- The **law of combining volume** states:
  - When gases react at the same temperature and pressure, their combining volumes are in ratios of simple, whole numbers
    - Example:

hydrogen + chlorine → hydrogen chloride

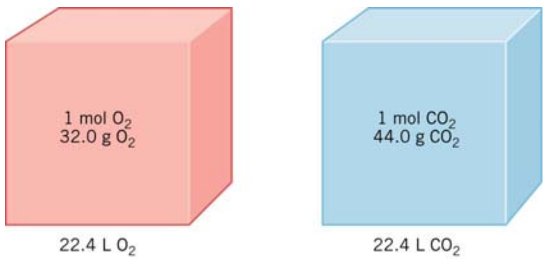
1 volume 1 volume 2 volumes

- Amedeo Avogadro studied this and devised **Avogadro's principle**:
  - When measured at the same temperature and pressure, equals volumes of gases contain equal number of moles

- A corollary to Avogadro's principle is:
  - The volume of a gas is directly proportional to its number of moles, *n*

$V \propto n$ (at constant T and P)

- Thus, the volume of one mole of any gas at standard temperature and pressure (STP) or 0^{o}C and 1 atm is 22.4 L (a constant for all ideal gases)
- This is called the standard molar volume of a gas



- The combined gas law can be generalized to include changes in the number of moles of sample
- The **ideal gas law** is

$$PV = nRT$$

$$R = \text{universal gas constant}$$

$$= 0.0821 \frac{\text{atm L}}{\text{mol K}}$$

What pressure (in torr) is exerted by 10.0 g of O_2 in a 2.50L container at a temperature of 27 degrees Celsius?

•The molecular mass is obtained by taking the *ratio* of mass to moles, which could be determined using the ideal gas law

•Question: A chemist isolated a gas in a glass bulb with a volume of 255 mL at a temperature of 25.0 degree Celsius and a pressure of 10.0 torr. The gas weighed 12.1 mg. What is the molecular mass of this gas?

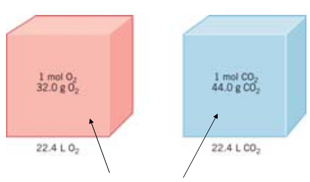
- Gas **densities** (*d*), a ratio of gas mass to volume, can be calculated by taking the ratio of the molar mass to molar volume
  - Example: The molar mass of oxygen is 32.0 g/mol. What is the density of oxygen at STP?

$$d_{O_2} = \frac{32.0 \text{ g/mol}}{22.4 \text{ L/mol}} = 1.42 \text{ g L}^{-1}$$

Molar Mass (or Molecular Mass) = $\frac{\mathbf{d\,R\,T}}{\mathbf{P}}$

Derive this equation starting with
the ideal gas law as you commonly know it…

PV = nRT



Up until now….we've been dealing with individual,
isolated gases (see above)….

But often we have MIXTURES of gases…what can we
say about these mixtures?

- One useful way to describe a composition of a mixture is in terms of its *mole fractions*
- The **mole fraction** is the ratio of the number of moles of a given component to the total moles of all components

- For a mixture of *A*, *B*, … substances, the mole fraction of substance *i* (X_i) is

$$X_A = \frac{n_A}{n_A + n_B + \ldots + n_Z}, \quad n_i = \text{moles of } i$$

- This provides a convenient way to `partition' the total pressure of a mixture of gases

A 22.4 L container at 0^{o}C contains 0.30 mol N_2, 0.20 mol O_2, 0.40 mol He, and 0.10 mol CO_2.

a) What are the mole fractions of each of the gases?

Dalton's law of partial pressures states: the total pressure of a mixture of gases is the sum of their individual partial pressures

- For a system of only gases, mole fractions and partial pressure partition the total pressure in the same fashion

$$
\begin{aligned}
P_{Total} &= P_A + P_B + ... \\
&= P_{Total} X_A + P_{Total} X_B + ... \\
1 &= X_A + X_B + ...
\end{aligned}
$$

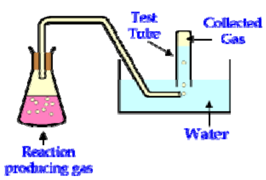
A 22.4 L container at 0°C contains 0.30 mol N_2, 0.20 mol O_2, 0.40 mol He, and 0.10 mol CO_2.

b) What are the partial pressures of each gas?

What is the mole fraction of methane in a gaseous mixture that consists of 8.00 g of methane and 12.00 g of ethane, C_2H_6, in a 3.50 liter container maintained at 35.20^oC?

http://www.cohenufo.org/Images.GIF/Mole.jpg

We are skipping the section on “collecting gases over water” …..well, for the most part!



http://www.gcsescience.com/a/Gas-Over-Water.gif

- Gas *volumes* can be used in stoichiometry problems

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$$

2 volumes   1 volume   2 volumes   (same temperature and pressure)

$$2 \text{ volumes } H_2(g) \Leftrightarrow 1 \text{ volume } O_2(g)$$
$$2 \text{ volumes } H_2(g) \Leftrightarrow 2 \text{ volumes } H_2O(g)$$
$$1 \text{ volumes } O_2(g) \Leftrightarrow 2 \text{ volumes } H_2O(g)$$

just as

$$2 \text{ moles } H_2(g) \Leftrightarrow 1 \text{ mole } O_2(g)$$
$$2 \text{ moles } H_2(g) \Leftrightarrow 2 \text{ moles } H_2O(g)$$
$$1 \text{ mole } O_2(g) \Leftrightarrow 2 \text{ moles } H_2O(g)$$

Write the balanced chemical equation for the combustion of methane, CH_4. All reactants and products are gases.

The combustion of 4.50 L of CH_4 consumes how many liters of O_2, assume both volumes measured at 25^{o}C and 740 torr?

STOICHIOMETRIC CALCULATION
(when some reactants and products are not gases)

Nitric acid is formed when NO_2 is dissolved in water:

$3NO_2$ (g) + H_2O(l) → $2HNO_3$(aq) + NO(g)

How many liters of NO_2 at 25^{o}C and 752 torr are needed to form 12.0 g of HNO_3?

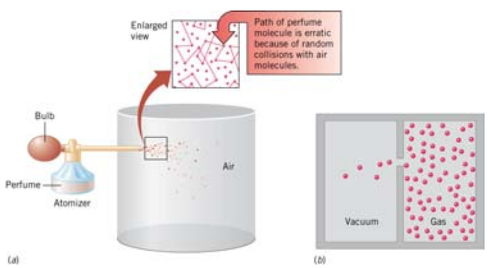
- **Diffusion** is the spontaneous intermingling of the molecules of one gas with another
- **Effusion** is the movement of gas molecules through a tiny hole into a vacuum
- The rates of both diffusion and effusion depend on the speed of the gas molecules
- The faster the molecules, the faster diffusion and effusion occur
- Thomas Graham studied effusion

- He found that the effusion rate of a gas was inversely proportional to the square root of the density (d)
- This is known as **Graham's law**

$$\text{effusion rate} \propto \frac{1}{\sqrt{\text{d}}} \qquad (\text{constant } P \text{ and } T)$$

$$\frac{\text{effusion rate }(A)}{\text{effusion rate }(B)} = \sqrt{\frac{d_B}{d_A}} = \sqrt{\frac{M_B}{M_A}}$$

- Where M_i is the molar mass of species i



(a) Diffusion (b) Effusion

Kinetic Theory of Gases

1. Gas particles are tiny and in constant, random motion

2. Gas particles occupy a VERY small volume in comparison to their container

3. Particles collide in perfectly elastic collisions, move in straight lines between collisions, and do not attract/repel each other

PV = nRT (The Ideal Gas Law)

Assumes the three postulates of the Kinetic theory hold true!

IS THIS “REAL”?

We’ll at LOW Temps and/or High Pressure…NO

Must use the van der Waals Equation (corrects for deviations From ideal behavior)