Gases

Ron Robertson

r2 f:\files\courses\1110-20\2010 possible slides for web\ch12tran.doc

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.

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

Intermolecular forces are very important in determining properties of a liquid.

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.

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 <u>memorize</u> the values for <u>standard pressure</u>

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

<u>Temperature</u>

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. "<u>STP</u>" 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

$$\boxed{PV = nRT}$$

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 62.4×10^3 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.