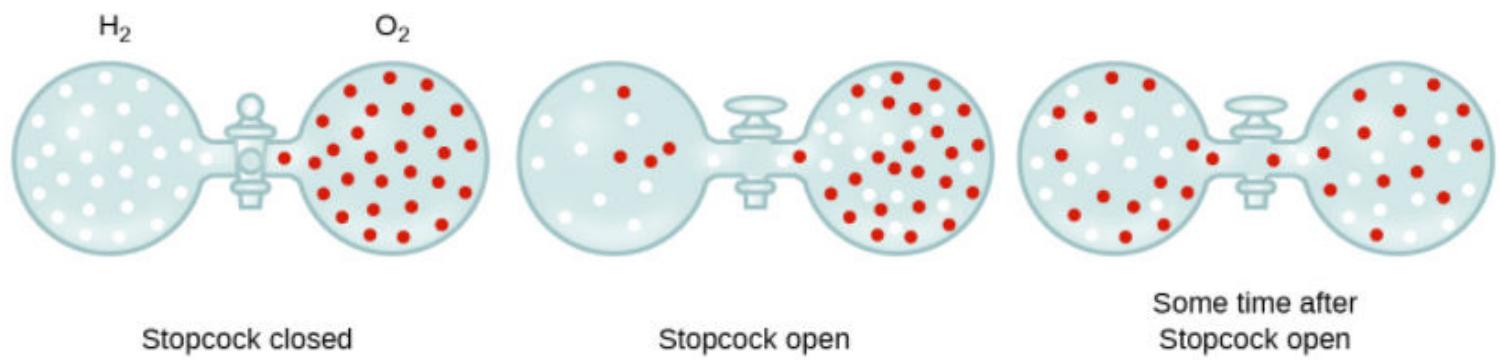
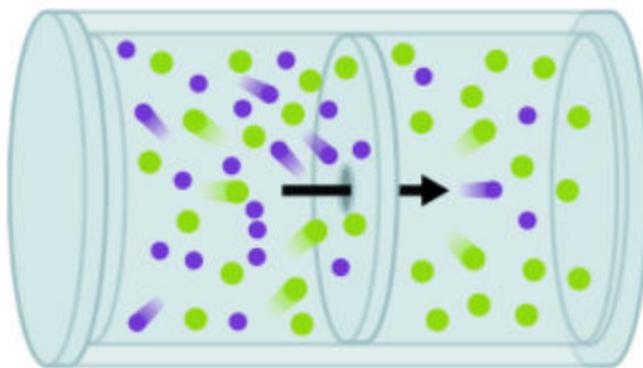


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Effusion: Escape of gas molecules from their container through a tiny pinhole. This is what is important to us for gases!



Effusion

Photo by CNS OpenStax / CC BY

According to the image above, do the green or purple particles have a faster rate of effusion?

 **WIZE CONCEPT**

A higher or faster rate of effusion means that those particles escape through the hole at a faster rate!

Do you think lighter or heavier molecules would be able to effuse faster? _____

Graham's Law of Effusion:

$$\frac{\text{rate of effusion A}}{\text{rate of effusion B}} = \frac{(u_{\text{rms}})_A}{(u_{\text{rms}})_B} = \sqrt{\frac{M_B}{M_A}}$$

where M=Molecular weight

A is for gas "A"

B is for gas "B"

Based on the equation, how is the rate of effusion for a particle related to its molecular weight?
(directly proportional, inversely proportional, other, not related??)

Practice: Graham's Law of Effusion

If we started with the same number of dark and light particles in one container and after 2 minutes we looked and saw that 4 dark particles and 2 light particles had effused to the other container, how much heavier/lighter are the dark particles than the light particles?

- A) Dark particles are two times heavier than the light particles
- B) Dark particles are two times lighter than the light particles
- C) Dark particles are four times lighter than the light particles
- D) Dark particles are four times heavier than the light particles

2.8.3

Practice: Speeds of Gases

An ambitious group of scientists conducted two separate experiments where they isolated 1000 atoms each of Li and Rb in a tiny volume of $5 \times 10^{-15} \text{ m}^3$ at a temperature of 1 Kelvin. How many times faster or slower would Li atoms be effusing than Rb atoms?

Li effuses 0.28x faster than Rb

Li effuses 3.5x faster than Rb

Li effuses 3.5x slower than Rb

Li effuses 0.28x slower than Rb