Gases

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Gas particles are very far apart -hundreds to thousands of times their own diameters. Thus interactions are usually small and if we neglect them we will not make a large error.

This allows us to "model" the behavior of gases by defining an *<u>ideal gas</u>* which is often close to a real gas.

1. **Ideal gas particles are point masses so the particles occupy no volume themselves.**
2. **Particles have no attraction for each other.**

Kinetic Theory

Principles

- **Matter is composed of particles in constant, random, motion**
- **Particles collide elastically if they do not react**
- **All particles at a given temp do not have the same KE**

Solids and liquids have particles that are relatively close together.

<u>Intermolecular forces and shapes of molecules or ions</u> are very important in determining properties of a solid.

<u>Intermolecular forces</u> are very important in determining properties of a liquid.

Gas Laws

Measurement- 4 variables specify the <u>state</u> of a gas -

- **Pressure - many units are used (we shall look at this next)**
- **Volume - m^3, liters**
- **Number of particles - moles**
- **Temperature - Kelvin (we shall see why soon!)**

<u>Pressure</u> = Force/Area

- **Units - lb/in^2, N/m^2= Pascal, bar =10^5 Pa, mm Hg=Torr, inches of Hg, Atmospheres**
- **It is more difficult to measure gas pressure in force/area units. That is why we often measure pressure in terms of what pressure can do - Torr, in Hg**
- ***<u>Manometers</u>* are used to measure the pressure of gases and are of 2 types – open and closed. *<u>Barometers</u>* are specifically designed to measure atmospheric pressure.**

- It is necessary to convert between these units. A convenient way to do this is to memorize the values for standard pressure

760. mm Hg	760. Torr	30.0 in Hg
1.00 Atm	14.7 lb/in^2	1.013 x 10^5 Pa
101.3 kPa	1.013 bar	

Temperature

The Kevin scale is used because it is an absolute temperature scale. This means that it begins at absolute zero - the point at which all translational motion would cease. For an ideal gas this means that volume would be zero since it is composed of point masses. K = ˚C + 273.15

Standard temperature is defined as 0°C or 273.15 K. "STP" represents standard temperature and pressure.

Boyle's Law $P \propto 1/V$ when n and T are constant

Charles' Law $V \propto T$ when n and P are constant

T must be expressed in Kelvin.

Ideal Gas Law Equation

PV = nRT

The constant R relates the other variables. From both experimental and theoretical work we know that 1 mole of a gas occupies 22.4 l at 273 K and 1.00 atm pressure. This gives a value of 0.0821 Atm - Liters/mole-K or 62.4 x 10^3 Torr-mL/mole-K.

PV = nRT can also be used to obtain a general use equation for changes in a gas.

$$\frac{P_1V_1}{P_2V_2} = \frac{n_1T_1}{n_2T_2}$$

We can also substitute for the (n) in PV = nRT.

n=m/M where m = mass in grams and M = molar mass in g/mole. This gives

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

Since Density = mass/volume, an expression for density can also be obtained for a gas

$$D = \frac{PM}{RT}$$

Particle speeds and Graham's Law

If different gases are at the same temperature, do they have the same average KE? do they have the same average speed?

Since temperature is a measure of KE they have the same average KE but not the same speed. In fact

$$\frac{v_1}{v_2} = \sqrt{\frac{M_2}{M_1}}$$

where v is velocity (speed) and M is molar mass.

Since distance = velocity x time we also find that

$$\frac{Dist_1}{Dist_2} = \sqrt{\frac{M_2}{M_1}} \quad \text{and} \quad \frac{t_2}{t_1} = \sqrt{\frac{M_2}{M_1}}$$

These speeds affect gaseous _diffusion_ (the gradual mixing of gases due to collisions) and _effusion_ (the process by which a gas escapes from a high pressure area to a low pressure area through a small opening). The relative rates of diffusion and effusion are the same for gases and obey the equations above.

What are the relationships between different gases in a mixture?

total moles = $n_x + n_o + n_z$

total volume = $V_x = V_o = V_z$

Temperature = $T_x = T_o = T_z$

total pressure = $P_x + P_o + P_z$

This last statement is called Dalton's Law of Partial Pressures. It means that each gas acts independently in the container and exerts its own pressure.

For each component $P_x = n_x RT/V$

For the total gas $P_{tot} = n_{tot} RT/V$

so P_x = (mole fraction of X) P_{tot}

Stoichiometry

Mass/volume or volume/mass
(1 reactant or product given)

2.02 g of hydrogen react with an excess of oxygen to produce how many liters of water vapor? The water vapor has a pressure of 300. torr and a temperature of 300. K.

Mass/volume or volume/mass (limiting reactant)
(more than one reactant or product given) -

2.02 g of hydrogen react with 8.00 g of oxygen to produce how many liter of water vapor? The water vapor has a pressure of 300. torr and a temperature of 300. K.

Volume/volume under same conditions

10.0 liters of hydrogen react with 8.00 liters of oxygen to produce how many liters of water vapor at the same T and P?

Use _Avogadro's hypothesis_ - equal volumes of gases at the same T and P contain equal numbers of molecules. This means that a ratio of volumes is the same as a ratio of moles.