

Chapter 5: Gases

5.1 Pressure

$$1 \text{ atm} = 760 \text{ Torr} = 760 \text{ mmHg} = 101.325 \text{ kPa}$$

$$\text{Pressure} = \text{Force/Area} = \text{N/m}^2 = \text{Pascal}$$

5.2 The Gas Laws of Boyle, Charles, and Avagadro Boyles Law 1

$$PV = k \quad (k = \text{constant}) \text{ @ constant } T^\circ$$

-there's an inverse relationship between V&P

-an "ideal gas" strictly obeys Boyle's law

$$P_1V_1 = P_2V_2$$

Boyle's law II

-PV is equal to 22.4 L atm at zero pressure for 1 mol of a gas

Charle's Law

-volume of gas is directly proportional to temperature and is zero when temperature is 273.15°C (0° K)

$$K = C^\circ + 273.15$$

$$V \propto T$$

$$V = bT \quad (b \text{ is proportionality constant})$$

(F.Y.I>).000001 K has been reached

$$V/T = b \text{ so } V_1/T_1 = V_2/T_2$$

(T must be in Kelvin)

Avogadro's Law

- equal volumes of gases contain the same number of particles at same T°&P

$$V = an \quad (n = \# \text{ of moles, } a = \text{proportionality constant})$$

$$V \propto n \text{ @ (constant } T^\circ \text{ \& } P)$$

$$V/n = a, \text{ so } V_1/n_1 = V_2/n_2$$

5.3 The Ideal Gas Law

$$\text{Boyle's law: } V = k/P$$

$$\text{Charle's law: } V = bT$$

$$\text{Avogadro's law: } V = an$$

$$V = R(Tn/P) \quad R = \text{universal gas constant}$$

$$R = 0.08206 \text{ L atm/K mol}$$

-this gas equation works the best for ideal gases at low temperatures

-(not to yourself) explain how the Ideal Gas Law can be used to derive Boyle's, Charle's, and the combined gas laws.

5.4 Gas Stoichiometry

$$V = nRT/P = (1.000 \text{ mol})(0.08206 \text{ L atm/K mol})(273.2\text{K})/1.0000 \text{ atm} = 22.42 \text{ L}$$

Molar volume of an ideal gas at STP = 22.42 L (0°C & 1 atm)

Molar Mass of a Gas

$$-N = \text{grams of gas/molar mass} = \text{mass/molar mass} = m/\text{molar mass}$$

$$-P = nRT/V = (m/\text{molar mass})TR/V = m(RT)/V \text{ (molar mass)}$$

$$-m/V = d, \text{ so } P = dRT / \text{molar mass}$$

5.5 Dalton's law of Partial Pressure

- total pressure is the sum of all pressures exerted by each gas with in a mixture
- $P_{\text{total}} = P_1 + P_2 + P_3 + \dots$
- $P_{\text{total}} = n_1RT/V + n_2RT/V + n_3RT/V + \dots$
- $= n_{\text{total}} (RT/V)$ (Independent of the identity of each gas)

mole fraction – ratio of the number of moles of a given component of a gas to the total # of moles within mixture

χ (chi) is the symbol used

$$n = P(V/RT) \quad X = n_1/n_1 + n_2 + n_3 \dots$$

$$\chi = P_1(V/RT)/P_1(V/RT) + P_2(V/RT) + P_3(V/RT)$$

$$\chi = P_1 / P_{\text{total}}$$

5.6 The Kinetic Molecular Theory of Gases

(KMT) = attempts to explain the behavior of an ideal gas

→4 postulates←

1. Volume of the individual gas particles can be assumed to be negligible because they are so small compared to the space between them.
2. Pressure is caused by the collisions of gas particles with the walls of a container (particles are in constant motion).
3. Particles are assumed to exert no forces one ach other; they are assumed to neither attract nor repel each other.
4. The average kinetic energy of gas particles is assumed to be directly proportional to its Kelvin temperature.

Root Mean Square Velocity (rms)

$$U_{\text{rms}} = \sqrt{u^2}$$

$$(KE)_{\text{Avg}} = N_A(1/2 mu^2) = 3/2 RT$$

$$u^2 = \frac{3RT}{NAM} \text{ then } U_{\text{rms}} = \sqrt{u^2} = \sqrt{\frac{3RT}{NAM}} = \frac{\sqrt{3RT}}{\sqrt{M}}$$

M = molar mass in Kg, m = mass of particle in Kg, N_A = mole of particle

R = 8.3145 J/K mol, J=Kg m²/s²

- there are numerous collisions in real gases and the average distance a particle travels between collisions in a gas sample is extremely small (mean free path)
- ex. O₂ 1*10⁻⁷ m at STP
- there are various molecular velocities of a gas within a sample and is dependent on temperature; the range and mean velocities increase with increased value of temperature.

5.7 Effusion and Diffusion

Diffusion- the mixing of gases

Effusion- the passage of a gas through a small hole into an evacuated chamber

Graham's law of effusion- rate of effusion is inversely proportional to the square root of the mass of its particles.

$$\text{(experimental)} \quad \frac{\text{rate of effusion for gas 1}}{\text{rate of effusion for gas 2}} = \frac{\sqrt{M_2}}{\sqrt{M_1}}$$

KMT supports Graham's law of effusion which directly affects effusion rate

$$\text{(theory)} \quad \frac{v_{\text{gas 1}}}{v_{\text{gas 2}}} = \frac{U_{\text{rms gas 1}}}{U_{\text{rms gas 2}}} = \frac{\sqrt{3RT/M_1}}{\sqrt{3RT/M_2}} = \frac{\sqrt{M_2}}{\sqrt{M_1}}$$

Diffusion

- There is no diffusion law due to the numerous collisions of mixing gases, which makes it quite complicated to describe quantitatively.

5.8 Real Gases

- KMT makes 2 drastic assumptions that sets up the Ideal Gas Law for failure (no particle interactions & zero volume for gas particle)
- Real gases exhibit ideal gas behavior at low pressure and high temperatures (fig 5.24 – 5.25)

$$P = \frac{nRT}{V-nb}$$

→ volume of real gases are smaller than ideal gases because the particles themselves take up space.

Real gas volume = $V-nb$

(n = # moles of gas and b is the correction factor that fits experimental results, V = volume of ideal gas)

- we must account for attractions between particles
- $P_{\text{obs}} = (P_{\text{IG}} - \text{correction factor}) = \frac{nRT}{V-nb} - \text{correction factor}$
- We have to now determine the size of the correction factor which is dependent on the concentration of gas particles (n/V)
- The interacting pairs is dependent on the square of the # of particles and thus the square of the concentration ($(n/V)^2$) (this is for every large #'s)

$$P_{\text{obs}} = P - a(n/V)$$

A = proportionality constant (fits observed behavior of that gas)

$$P_{\text{obs}} = \frac{nRT}{V-nb} - a(n/v)^2 - \text{pressure correction}$$

$$\text{Van der Waals equation } [P_{\text{obs}} + a(n/V)^2] * (V-nb) = nRT$$

- values of a & b are different for each gas (table 5.3)
- at large volumes (low pressure) gases behave ideally
- at small volumes (high pressure) gases don't behave ideally because particle volume is significant
- the size of volume correction constant b generally increase with increased particle size.
- At high temperatures are moving so rapidly that the effects of interparticle attraction is negligible.

5.9 Chemistry in the Atmosphere

- atmosphere contains primarily N_2 & O_2
- heavier molecules are found near the earth's surface due to gravity
- atmosphere is divided based on temperature

- the upper atmosphere prevents high energy radiation from reaching the earth's surface
- the troposphere (lowest layer) is influenced by activities by man
- Common pollutants by the Combustion of Petroleum

CO NO

CO₂ NO₂

SO₂ unburned hydrocarbons

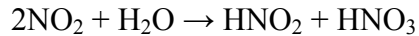
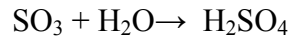
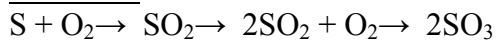
-light can have a significant effect on a mixture of polluted air

O₃ can be produced

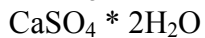
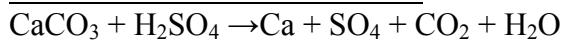
HNO₃ can be produced

And the result is photochemical smog

Acid Rain

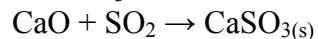


Limestone/Marble Deterioration



Gypsum

Scrubbers (Acid Rain Prevention)



waste