

# Queen's **CHEM 112A**

**Fall 2024, Chapter 2 Notes**



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# 2. Gases

## 2.1 Pressure

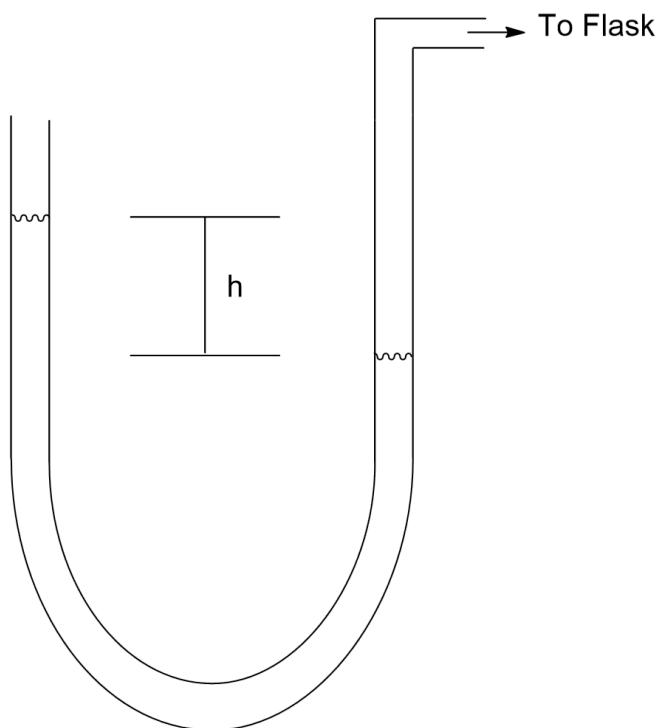
2.1.1

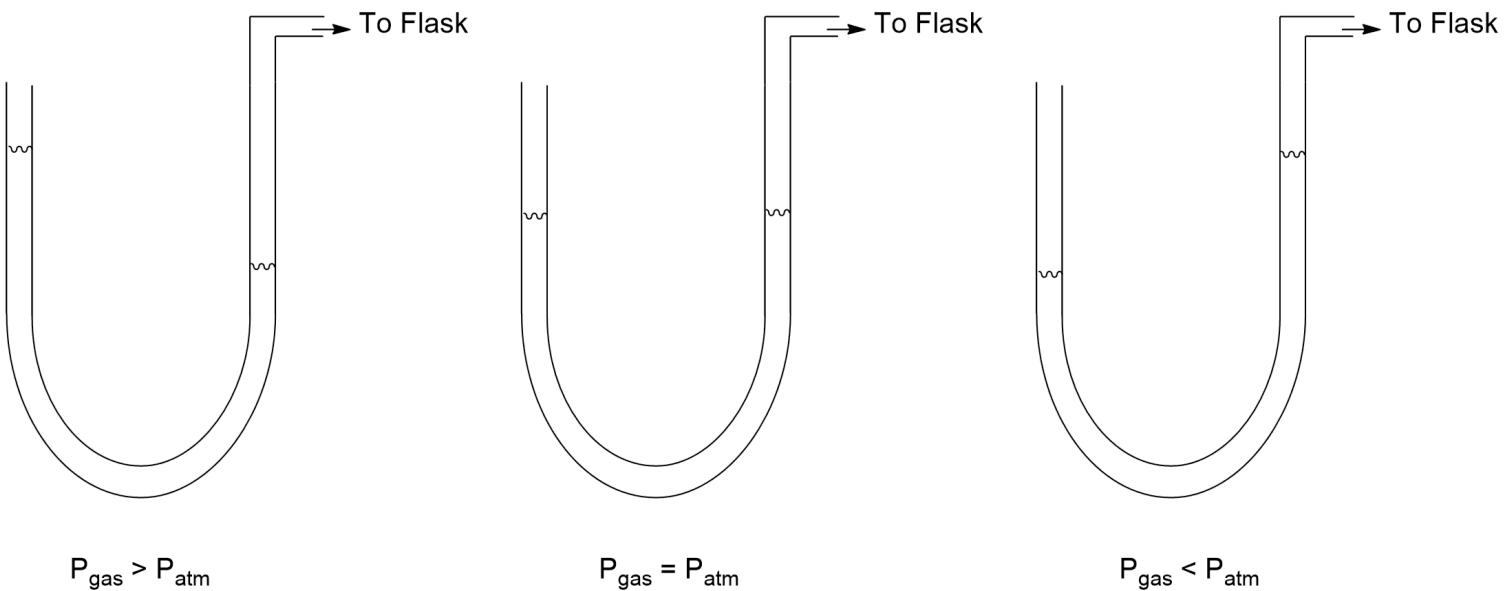
### Measuring Pressure in the Lab

There are numerous ways to measure atmospheric pressure but the original and simplest method involves the use of a manometer.

**Manometer:**

- A manometer is a u shaped tube partially filled with mercury and open at both ends. To one of those ends is attached a hose to the flask or system that you want to measure. the other is left open to the atmospheric pressure.
- If the flask pressure is higher than the atmospheric pressure it will push the mercury down one arm and up the other towards the atmospheric pressure
- If the flask pressure is lower than atmospheric pressure, the atmospheric pressure will push the mercury down one arm and up the other towards the flask.
- If the levels of mercury are even, then the pressure of the flask is equal to atmospheric pressure.





$$\Delta P = \rho g h$$

If the pressure of the flask is higher:

$$P_{\text{gas}} = P_{\text{atm}} + |\Delta P|$$

If the pressure of the flask is lower:

$$P_{\text{gas}} = P_{\text{atm}} - |\Delta P|$$

---

### 2.1.2

You are visiting Hawaii and decide to take your barometer with you. When sitting on the beach you notice that the barometer is reading roughly 1 atm. The next day you take your barometer up to the top of Mauna Kea (roughly 4200m above sea level). You notice that the barometer has dropped and the column of mercury is only 45.6 cm high. The density of mercury is  $13.595 \text{ g cm}^{-3}$  and the force of gravity is  $9.8 \text{ m s}^{-2}$ , what is the atmospheric pressure on top of Mauna Kea? Please report your answer in atm.

0.1 atm

0.6 atm

1.2 atm

2.4 atm

## 2.2 Ideal Gas Law

2.2.1

### 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/m<sup>2</sup>=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 <sup>2</sup> /s <sup>2</sup> 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!)

---

## 2.2.2

### 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?

2.2.4

## Practice: Stoichiometry and the Ideal Gas Law

The combustion of methanol ( $\text{CH}_3\text{OH}$ ) is shown below. If 16g of methanol is burned in oxygen, what volume of  $\text{CO}_2$  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



---

## 2.3 Gas Laws

2.3.1

### 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?

## 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

2.3.4

## Practice: Gas Laws

A 15 L sample of N<sub>2</sub> 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



## 2.4

# 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 ( $\delta$ ) 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  $273\text{ K}$ ), 1 mol of any gas occupies 22.4L! \*\*Memorize this!\*\*

**STP=Standard Temperature and Pressure**

@ STP:

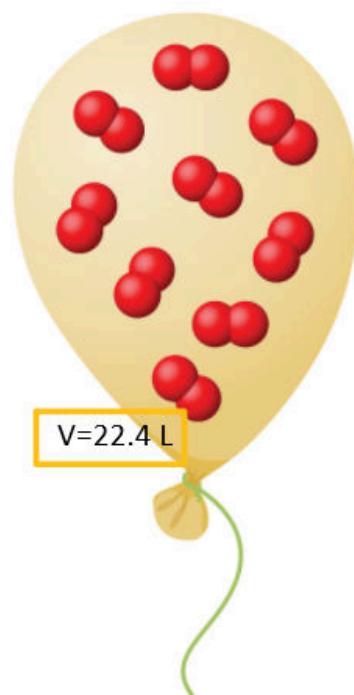
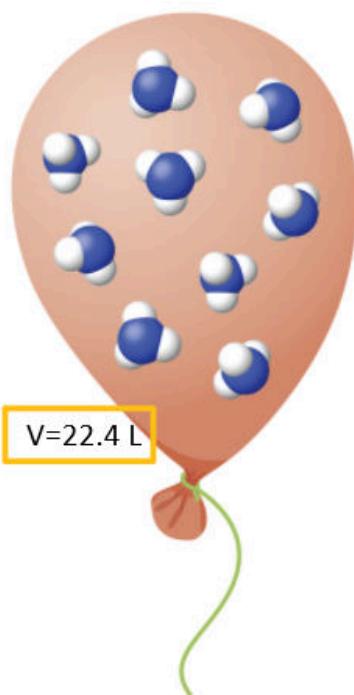
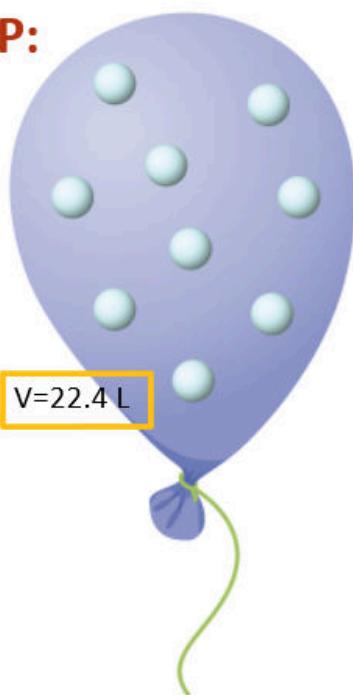


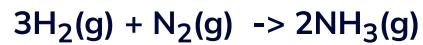
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#### 2.4.3

### Example: Standard Molar Volume

In the following reaction, if you start with 4L H<sub>2</sub>(g) and 1.5L N<sub>2</sub>(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?

H<sub>2</sub> (g)

He (g)

N<sub>2</sub> (g)

F<sub>2</sub> (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

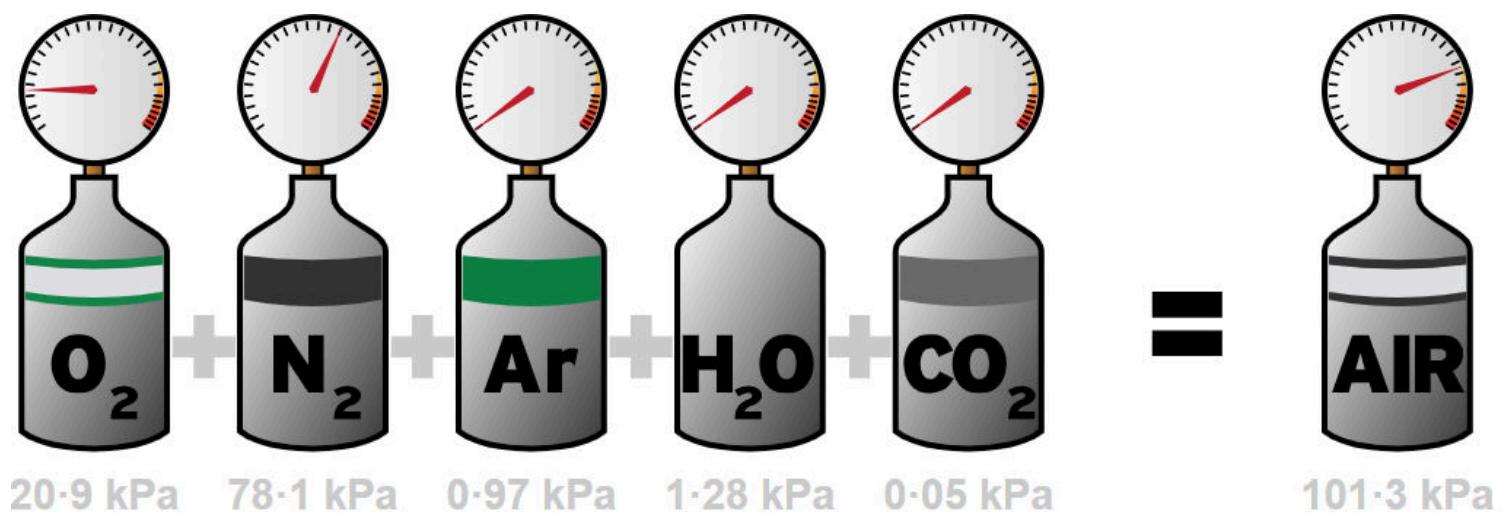
## 2.5

# 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

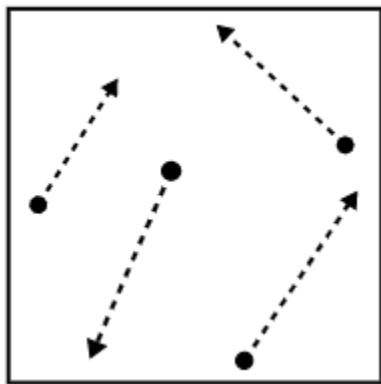
# Kinetic Molecular Theory of Gases

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?)

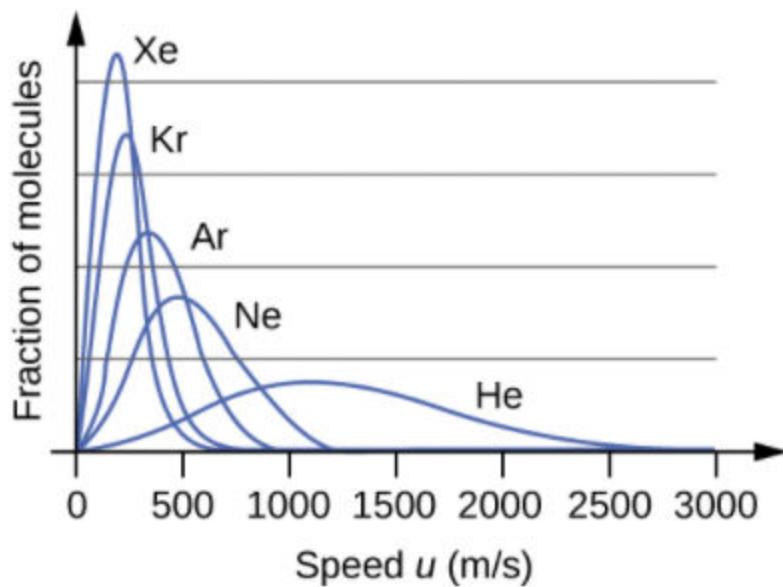


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Therefore, a heavier gas will move at a (faster/slower) \_\_\_\_\_ average speeds than a lighter gas at the same temperature.

**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) \_\_\_\_\_.

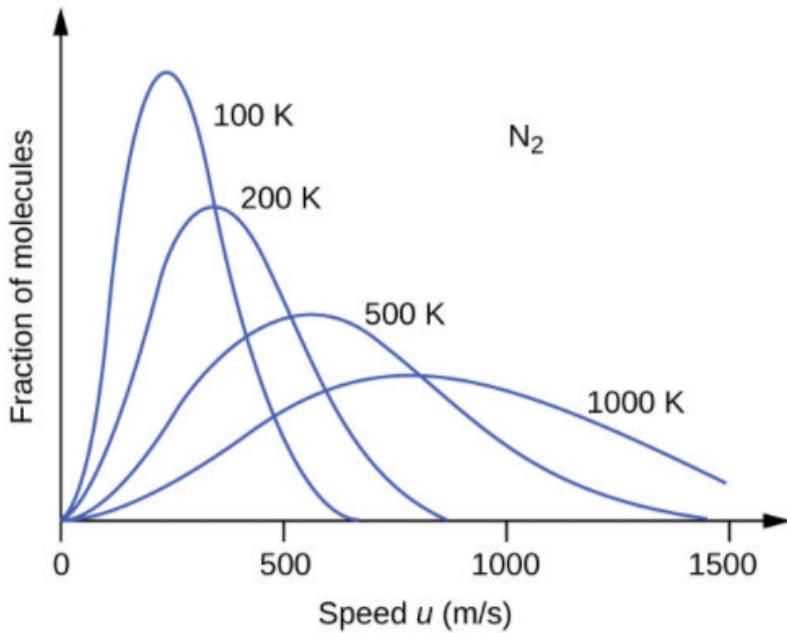


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## Root Mean Square Speed ( $u_{rms}$ )

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

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

T is **temperature in K**

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

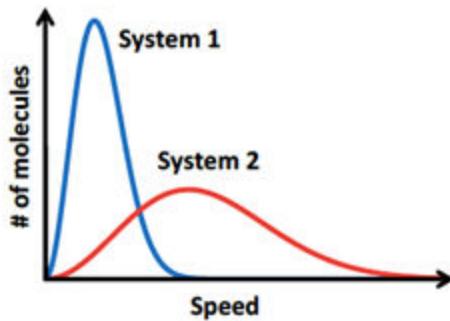
### ! WATCH OUT!

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  $u_{\text{rms}}$ ?

A) NO<sub>2</sub> at 273 K

B) NO<sub>2</sub> at 453 K

C) O<sub>2</sub> 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 H<sub>2</sub> at 297 K
- IV O<sub>2</sub> at 212 K
- V SO<sub>2</sub> 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!)

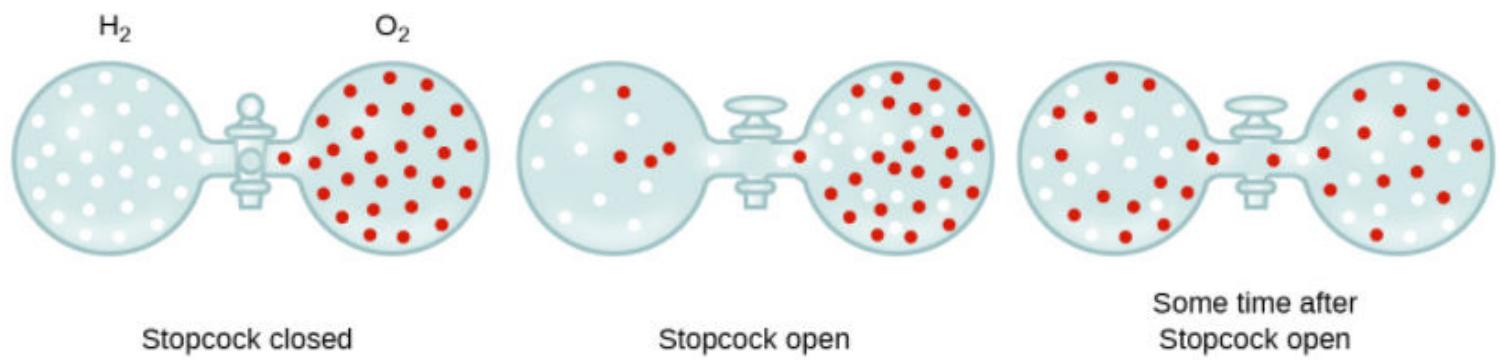
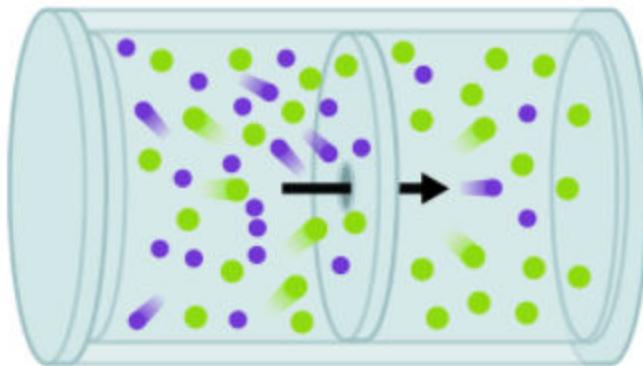


Photo by CNS OpenStax / CC BY

**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

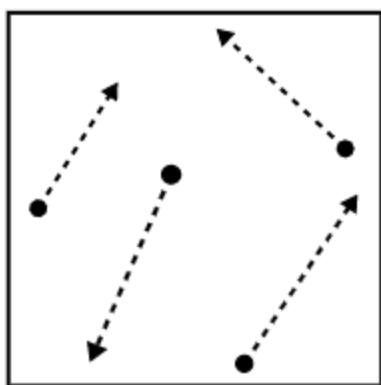
Li effuses 0.28x slower than Rb

## 2.9 Real vs Ideal Gases

2.9.1

### Real Gases and the Van der Waals Equation

#### Ideal Gases



##### i WIZE TIP

Ideal gases have **negligible volume**.

And there are **no intermolecular forces** between gas particles.

Note:  $PV=nRT$  is the **IDEAL gas law** and describes **IDEAL gases**

## Real Gases

### WIZE TIP

Real gases **have volume**.

And they **have intermolecular forces** (attraction between gas molecules).

Real gases are most like ideal gases when temperature is high (well above the condensation temp for that gas), pressure is low (usually less than 1 atm)

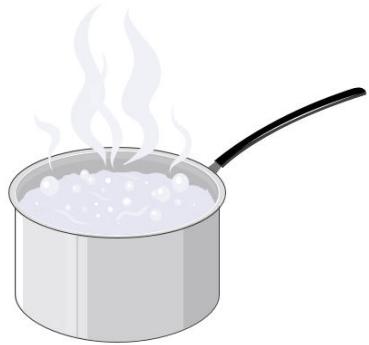
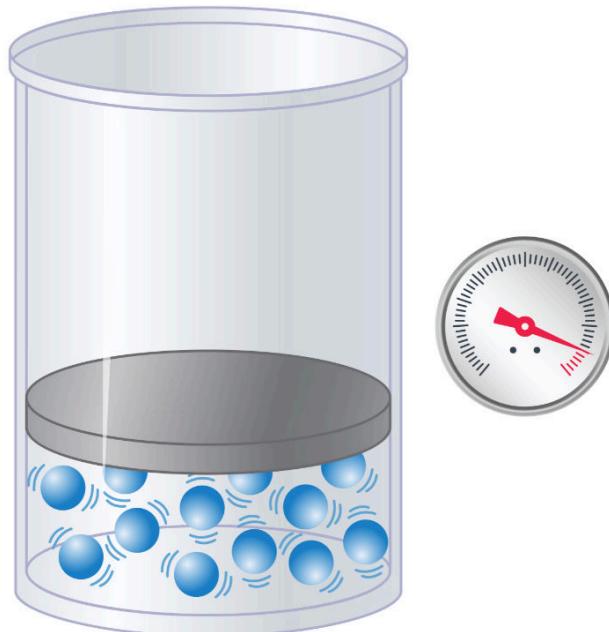


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### WIZE CONCEPT

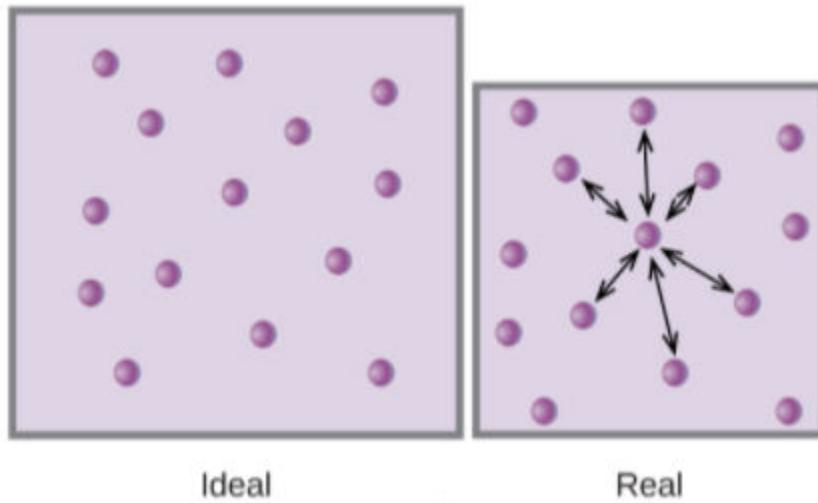
Real gases deviate more from ideal gases when **temperature is low** (near bp) and **pressure is high** ( $\sim > 10\text{ atm}$ )

- With a lower temperature and higher pressure just think that we are going more towards liquids than gases
- High pressures help to compress and push molecules closer together, resulting in less distance between molecules. As a result, the volume of the gas particles becomes more significant.
  - Both of these factors (low temp and high pressure) cause gas molecules to interact more with each other (have intermolecular forces) and interact less with the walls of the container (decreasing pressure).



**High pressure**

## Ideal vs Real Gas in a Constant Pressure Container (Volume of the Container Can Change)



### Ideal Gas

- Has no intermolecular forces between gas particles
- Has no volume
- Container remains unchanged

### Real Gas

- Real gases have a significant volume!
- Since they **have volume**, real gas particles want to interact with other gas particles (**intermolecular attractions**)
- This interaction causes the particles to come closer together and the **volume of the container decreases**

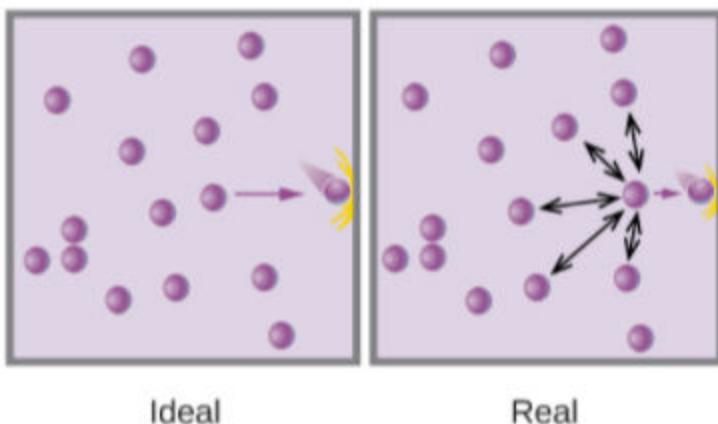
#### WIZE CONCEPT

For real gases:

$$V_{\text{(gas)}} > 0$$

$V_{\text{(container)}}$  less than the volume of a container for an ideal gas.

## Ideal vs Real Gas in a Constant Volume Container (Pressure in the Container Can Change)



### Ideal Gas

- Ideal gases have no intermolecular forces

### Real Gas

- Real gases **have intermolecular forces**
- The intermolecular forces between real gas particles causes real gas particles to spend (more/less) \_\_\_\_\_ time close to other gas particles compared to ideal gases
- As a result, real gas particles hit the walls of the container (more/less) \_\_\_\_\_ and with (more/less) \_\_\_\_\_ force
- The pressure inside the container for a real gas is therefore (higher/lower) \_\_\_\_\_ than the pressure inside the container for an ideal gas!

#### WIZE CONCEPT

Real gases have intermolecular attractions and this results in the pressure of the container decreasing.

# Van der Waals Equation

*ideal gases*  $\rightarrow PV = nRT$

$$\boxed{\text{real gases} \rightarrow \left(P + \textcolor{blue}{a} \frac{n^2}{V^2}\right)(V - nb) = nRT}$$

For real gases, we need to adjust the ideal gas law. The adjusted version is called the **Van der Waals equation**.

**Real gas particles are attracted to each other (they have intermolecular attractions).**

- In the Van der Waals equation, to "adjust" the equation to be more like ideal gases there is a constant, "a"
  - **a=attractive force correction**
  - **a increases with increasing intermolecular attraction between molecules**

**Real gas particles also have volume.**

- In the Van der Waals equation, to "adjust" the equation to be more like ideal gases there is a constant, "b"
  - **b=bulkiness correction**
  - **b increases with molecular weight** of gas molecule (more volume)

**Note:**  $a$  and  $b$  are Van der Waals constants and experimentally determined parameters that are different for every gas.

**i WIZE TIP**

**Know the differences between ideal and real gases** (you are most likely to be tested on this)

Know what the constants **a** and **b** mean in the Van der Waals equation.

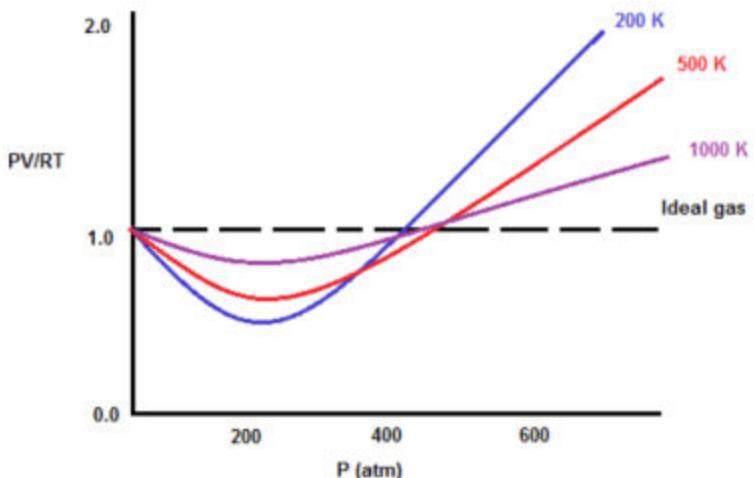
You don't need to memorize the Van der Waals equation :)

## Real Gases vs Ideal Gases & The Compressibility Factor

Review: There are two main differences between real gases and ideal gases: (\*\*you should know this!!)

Review: Real gases are more likely to deviate from ideal behaviour under what conditions? (\*\*you should know this!!)

At \_\_\_\_\_ pressure and \_\_\_\_\_ temperature real gases deviate most from ideal gas behaviour.



Let's interpret the graph based on our knowledge!

- $PV=nRT$
- $PV/RT=n$  so the y axis is  $n$  and on the graph we have ideal gas labelled as 1 mol
- In lower temperatures did we say we would see more/less of a deviation from an ideal gas?

- 
- This is why we see the 200K line curve farthest away from the horizontal line representing an ideal gas!

## Compressibility Factor (Z)

- We can calculate how close a real gas behaves like an ideal gas by measuring the **compressibility factor, Z**.
- For an **ideal gas,  $Z=1$**
- In this equation  $V$  is the molar volume ( $V/n$ )

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

## Example: Van der Waals Equation

3 moles of He gas are released into a 4 L container at 275 K. If He is a real gas, what is the pressure of the gas? (For He,  $a = 0.035 \frac{L^2 atm}{mol^2}$ , and  $b = 0.2371 \text{ L/mol}$ ). \*Please note the mistake in the video for the units of  $a$

As this is a real gas, we need to setup the Van der Waal's equation:

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

## Example: Real vs Ideal Gases

Find the difference between the ideal pressure and the Van der Waals pressure for 2 mol of CO<sub>2</sub> in a 1L vessel at room temperature. CO<sub>2</sub> ( $a = 3.640 \text{ L}^2\text{atm/mol}^2$ ;  $b = 0.04267 \text{ L/mol}$ ).

**2.9.5**

Out of the following, which gas would deviate the most from ideal behavior at a low temperature and high pressure?

$\text{Cl}_2$

$\text{CO}_2$

$\text{H}_2$

$\text{N}_2$