AP Chemistry
Chapter 10 Outline

a) Characteristics of Gases
   i) Monatomic gases = noble gases He, Ne, Ar, Kr, Xe
   ii) Substances that are liquids or solids under ordinary conditions can also exist in the gaseous state, when they are referred to as vapors
   iii) Properties of gases—because the individual particles are relatively far apart.
      (1) A gas expands spontaneously to fill its container.
      (2) Gases are highly compressible.
      (3) Gases form homogeneous mixtures with each other, regardless of the identities or relative proportions of the component gases.

b) Pressure
   i) Pressure = the force, F, that is applied on a given area, A
   ii) $P = \dfrac{F}{A}$
   iii) Gases exert a pressure on any surface with which they are in contact.
      (1) Atmospheric pressure can be measured with a barometer.
      (2) The pressure of a confined gas can be measured with a manometer.
   iv) Units of pressure
      (1) 1 atm = 760 mm Hg = 760 torr = 1.01325 x 10^5 Pa = 101.325 kPa

c) The Gas Laws
   i) Boyle's Law: Pressure and volume are inversely related, for a fixed sample of gas at constant temperature.
      (1) $P_1V_1 = P_2V_2$
      (2) The graph of V vs. P shows an inverse relationship curve.
      (3) The graph of V vs. 1/P is linear.
   ii) Charles' Law: The volume of a fixed amount of gas maintained at constant pressure is directly proportional to its absolute temperature.
      (1) $\dfrac{V_1}{T_1} = \dfrac{V_2}{T_2}$
      (2) The graph of V vs. T (in Kelvin) is linear.
   iii) Avogadro's Law
      (1) The Law of Combining Volumes: At a given temperature and pressure, the volumes of gases that react with each other are in the ratios of small whole numbers.
      (2) Avogadro's Hypothesis: Equal volume of gases at the same temperature and pressure contain equal numbers of molecules.
      (3) Avogadro's Law: The volume of a gas maintained at constant temperature and pressure is directly proportional to the number of moles of the gas.

d) The Ideal-Gas Equation
   i) $PV = nRT$ know this equation—it is incredibly useful!
   ii) R = the gas constant; on the green sheet
   iii) Standard temperature and pressure (STP) = $0°C$, 1 atm Know this!
   iv) <u>One mole of an ideal gas at STP occupies 22.4 L</u>.   Know this!

e) Further Applications of the Ideal-Gas Equation
   i) Use it to calculate gas density from the molar mass, pressure, and temperature of the gas.
      (1) $Density = \dfrac{P * gfm}{RT}$ You should be able to derive this directly from the ideal gas law
   ii) The ideal gas law can also be rearranged to solve for the molar mass (gfm) of a gas
   iii) The ideal gas law is also very useful in gas stoichiometry problems

f) Gas Mixtures and Partial Pressures
   i) <u>Dalton's Law of Partial Pressures</u> = the total pressure of a mixture of gases equals the sum of the pressures that each gas would exert if it were present alone
      (1) Partial pressure = the pressure exerted by a particular component of a mixture
      (2) $P_{tot} = P_1 + P_2 + P_3 + \ldots$
      (3) The total pressure at constant temperature and constant volume is determined by the total number of moles present.
   ii) Mole fraction of gas i = $X_i = \dfrac{moles\ of\ gas\ i}{total\ moles\ of\ gas}$

      (1) $P_i = X_i P_t$
      (2) Dalton's law must be used when <u>collecting a sample of gas over water</u>, to correct for the pressure exerted by the water vapor.
         (a) $P_{tot} = P_{gas} + P_{H2O}$

g) <u>Kinetic-Molecular Theory</u>
   i) Developed over 100 year period; published in main form by Clausius
      (1) Gases consist of large numbers of particles that are in continuous, random motion.
      (2) The combined volume of all the gas particles is negligible relative to the total volume in which the gas is contained.
      (3) Attractive and repulsive forces between gas particles are negligible.
      (4) Collisions between particles are completely elastic. (Energy can be transferred between particles during collisions, but the average kinetic energy of the particles does not change, as long as the temperature of the sample remains constant.)
      (5) The average kinetic energy of the particles is proportional to the absolute temperature.
   ii) The pressure of a gas is caused by collisions of the particles with the walls of the container.
   iii) The absolute temperature of a gas is a measure of the average kinetic energy of its particles.
      (1) <u>Individual particles move at varying speeds</u>. This is often plotted graphically, in a Gaussian distribution.

(a) The peak of each curve represents the most probable speed.

(2) <u>As temperature increases</u>, the curve flattens out and the most probable speed rises

(3) Root-mean-square (rms) speed, μ = the speed of a particle possessing average kinetic energy

h) Molecular Effusion and Diffusion

i) The average kinetic energy of any collection of gas particles has a specific value at a given temperature.

ii) However, average kinetic energy is proportional to molar mass; since $KE = \frac{1}{2} m\mu^2$

iii) $\mu = \sqrt{\dfrac{3RT}{gfm}}$ (this equation is on your green sheet)

iv) Effusion = the escape of gas particles through a tiny hole into an evacuated space

v) Diffusion = the spread of one substance throughout a space or a second substance by random mixing

vi) <u>Graham's Law of Effusion</u> = the effusion rate of a gas is inversely proportional to the square root of its molar mass (i.e., more massive particles move at a slower velocity than less massive particles)

(1) $\dfrac{r_1}{r_2} = \sqrt{\dfrac{gfm_2}{gfm_1}}$

vii) Because of collisions, diffusion is more complicated (and slower!) than effusion.

(1) The direction of a gas molecule is constantly changing.

(2) Mean free path = the average distance traveled by a molecule between collisions

i) Real Gases: Deviations from Ideal Behavior

i) <u>All real gases fail to obey the ideal gas law to some degree</u>.

(1) Real gases do not obey ideally at high pressure. At lower pressures (below 10 atm), the deviation from ideal behavior is small.

(2) In general, the deviations from ideal behavior increase as temperature decreases.

ii) Molecules of an ideal gas are assumed to occupy no space and have no attractions for each other.

(1) Real gases do have finite volumes. ("excluded volume")

(2) Real gas particles do have attractions for each other. (intermolecular attractions, in Chapter 11)

(3) At high pressure, the excluded volume has the dominant effect on deviations.

(4) At low temperatures, gas molecules have low average kinetic energies, but their intermolecular attractions are unaffected.

iii) <u>The van der Waals equation</u>

(1) Introduced two constants, a and b, to account for the excluded volume and attractive forces, respectively

(2) $\left(P + \dfrac{n^2 a}{V^2} \right)(V - nb) = nRT$ this formula is on the green sheet

(3) The values of a and b are different for each gas