Chemistry 6A F2007

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Monday

10/29/07

Exam 2: Friday 11/2/07 (here in lecture)

What will be covered on the exam?

- Chapter 4: (4.6-4.9 and 4.11)
- Chapter 5: All
- Chapter 6: (6.1-6.8)
- Any thing from lab as well

What do I need to bring?

Bring a Pencil, Eraser, Calculator and scamtron form 882

YOU NEED TO KNOW YOUR LAB SECTION NUMBER!

Combining Avagodro's Law with the general gas law…

This in known as the "***Ideal Gas Law***"

$$\frac{P \times V}{n \times T} = \text{constant}$$

$$P \times V = n \times R \times T$$

$$R = \text{"gas constant"} = 0.08206 \frac{L \cdot atm}{mol \cdot K}$$

$$PV = nRT$$

Rules for Ideal Gas Law Calculations:

- Always convert the temperature to Kelvin (K = 273.15 + ^{o}C)
- Convert from grams to moles if necessary.
- Be sure to convert to the units of "R" (L, atm, mol & K).

Types of Ideal Gas Law problems you may encounter:

- Determination one unknown quantity of one gas variable (P, V, T, or n) given the other values directly or indirectly.
- Determine the new values of P, V, T, or n after a gas undergoes a change.
- Stoichiometry problems.
- Gas density and molar mass problems.

Calcium carbonate decomposes upon heating to form calcium oxide and carbon dioxide. If a 0.250g sample of calcium carbonate is heated to 250°C in a 1.55L vessel, what is the pressure in torr?

Step 1: Recognize that the carbon dioxide liberated is a gas.

$$CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$$

Step 2: Determine the moles of CO_2 that form from the $CaCO_3$

$$0.250\cancel{g\ CaCO_3} \times \frac{1\ \cancel{mol\ CaCO_3}}{100.1\cancel{g\ CaCO_3}} \times \frac{1\ mol\ CO_2}{1\ \cancel{mol\ CaCO_3}}$$

$$= 0.00250\ mol\ CO_2$$

Calcium carbonate decomposes upon heating to form calcium oxide and carbon dioxide. If a 0.250g sample of calcium carbonate is heated to 250°C in a 1.55L vessel, what is the pressure in torr?

$$PV = nRT \dashrightarrow P = \frac{nRT}{V}$$

$$P = \frac{0.00250\ \cancel{mol} \times 0.0821\frac{\cancel{L}\cdot atm}{\cancel{mol}\cdot \cancel{K}} \times (250.0 + 273.15)\cancel{K}}{1.55\cancel{L}}$$

$$= 0.0692\ atm$$

$$0.0692 \times \frac{760torr}{1atm} = 52.6\ torr$$

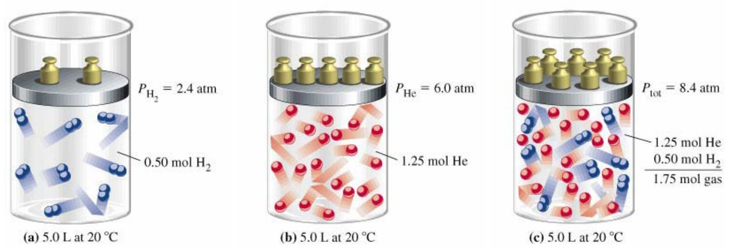
Dalton's Law of Partial Pressure

Daltons Law of partial pressures states that:

"The total pressure of a system of gases is the sum of the partial pressure of each individual gas"

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

The total pressure of a system of gases is equal to the sum of the individual pressures.



(a) 5.0 L at 20 °C **(b)** 5.0 L at 20 °C **(c)** 5.0 L at 20 °C

P = 2.4 atm

P = 6.0 atm

P = 2.4 atm
+ 6.0 atm
= 8.4 atm

Calculate the pressure in a 2.0L tank that contains 1.60g of hydrogen and 15.6g of oxygen at 32.0 °C.

$$P_{\text{total}} = \frac{n_{H_2}RT}{V} + \frac{n_{O_2}RT}{V}$$

Since both gases are in the same container, one can write:

$$P_{\text{total}} = \frac{(n_{H_2} + n_{O_2}) \times RT}{V}$$

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

Calculate the pressure in a 2.0L tank that contains 1.60g of hydrogen and 15.6g of oxygen at 32.0 °C.

Solving for the moles of each:

1.60g H_2 = 0.792 mols 15.6g O_2 = 0.488 mols

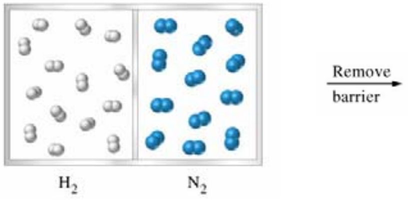
n_{total} = 1.280 mols

$$P = \frac{1.280\ \cancel{\text{mols}} \times 0.0821 \frac{\cancel{L} \cdot \text{atm}}{\cancel{\text{mol}} \cdot \cancel{K}} \times (32.0 + 273.15)\cancel{K}}{2.0\cancel{L}}$$

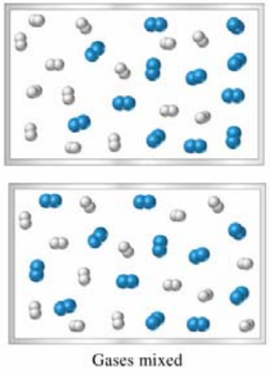
= 16 atm

Diffusion is the process of gas migration due to the random motions and collisions of gas particles. It is diffusion that results in a gas completely filling its container

gases separated

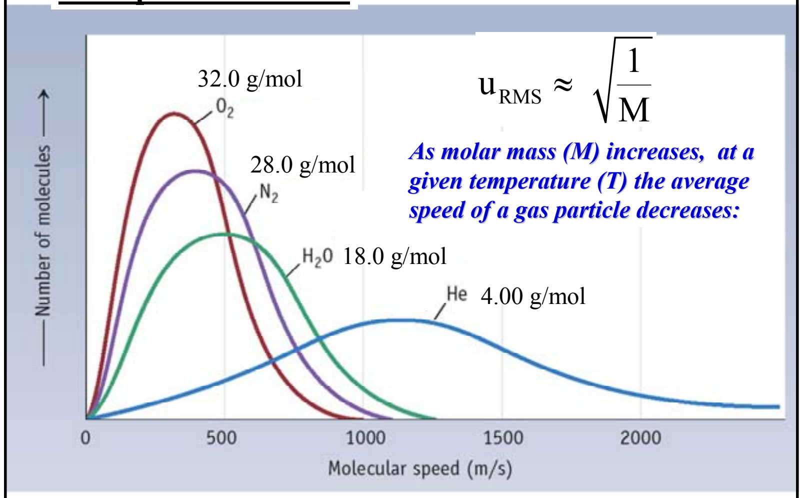


gases begin to mix



over time, the mixture becomes homogeneous.

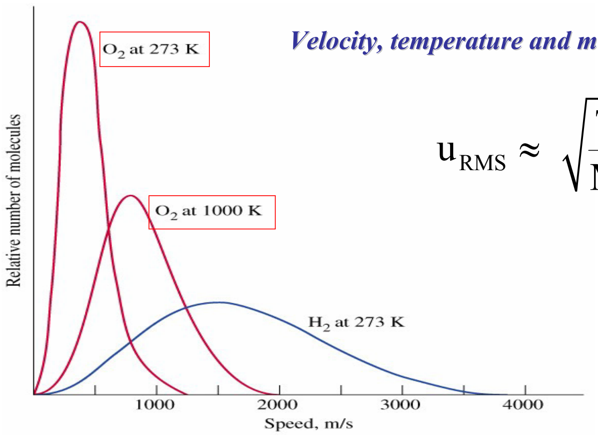
The Speeds of Gasses:



$$u_{RMS} \approx \sqrt{\frac{1}{M}}$$

As molar mass (M) increases, at a given temperature (T) the average speed of a gas particle decreases:

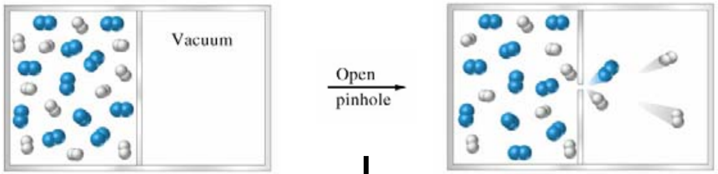
As the temperature increases, the average speed of a gas increases and the distribution broadens as is does for lighter massed particles.

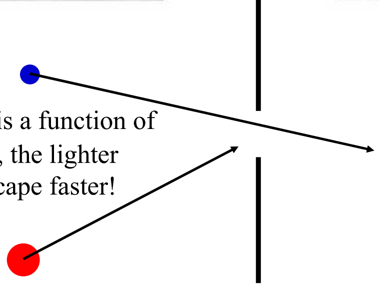


Velocity, temperature and molar mass

$$u_{RMS} \approx \sqrt{\frac{T}{M}}$$

<u>Effusion</u> *<u>is the escape of a gas through an orifice or "pin hole".</u>*





since u_{RMS} is a function of molar mass, the lighter particles escape faster!

Effusion is governed by ***Graham's Law:***

The rate of effusion of a gas is proportional to its u_{RMS}.

$$\text{Rate} \approx u_{RMS} \approx \sqrt{\frac{T}{M}}$$

Where M is the molar mass of a substance.

This implies that heavier gases will effuse slower than lighter gases.

So knowing the rate of effusion of one known gas, one can determine the molar mass of an unknown gas based on its effusion rate.

$$\frac{\text{effusion rate of A}}{\text{effusion rate of B}} = \sqrt{\frac{\text{molecular mass of B}}{\text{molecular mass of A}}}$$

Example: Carbon dioxide effuses through a pinhole at a rate of 0.232 ml/min. Another gas effuses at a rate of 0.363 ml/min. What is the molar mass of the gas?

$$\text{Rate} \propto u_{RMS} \propto \sqrt{\frac{T}{M}}$$

Comparing the rate of effusion of CO_2 vs. the unknown gas:

$$\frac{\text{Rate}_{CO_2}}{\text{Rate}_{unk.}} = \frac{u_{CO_2}}{u_{unk.}} = \frac{\sqrt{\frac{\cancel{T}}{M_{CO_2}}}}{\sqrt{\frac{\cancel{T}}{M_{unk.}}}} = \sqrt{\frac{M_{unk.}}{M_{CO_2}}}$$

Example: Carbon dioxide effuses through a pinhole at a rate of 0.232 ml/min. Another gas effuses at a rate of 0.363 ml/min. What is the molar mass of the gas?

$$\frac{\text{Rate}_{CO_2}}{\text{Rate}_{unk.}} = \sqrt{\frac{M_{unk.}}{M_{CO_2}}}$$

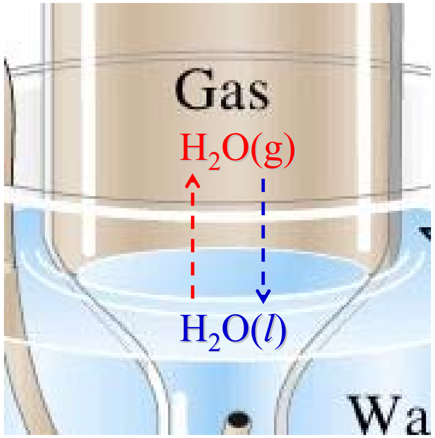
Solving for the molar mass of the unknown gas:

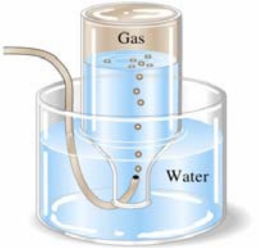
$$M_{unk.} = M_{CO_2} \times \left(\frac{\text{Rate}_{CO_2}}{\text{Rate}_{unk.}}\right)^2$$

$$M_{unk.} = 44.0\frac{g}{mol} \times \left(\frac{0.232\frac{ml}{min}}{0.363\frac{ml}{min}}\right)^2 = 18.0\frac{g}{mol}$$

Water!

Vapor Pressure of a Liquid:





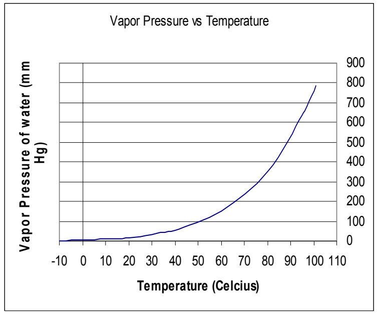
When a gas is collected over water, in addition to the gas in the container, there is water in the gas phase!

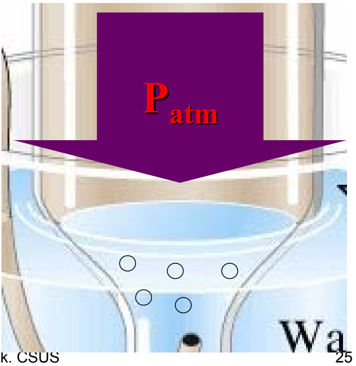
Some of the $H_2O(l)$ escapes in to the gas phase establishing a water-vapor equilibrium.

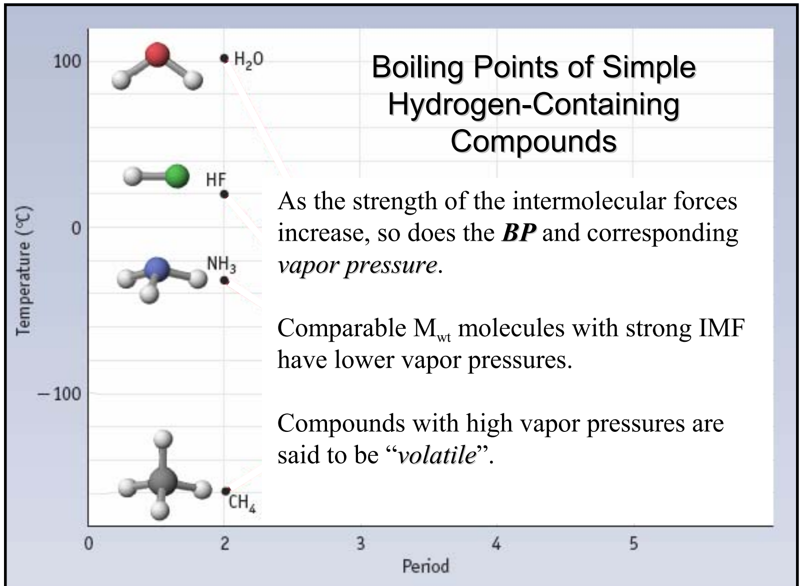
Vapor Pressure of a Liquid:

The pressure exerted by the liquid is know as the "*Vapor Pressure*"

The vapor pressure of a liquid increases with temperature until it reaches it's normal boiling point, where it equals the pressure of the atmosphere pushing down upon the surface.







Boiling Points of Simple Hydrogen-Containing Compounds

As the strength of the intermolecular forces increase, so does the ***BP*** and corresponding *vapor pressure*.

Comparable M_{wt} molecules with strong IMF have lower vapor pressures.

Compounds with high vapor pressures are said to be "*volatile*".