Gases

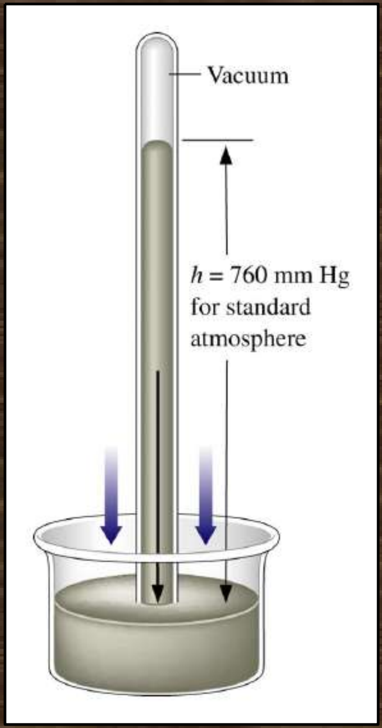
Chapter 5

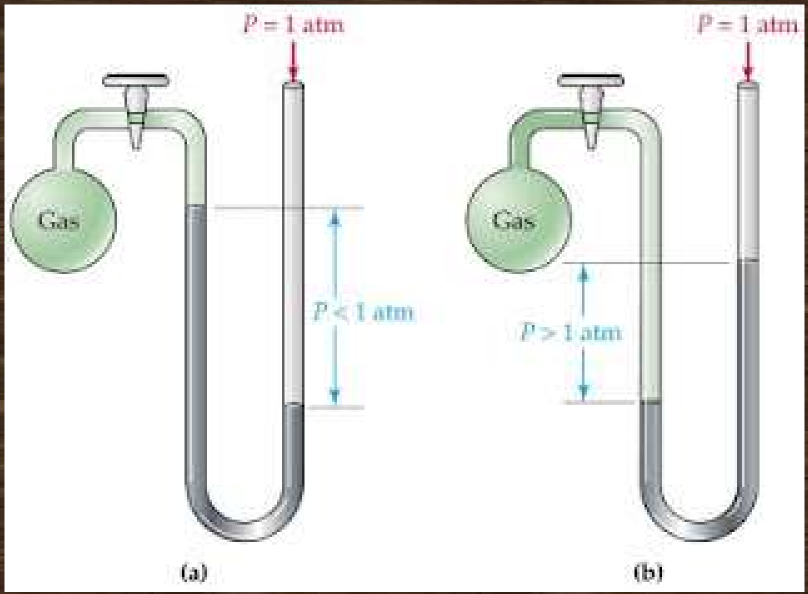
General Properties

- Fills the container
- Easily compressed
- Mixes completely with all other gases
- Exerts pressure on its surroundings
- Has no attraction for other gas particles (Ideally)

Measuring Pressure

- Barometer:
  - Measures atmospheric pressure
- Atmospheric Pressure:
  - Pressure exerted by atmosphere (varies)
- Manometer:
  - Measures the pressure of a gas in a container





Pressure

- Pressure = force/area

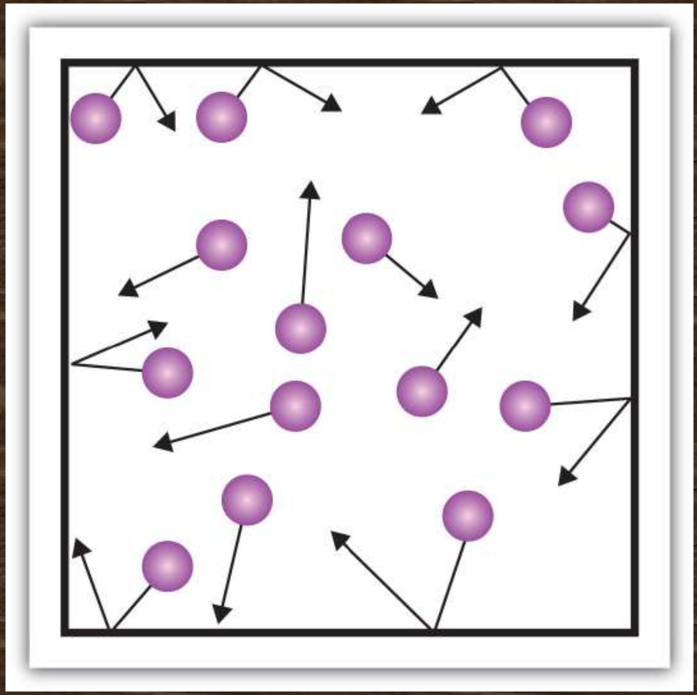
- Common Units

  mmHg

  torr

  atm (standard atmosphere)

  Pa = pascal = N/m2

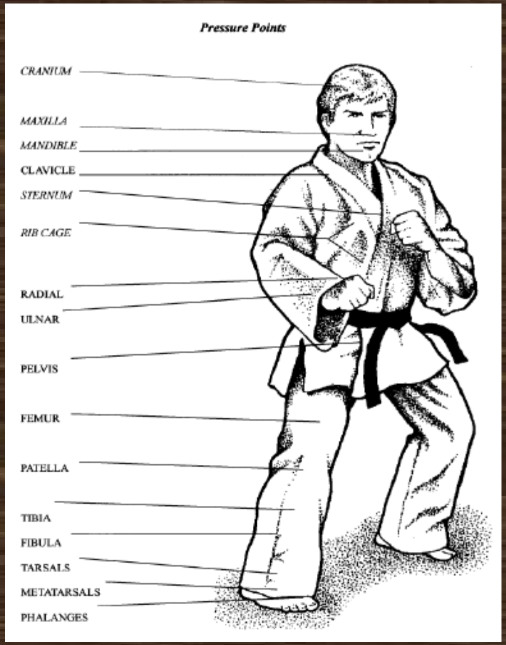


Pressure Conversion

- Standard pressure
- 1 atm=760 mmHg=760 torr= 101.3 kPa

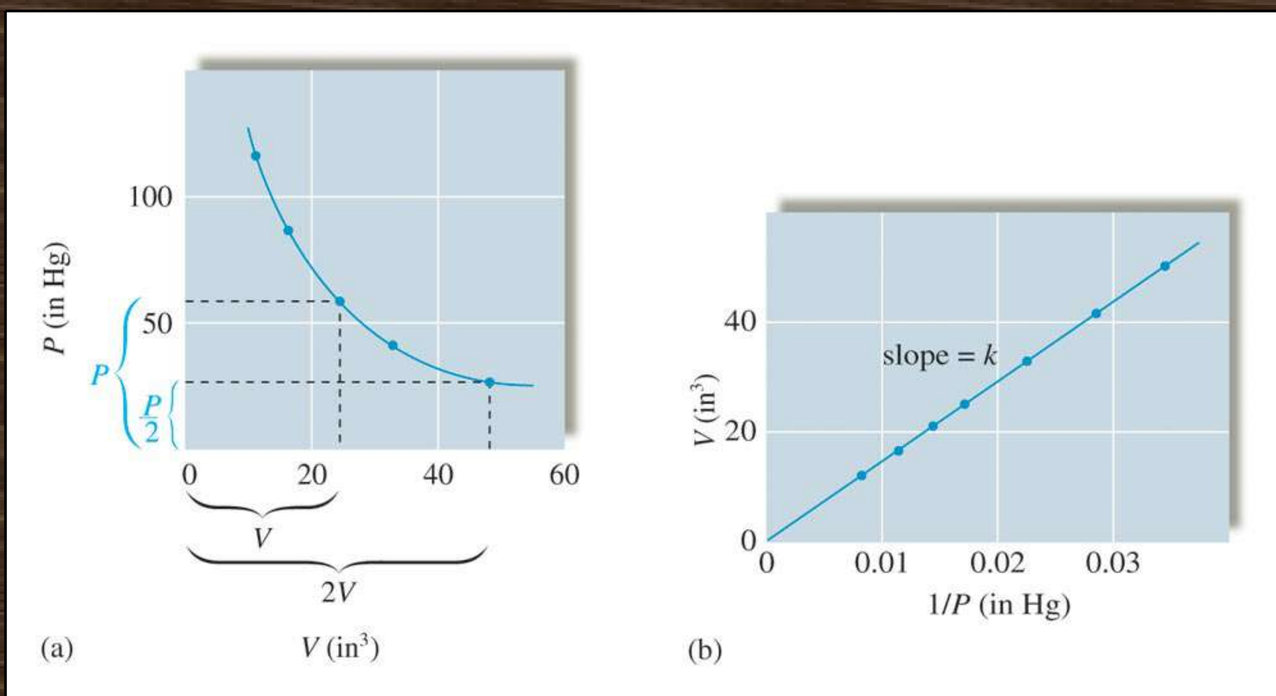
- Convert:
- 520mmHg = ___atm = ___kPa
- 235kPa = ___torr = ___atm

HW: #27, 28, 31



Boyle's Law

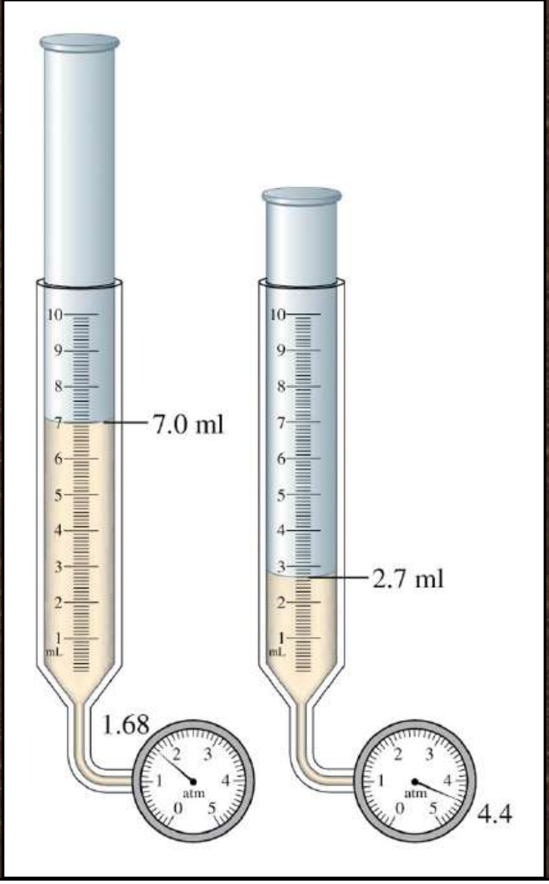
- **Pressure increases as volume decreases.**
- **Inverse relationship: ½ V = 2P ¼ V = 4P**
- **PV = k (where k is a constant)**
- **$P_1V_1 = k = P_2V_2$ OR $P_1V_1 = P_2V_2$**

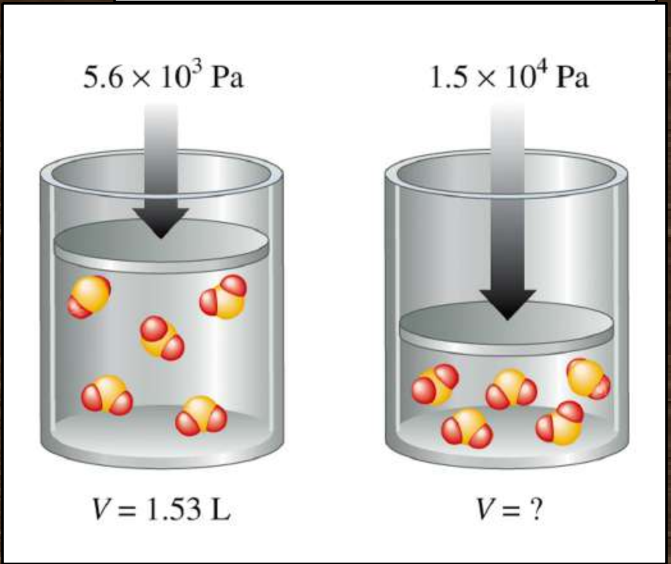


Boyle's Law Example

1. A sample of gas occupies 250. mL at 740. torr pressure. What volume will it occupy at 370. torr pressure?
2. A sample of fluorine gas exerts a pressure of 900. torr. When the pressure is changed to 1.50 atm, its volume is 250. mL. What was the original volume?

HW: #33, 101

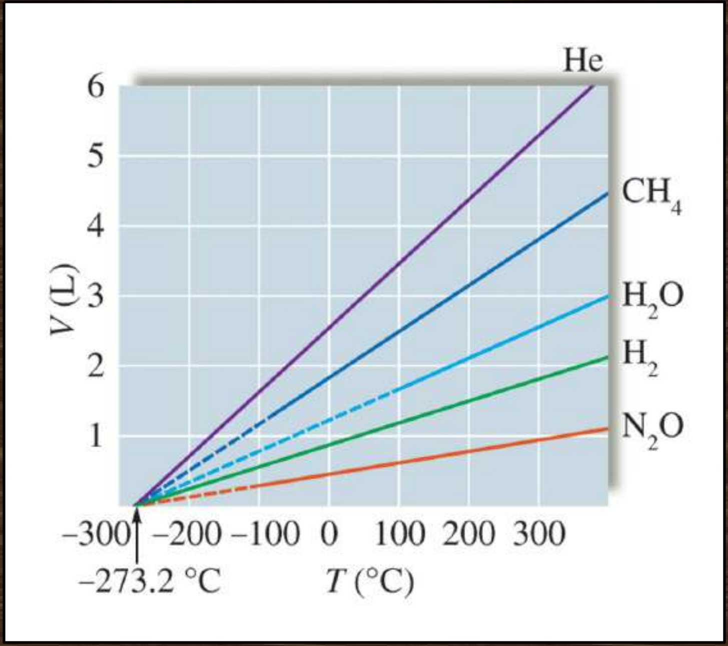


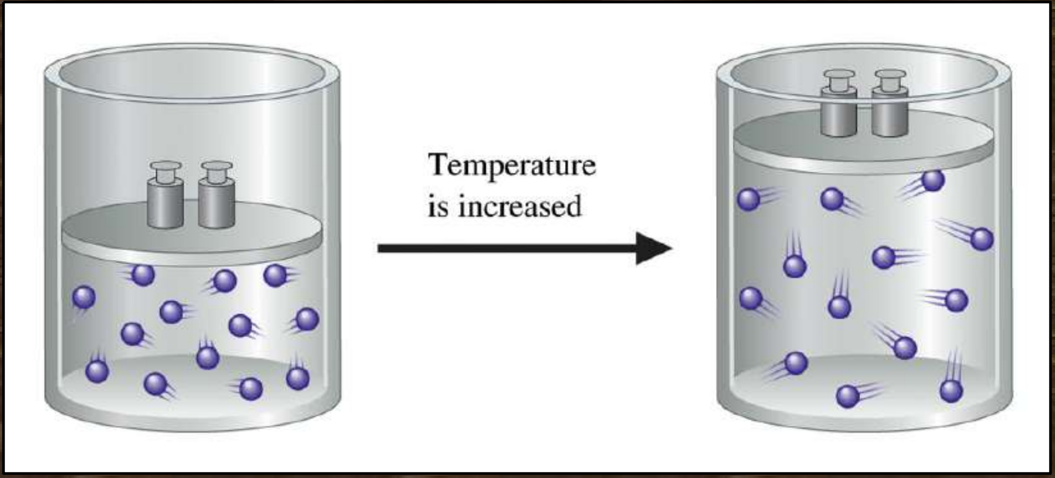


Charles's Law

Temp in Kelvin!!!!!!!!

- **Volume increases as temperature increases.**
- **Directly proportional: ½ T = ½ V4T = 4V**
- **V = bT OR V/T = b (where b is a constant)**
- **$V_1/T_1 = b = V_2/T_2$ OR$V_1/T_1 = V_2/T_2$**





Absolute Zero

Charles's Law Examples

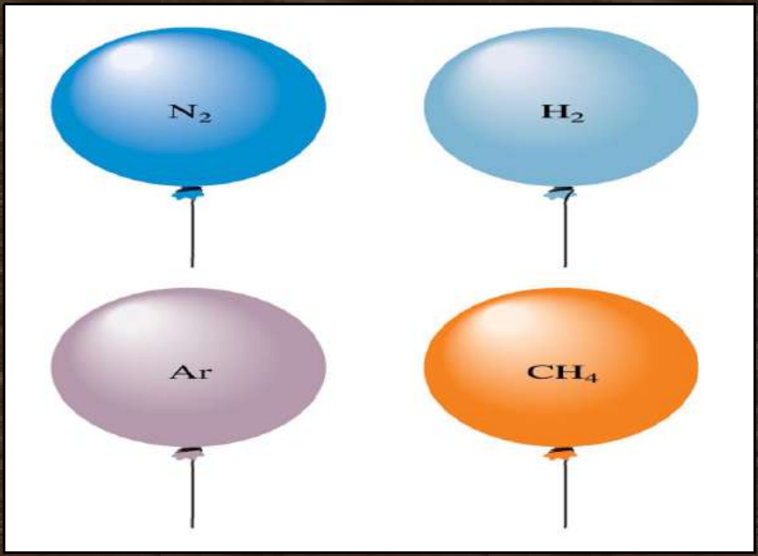
1. Hydrogen gas was cooled from 150°C to 50.$^{\circ}$C. Its new volume in 75 mL. What was its original volume?

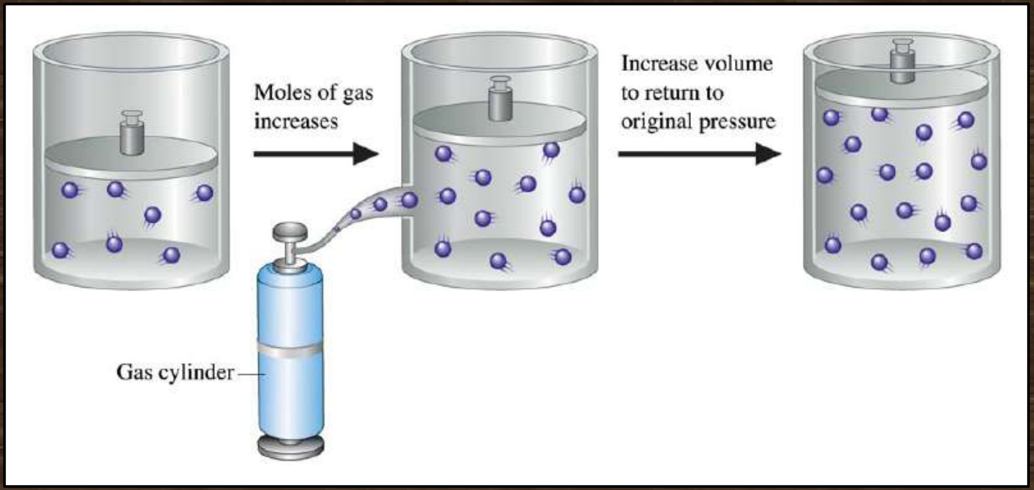
2. A sample of argon gas is cooled and its volume went from 380mL to 250mL. If its final temperature was 55°C, what was its original temperature?

HW: # 34

Avogadro's Law

- Volume of gas is directly proportional to the # of moles of the gas.
- Volume increases as moles increase
- $V = an$ (n is # of moles and a is a constant)
- $V_1/n_1 = a = V_2/n_2$ OR $V_1/n_1 = V_2/n_2$





1 mole of Gas = 22.4 L of gas at STP

Avogadro's Examples

1. Suppose 5.00 L of a gas is known to contain 0.965 mol. If the amount of gas is increased to 1.80 mol, what new volume will result?
2. Ammonia is manufactured for fertilizer. The truck hauling the ammonia can hold 450 kiloliters of the gas, which is 20 moles. If a buyer only needs to purchase 225 kiloliters of the gas, how many moles is the buyer receiving? How many grams?

HW: #35, 36

Ideal Gas Law

- Boyle's Law: $V=k/P$
- Charles's Law: $V=bT$
- Avogadro's Law: $V=an$

- $V=(kba)nT/P$

(kba) = Univ Gas Constant (R)

- **$PV=nRT$**

$R = 0.08206\ L \cdot atm / mol \cdot K$

Ideal Problems 0.08206 L·atm/mol·K

1. What volume will 1.50 moles of nitrogen occupy at 720 torr and 20.0 ^{o}C?
2. What pressure will be exerted by 45 grams of CO_2 held in a ½ liter container 25^{o}C?
3. At what temp. will 35.5 g of chlorine gas exert a pressure of 900. torr at a volume of 750. mL?
4. A sample of gas occupies 3.0 L at 850 mmHg and 130^{o}C. The gas is transferred to a 5.0 L container where it exerts a pressure of 690 mmHg. What is the temperature of the gas in the new container?

HW: #37-49 Odd

Stoichiometry…again

- Standard Temperature and Pressure:
  - $0^{o}C$ or 273K
  - 1 atm, 760 torr, 760 mmHg, 101.3kPa
  - 14.7 psi
- Molar Volume of Ideal Gas:
  - Volume of 1 mole of an ideal gas
  - 22.42 L at STP

Stoichiometry Problems

1. 4.0g of hydrogen gas at 0^{o}C and 760torr occupies what volume?

$$H_{2\,(g)} \quad + \quad Cl_{2\,(g)} \quad ===> \quad HCl_{\,(g)}$$

2. What volume of gaseous reactants, at STP, are needed to produce 156g of hydrogen chloride?
3. If liquid carbon disulfide reacts with 750 mL of oxygen gas to produce carbon dioxide and sulfur dioxide gases, what is the volume of each product?

HW: #51, 53, 59

Molar Mass of a Gas

- If we know the density of a gas, we can calculate its molar mass. OOOHHH!!!!
- PV=nRT
- n=m/MM (mass/molar mass in g/L)
- PV=mRT/MMORP=mRT/V·MM
- We know that density(d)=m/V
- P=dRT/MMORMM=dRT/P

- Do not memorize: Try to follow the math so you can derive on your own.

Molar Mass Problems

1. An unknown diatomic gas is known to have a density of 1.428 g/L at STP. What is the identity of the gas?
2. A compound has the empirical formula CHCl. A 256 mL flask, at 100.^{o}C and 750. torr, contains 0.800 g of the gaseous compound. Give the molecular formula. (#62 in text)
3. Determine the density of methane (CH_4) at 53.0 ^{o}C and 925 mmHg.

HW: #61, 63

Dalton's Law of Partial Pressure

- Total pressure of a mixture of gases is equal to the sum of the individual pressures.

$$P_{total} = P_1 + P_2 + P_3 + \ldots$$

- P_1, P_2, P_3 represent the PARTIAL PRESSURE of each of the gases (gas 1, gas 2, gas 3, …)
- Since **$P=nRT/V$** and RT/V is the same for all
- **$P_{total} = P_1 + P_2 + P_3 + \ldots = (n_1 + n_2 + n_3\ldots)(RT/V)$**

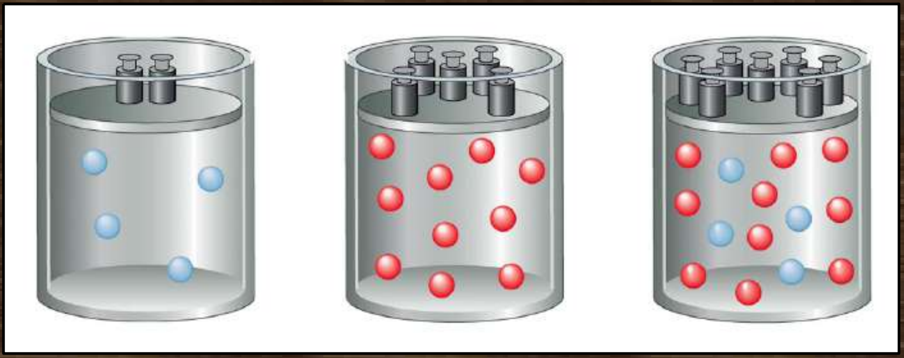
  $= n_{total}\,(RT/V)$

- Don't Freak Out!!! We shall go slow.

Partial Pressure Problems

1. Three gases are all held in the same container. If the pressure of gas 1 is 1.5 atm, gas 2 is 4.5 atm and the total pressure is 9.0 atm, what is the pressure of gas 3?
2. A mixture of 14.0 g of nitrogen gas and 48.0g of oxygen gas is placed in a 1.00L container at 27°C. Calculate the partial pressure of each gas and the total pressure of the mixture.
3. #68 from textbook

HW: # 65-67



Mole Fractions

- Ratio of moles of a thing to the total moles.
- $c_1 = \dfrac{n_1}{n_{total}} = \dfrac{n_1}{n_1 + n_2 + n_3 \ldots}$
- Since $PV=nRTn= PV/RT = P(V/RT)$
- $n_1 = P_1(V/RT)n_2 = P_2(V/RT)$
- $c_1 = n_1/n_{total} = P_1(V/RT)/\ P_1(V/RT)+P_2(V/RT)+\ldots$

  $=$ (V/RT) cancels leaving

$n_1/n_{total} = P_1\ /\ P_{total}$

Show on Board

Mole Fraction Problems

The partial pressure of hydrogen is 1.50 atm and the partial pressure of nitrogen is 2.50 atm in a mixture of the two gases

1. What is the mole fraction of each gas in mixture?
2. If the mixture occupies 15.0 L at 27.0 ^{o}C, calculate the total number of moles.
3. How many grams of each gas are present?

HW: # 69, 70

More Mole Fraction Problems

$c_1 = n_1/n_{total} = P_1/\ P_{total}$ (c$_1$ = mole fraction)

$c_1 = P_1/P_{total}$ OR $P_1 = c_1 \cdot P_{total}$

1. The mole fraction of fluorine in a mixture of gases at 900. torr is 0.50. Calculate the partial pressure of fluorine.
2. The mixture also contains oxygen (c_O = 0.333) and nitrogen. What are the partial pressures of all 3 gases when the total pressure is 600.mmHg?

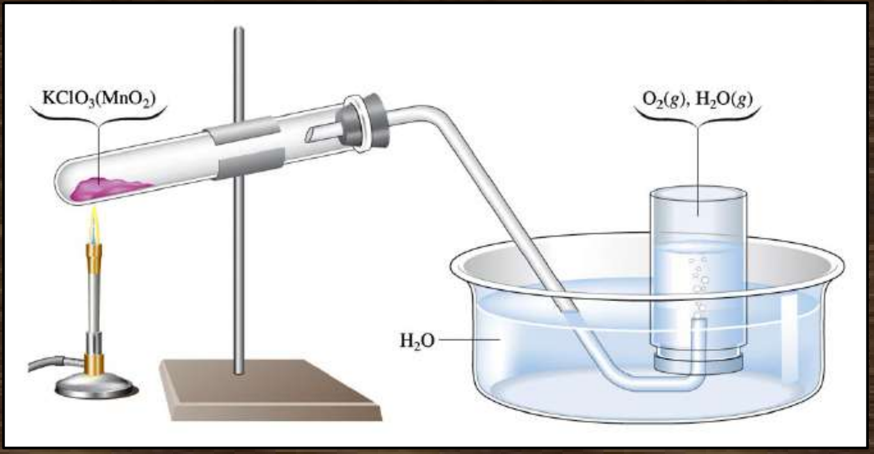
HW: # 75, 76

Collecting Over H_20 / H_20 Displacement

- Partial pressure problem where one of the gases present is water vapor.
- Example: Helium is collected over water at 25°C and 1.00 atm. What volume of gas must be collected to obtain 1.00g of helium?

(The vapor pressure of H_2O at 25°C is 23.8torr)

HW: # 71, 73



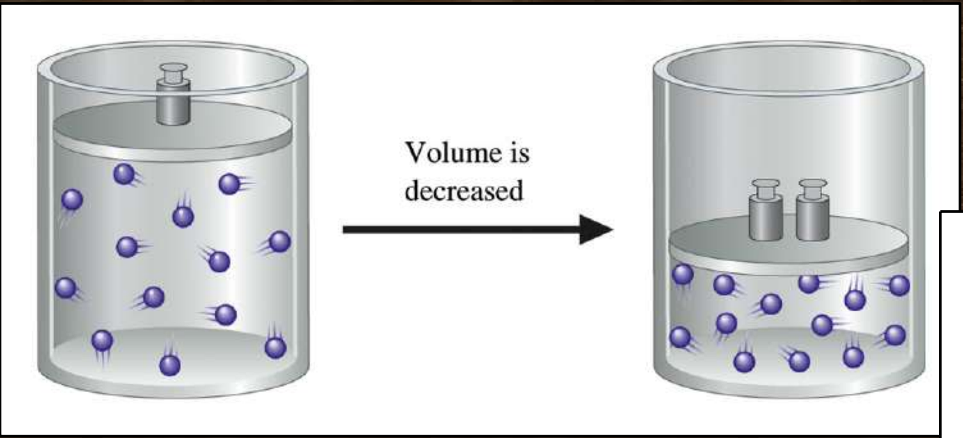
Kinetic Molecular Theory (KMT)

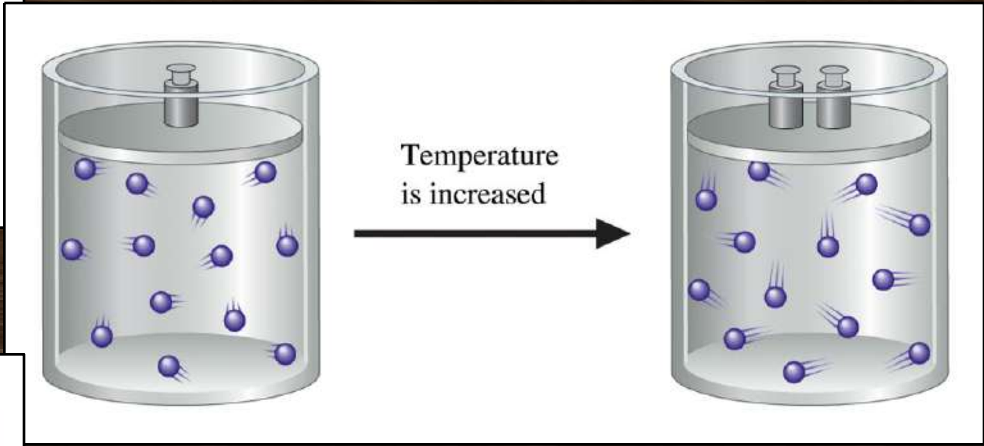
Simple model used to explain IDEAL gas behavior.

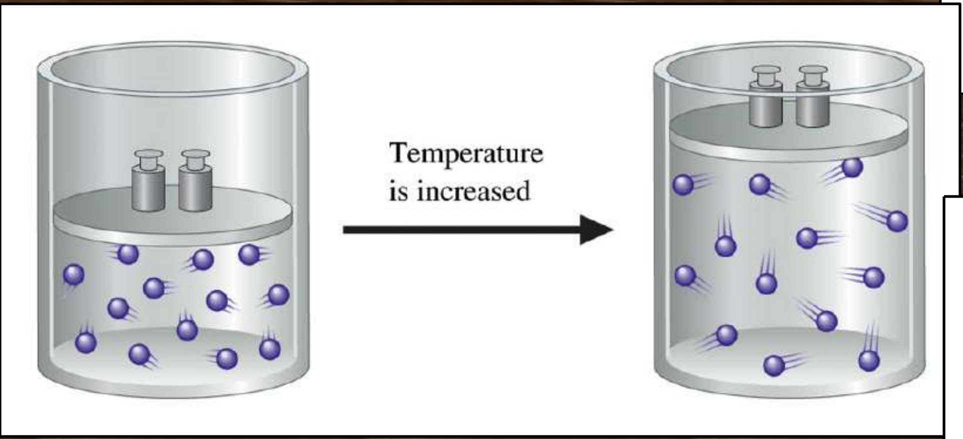
4 assumptions of KMT (you know what they say)

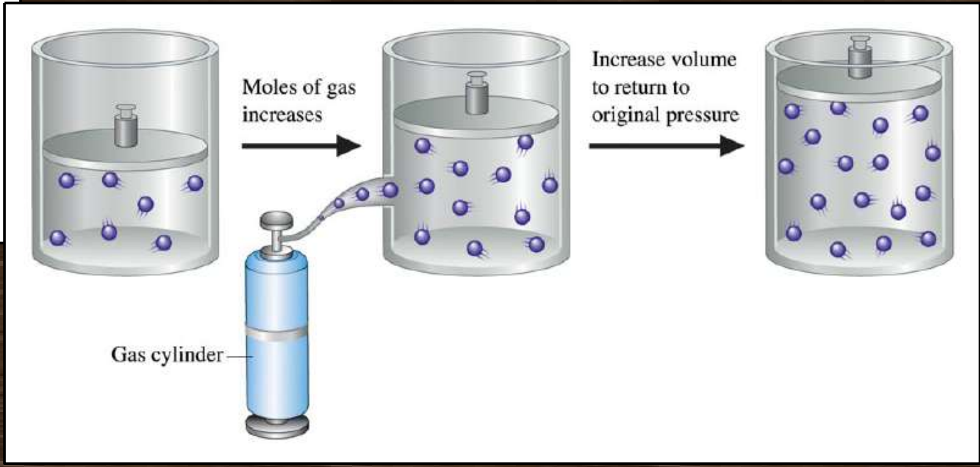
1. Particles are so small we ignore their individual volume.
2. Particles in constant motion. Collisions w/ container cause the exerted pressure.
3. Particles exert no forces on each other.
4. Average Kinetic Energy of a gas is directly proportional to Kelvin Temp.

How KMT Accounts For Stuff





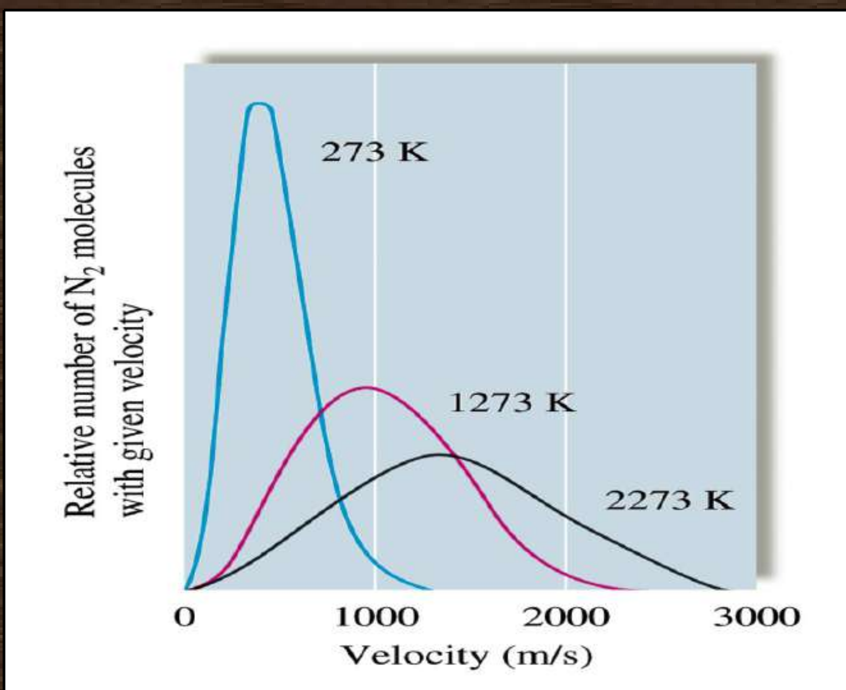


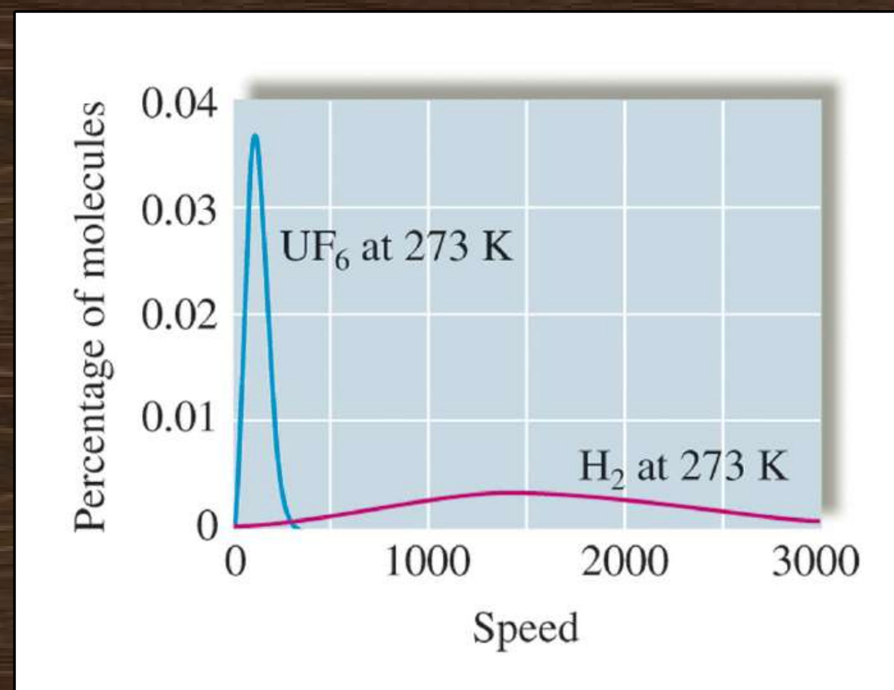


HW: # 81, 82

What Temp Means

- **An index of the random motions of the particles**
- Relates to kinetic energy (KE) with equation
- $(KE)_{avg} = 3/2\ RT$ Don't worry about this for now
- Higher temps mean higher average KE





Root Mean Square u_{rms}

- Square root of the average velocity squared
- $U_{rms} = \sqrt{(\frac{3RT}{M})}$

- **M = Molar mass in kg**
- **R = 8.31 J mol^{-1} K^{-1}**
- **J = kg · m^{2} sec^{-2}**

LOOKING AT UNITS

- $U_{rms} = \sqrt{\frac{J\ mol^{-1}\ K^{-1} \cdot K}{kg\ mol\text{-}1}} = \sqrt{J/kg} = \sqrt{\frac{kg\ m^2}{kg\ s^2}} = m/s$

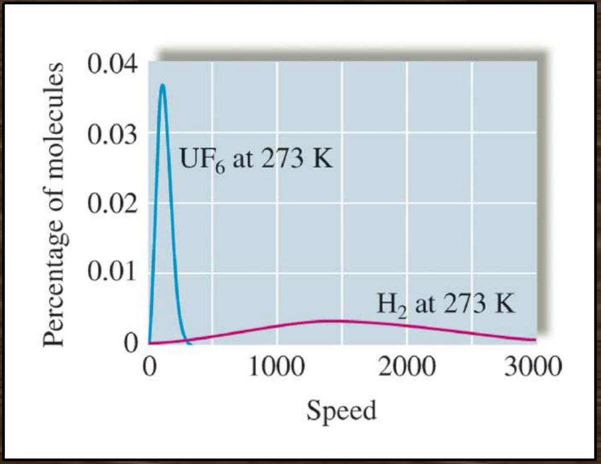
u_{rms} Problems

1. Calculate the root mean square velocities of O_2 and H_2O at 100K and 200K.
2. Separate samples of H_2 and N_2 gas are in held is 1.0L containers at STP. Which sample has the greatest average kinetic energy? Which gas has the greatest average velocity?

HW: # 79, 80

Effusion and Diffusion

- Diffusion – Gas moving from high to low concentration
- Effusion – The passage of a gas through a tiny orifice (hole) into an empty chamber
- Both are related to the velocity of the particles
- Therefore both are related to the molar mass of the particles in question.



Problems

$$\frac{\text{Effusion rate 1}}{\text{Effusion rate 2}} = \frac{\sqrt{M_2}}{\sqrt{M_1}}$$ = **Graham's Law of Effusion**

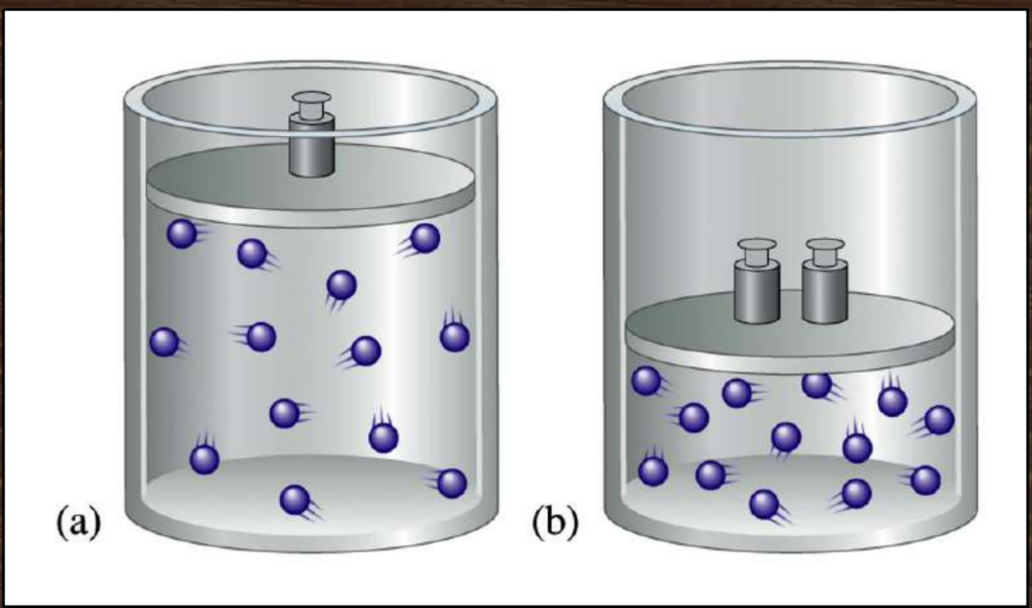
1. Calculate the ratio of effusion rates for H_2 and N_2.

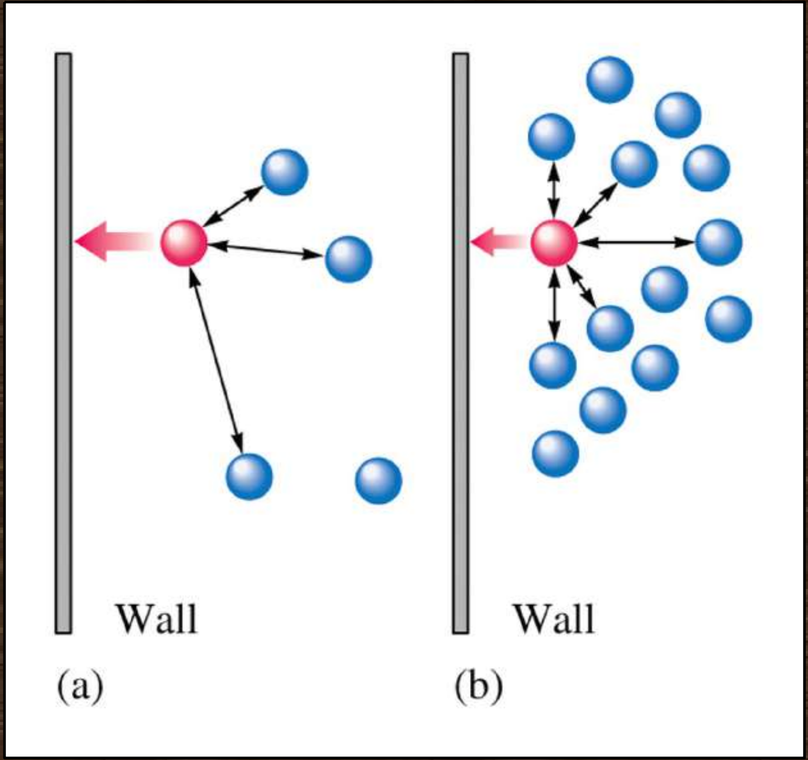
2. If the rate of effusion of O_2 is measured under certain circumstances to be 25.0 mL/min, calculate the rate of effusion of CH_4

HW: # 87, 88, 86

Real Gases

- Real gases are not IDEAL.
- Real gases have volume and do interact with each other.





Real Gases

- The volume of individual particles is going to become important when the volume is small or the pressure is high.
- Interactions among particles (intermolecular forces) are going to become more important when the amount (mole) of gas is high or the temperature is low.

250. mL of an unknown gas is collected over water at 25°C and 1.50 atm. (Vapor pressure of water at 25°C is 23.8 torr)

1. What is the partial pressure of the unknown gas?
2. How many moles of the gas are present?
3. If 0.421 grams of the gas are present, calculate the molar mass of the gas.
4. The gas is diatomic, what is its identity?
5. What is the root mean square of the gas at 400K?
6. What is the rate of effusion for the unknown gas if the rate of effusion for butane (C4H10) is 60.0mL/min