Chemistry 2AP Assignments

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Text: Chemistry The Central Science, Eighth Edition; Brown, LeMay, Bursten
Student's Guide: Chemistry The Central Science, Eighth Edition; Brown, LeMay, Bursten

Complete these assignments as we cover the pertinent material in class.

Read Chapter 10 (take notes and review all sample exercises), pgs. 353-383.
Text Problems: Pg. 383-391: #'s 10.5, 10.7-10.9, 10.11, 10.14, 10.17, 10.20-10.23, 10.25-10.27, 10.31-10.33, 10.37, 10.39-10.42, 10.44, 10.46-10.48, 10.50-10.53, 10.55, 10.58, 10.59, 10.61-10.65, 10.67, 10.68, 10.70, 10.72, 10.76, 10.81, 10.84, 10.86, 10.88-10.91, 10.93-10.97, 10.99.
Check your answers to red textbook problems in back of book.
Student's Guide Chapter 10, read pgs. 182-193:

Exercises 1-15

Problems: pgs. 194-199: #'s 10.36-10.50.

Turn to page 220 (Cumulative Quiz 4) in the Student's Guide and tape a opaque flap (allow it to still flip up as if it were on a hinge) over the Answers so you cannot "sneak a peek" when you self-test for chapters 10-11.

Read Chapter 11 (take notes and review all sample exercises), pgs. 393-427.
Text Problems: Pg. 427-433: #'s 11.6, 11.7, 11.10, 11.12, 11.14-11.16, 11.18, 11.20, 11.21, 11.23-11.27, 11.29, 11.30, 11.32-11.35, 11.37, 11.39-11.41, 11.44, 11.45, 11.47, 11.48, 11.50, 11.51, 11.53, 11.58, 11.65, 11.68-11.71, 11.74-11.78, 11.81-11.84, 11.86, 11.88, 11.93, 11.95, 11.97, 11.100.
Student's Guide Chapter 11, read pgs. 200-210:

Exercises 1-13.

Problems: pgs. 210-216: #'s 11.45, 11.50-11.53, 11.56-11.70.

Read Chapter 12 (take notes and review all sample exercises), pgs. 435-462.
Text Problems: Pg. 462-467: #'s 12.8, 12.11, 12.12, 12.17, 12.19-12.22, 12.29, 12.40, 12.44, 12.45, 12.49, 12.54, 12.55.
Student's Guide Chapter 12, read pgs. 221-229 (a casual understanding is all that is necessary):

Exercises: 1-9.

Problems: pgs. 229-233: #'s 12.31-12.39, 12.42-12.43.

- Take Cumulative Quiz 4 in the Student's Guide: Pgs. 217-220, #'s 1-30. Check your answers on pg. 220. Note any difficulties you may have and be prepared to ask questions about them..

Chapter 10. Gases

10.1 Characteristics of Gases

- All substances have three phases: solid, liquid, and gas.
- Substances that are liquids or solids under ordinary conditions may also exist as gases.
  - These are often referred to as **vapors**.
- Many of the properties of gases differ from those of solids and liquids:
  - Gases are highly compressible and occupy the full volume of their containers.
  - When a gas is subjected to pressure, its volume decreases.
  - Gases always form homogeneous mixtures with other gases.
- Gases occupy only a small fraction of the volume of their containers.
  - As a result, each molecule of a gas behaves largely as though other molecules were absent.

10.2 Pressure

Atmospheric Pressure and the Barometer

- **Pressure** is the force acting on an object per unit area:

$$P = \frac{F}{A}$$

- The SI unit of pressure is the **pascal**.
  - 1 Pa = 1 N/m^2
  - 1 N = 1 $kg{\cdot}m/s^2$
- Gravity exerts a force on the Earth's atmosphere.
- A column of air 1 m^2 in cross section extending to the top of the atmosphere exerts a force of 10^5 N.
  - Thus, the pressure of a 1-m^2 column of air extending to the top of the atmosphere is 100 kPa.
- Atmospheric pressure is measured with a barometer.
- If a tube is completely filled with mercury and then inverted into a container of mercury open to the atmosphere, the mercury will rise until the pressure due to the mass of the mercury column is the same as atmospheric pressure.
  - Standard atmospheric pressure is the pressure required to support 760 mm of Hg in a column.
  - Units: 1 atm = 760 mm Hg = 760 torr = 1.01325 x 10^5 Pa = 101.325 kPa.
- Atmospheric pressure is sometimes reported using a related unit, the **bar**.
  - 1 bar = 10^5 Pa

Pressures of Enclosed Gases and Manometers

- In the laboratory the pressures of gases not open to the atmosphere are measured using manometers.
- A manometer consists of a bulb of gas attached to a U-tube containing Hg.
- If the U-tube is closed, then the pressure of the gas is the difference in height of the liquid (usually Hg).
- If the U-tube is open to the atmosphere, a correction term needs to be added:
  - If $P_{gas} < P_{atm}$, then $P_{gas} + P_h = P_{atm}$.
  - If $P_{gas} > P_{atm}$, then $P_{gas} = P_{atm} + P_h$.

10.3 The Gas Laws

- The equations that express the relationships among *T* (temperature), *P* (pressure), *V* (volume), and *n* (number of moles of gas) are known as *gas laws*.

The Pressure-Volume Relationship: Boyle's Law''

- Weather balloons are used as a practical application of the relationship between pressure and volume of a gas.
  - As the weather balloon ascends, the volume increases.
  - As the weather balloon gets farther from Earth's surface, the atmospheric pressure decreases.
- **Boyle's law**: The volume of a fixed quantity of gas, at constant temperature, is inversely proportional to its pressure.
- Mathematically:

$$V = \text{constant} \times \frac{1}{P} \text{ or } PV = \text{constant}$$

- A plot of V versus P is a hyperbola.
  - Similarly, a plot of V versus $1/P$ is a straight line passing through the origin.
- The process of breathing illustrates Boyle's law:
  - As we breathe in, the diaphragm moves down, and the ribs expand. Therefore, the volume of the lungs increases.
  - According to Boyle's law, when the volume of the lungs increases, the pressure decreases. Therefore, the pressure inside the lungs is less than atmospheric pressure.
  - Atmospheric pressure then forces air into the lungs until the pressure once again equals atmospheric pressure.
  - As we breathe out, the diaphragm moves up and the ribs contract. Therefore, the volume of the lungs decreases.
  - By Boyle's law, the pressure increases and air is forced out.

The Temperature-Volume Relationship: Charles's Law'

- We know that hot-air balloons expand when they are heated.
- **Charles's law:** The volume of a fixed quantity of gas at constant pressure is directly proportional to its absolute temperature.
- Mathematically:

$$V = \text{constant} \times T \text{ or } \frac{V}{T} = \text{constant}$$

- Note that the value of the constant depends on the pressure and number of moles of gas.
  - A plot of V versus T is a straight line.
  - When T is measured in °C, the intercept on the temperature axis is -273.15°C.
  - We define absolute zero, 0 K = -273.15°C.

The Quantity-Volume Relationship: Avogadro's Law' '

- Gay-Lussac's law of combining volumes: At a given temperature and pressure the volumes of gases that react with one another are ratios of small whole numbers.
- **Avogadro's hypothesis:** Equal volumes of gases at the same temperature and pressure contain the same number of molecules.
- **Avogadro's law**: The volume of gas at a given temperature and pressure is directly proportional to the number of moles of gas.
  - Mathematically:

$$V = \text{constant x } n$$

  - We can show that 22.4 L of any gas at 0°C and 1 atm contains 6.02 x 10^{23} gas molecules.

10.4 The Ideal-Gas Equation

- Summarizing the gas laws:
  - Boyle: $V \propto 1/P$ (constant n, T)
  - Charles: $V \propto T$ (constant n, P)
  - Avogadro: $V \propto n$ (constant P, T)
  - Combined: $V \propto nT/P$
- Ideal-gas equation: $PV = nRT$
  - Where R = **gas constant** = 0.08206 L·atm/mol·K
  - An **ideal gas** is a hypothetical gas whose P, V, and T behavior is completely described by the ideal-gas equation.
- Define **STP (standard temperature and pressure)** = 0°C, 273.15 K, 1 atm.
  - Volume of 1 mol of gas at STP is 22.4 L.

Relating the Ideal-Gas Equation and the Gas Laws

- If $PV = nRT$ and n and T are constant, then PV = constant, and we have Boyle's law.
  - Other laws can be generated similarly.
- In general, if we have a gas under two sets of conditions, then

$$\frac{P_1V_2}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$$

- We often have a situation in which P, V, and T all change for a fixed number of moles of gas.
  - For this set of circumstances,

$$\frac{PV}{T} = nR = \text{constant}$$

10.5 Further Applications of the Ideal-Gas Equation

Gas Densities and Molar Masses

- Density has units of mass over volume.
- Rearranging the ideal-gas equation with M as molar mass we get:

$$\frac{n}{V} = \frac{P}{RT}$$

$$\frac{nM}{V} = \frac{PM}{RT}$$

$$\therefore d = \frac{PM}{RT}$$

The molar mass of a gas can be determined as follows:

$$M = \frac{dRT}{P}$$

Volumes of Gases in Chemical Reactions' ' '

- The ideal-gas equation relates P, V, and T to the number of moles of gas (n).
- The n can then be used in stoichiometric calculations.

10.6 Gas Mixtures and Partial Pressures

- Since gas molecules are so far apart, we can assume that they behave independently.
- Dalton observed:
  - The total pressure of a mixture of gases equals the sum of the pressures that each would exert if present alone.
  - **Partial pressure** is the pressure exerted by a particular component of a gas mixture.
- **Dalton's law of partial pressures**: In a gas mixture the total pressure is given by the sum of partial pressures of each component:

$$P_t = P_1 + P_2 + P_3 + \ldots + P_n$$

- Each gas obeys the ideal-gas equation.
  - Thus,

$$P_t = (n_1 + n_2 + n_3 + \cdots)\frac{RT}{V} = n_t \frac{RT}{V}$$

Partial Pressures and Mole Fractions

- Let n_1 be the number of moles of gas 1 exerting a partial pressure P_1, then

$$P_1 = X_1 P_t.$$

  - Where X_1 is the **mole fraction** (n_1/n_t).

Collecting Gases over Water

- It is common to synthesize gases and collect them by displacing a volume of water.
- To calculate the amount of gas produced, we need to correct for the partial pressure of the water:

$$P_{total} = P_{gas} + P_{water}$$

- The vapor pressure of water varies with temperature.
  - Values can be found in Appendix B

10.7 Kinetic-Molecular Theory'

- The **kinetic-molecular theory** was developed to *explain* gas behavior.
  - Theory of moving molecules.
- Summary:
  - Gases consist of a large number of molecules in constant random motion.
  - The volume of individual molecules is negligible compared with the volume of the container.
  - Intermolecular forces (forces between gas molecules) are negligible.
  - Energy can be transferred between molecules during collisions, but the average kinetic energy is constant at constant temperature.
    - The collisions are perfectly elastic.
  - The average kinetic energy of the gas molecules is proportional to the absolute temperature.
- Kinetic-molecular theory gives us an *understanding* of pressure and temperature on the molecular level.
  - The pressure of a gas results from collisions of the molecules with the walls of container.
  - The magnitude of the pressure is determined by how often and how hard the molecules strike.
- The absolute temperature of a gas is a measure of the average kinetic energy.'
  - Some molecules will have less kinetic energy or more kinetic energy than the average (distribution).
  - There is a spread of individual energies of gas molecules in any sample of gas.
  - As the temperature increases, the average kinetic energy of the gas molecules increases.
- As kinetic energy increases, the velocity of the gas molecules increases.
  - **Root-mean-square (rms) speed**, u, is the speed of a gas molecule having average kinetic energy.
  - Average kinetic energy, ε, is related to rms speed:

$$\varepsilon = \tfrac{1}{2}mu^2$$

  Where m = mass of the molecule.

Application to the Gas Laws

- We can understand empirical observations of gas properties within the framework of the kinetic-molecular theory.
- Effect of an increase in volume (at constant temperature):
  - As volume increases at constant temperature, the average kinetic energy of the gas remains constant.
  - Therefore, u is constant.
  - However, volume increases, so the gas molecules have to travel farther to hit the walls of the container.
  - Therefore, pressure decreases.
- Effect of an increase in temperature (at constant volume):
  - If temperature increases at constant volume, the average kinetic energy of the gas molecules increases.
  - There are more collisions with the container walls.
  - The change in momentum in each collision increases (molecules strike harder).
  - Therefore, pressure increases.

10.8 Molecular Effusion and Diffusion

- The average kinetic energy of a gas is related to its mass:

$$\epsilon = \tfrac{1}{2}mu^2$$

- Consider two gases at the same temperature: the lighter gas has a higher rms speed than the heavier gas.
  - Mathematically:

$$u = \sqrt{\frac{3RT}{M}}$$

- The lower the molar mass, M, the higher the rms speed for that gas at a constant temperature.
- Two consequences of the dependence of molecular speeds on mass are:
  - **Effusion** is the escape of gas molecules through a tiny hole into an evacuated space.
  - **Diffusion** is the spread of one substance throughout a space or throughout a second substance.

Graham's Law of Effusion'
- The rate of effusion can be quantified.
- For two gases with molar masses M_1 and M_2. The relative rate of effusion is given by **Graham's law**:

$$\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$$

- Only those molecules that hit the small hole will escape through it.
  - Therefore, the higher the rms speed the more likely that a gas molecule will hit the hole.
  - We can show

$$\frac{r_1}{r_2} = \frac{u_1}{u_2} = \sqrt{\frac{M_2}{M_1}}$$

Diffusion and Mean Free Path
- Diffusion is faster for light gas molecules.
- Diffusion is significantly slower than the rms speed.
  - Diffusion is slowed by collisions of gas molecules with one another.
  - Consider someone opening a perfume bottle: It takes a while to detect the odor, but the average speed of the molecules at 25°C is about 515 m/s!
- The average distance traveled by a gas molecule between collisions is called the **mean free path**.
- At sea level, the mean free path for air molecules is about 6×10^{-6} cm.

10.9 Real Gases: Deviations from Ideal Behavior
- From the ideal gas equation:

$$\frac{PV}{RT} = n$$

- For 1 mol of an ideal gas, $PV/RT = 1$ for all pressures.
    - In a real gas, PV/RT varies from 1 significantly.
    - The higher the pressure the more the deviation from ideal behavior.
  - For 1 mol of an ideal gas, $PV/RT = 1$ for all temperatures.
    - In a real gas, PV/RT varies from 1 significantly.
    - As temperature increases, the gases behave more ideally.
- The assumptions in the kinetic-molecular theory show where ideal-gas behavior breaks down:
  - The molecules of a gas *have* finite volume.
  - Molecules of a gas *do* attract each other.
- As the pressure on a gas increases, the molecules are forced closer together.
  - As the molecules get closer together, the volume of the container gets smaller.
  - The smaller the container, the more of the total space the gas molecules occupy.
  - Therefore, the higher the pressure, the less the gas resembles an ideal gas.
  - As the gas molecules get closer together, the intermolecular distances decrease.
  - The smaller the distance between gas molecules, the more likely that attractive forces will develop between the molecules.
  - Therefore, the less the gas resembles an ideal gas.
- As temperature increases, the gas molecules move faster and farther apart.
  - Also, higher temperatures mean that more energy is available to break intermolecular forces.
  - As temperature increases, the negative departure from ideal-gas behavior disappears.

The van der Waals Equation

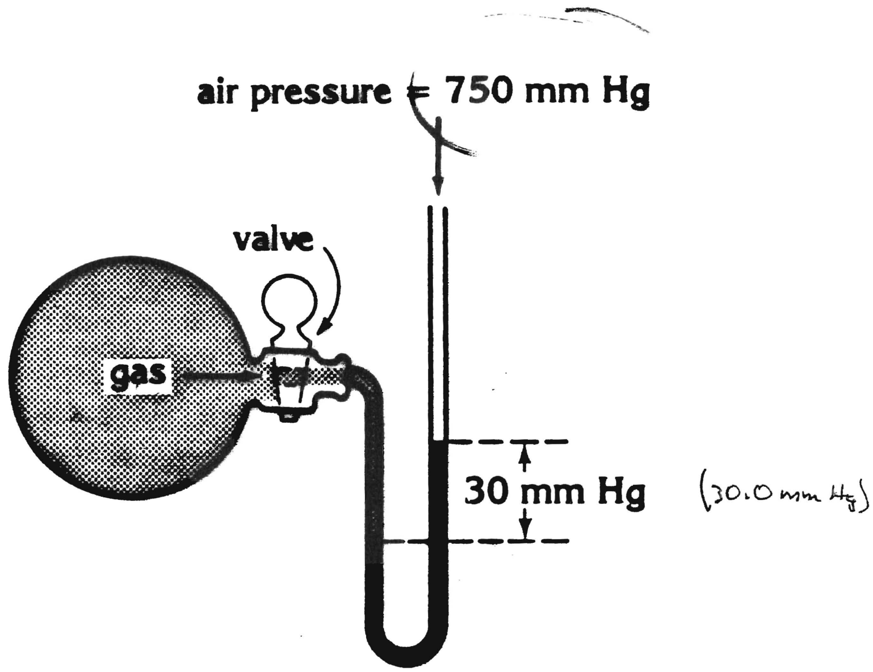
- We add two terms to the ideal-gas equation to correct for
  - The volume of molecules: $(V - nb)$
  - Molecular attractions:

$$\left(\frac{n^2 a}{V^2}\right)$$

$$\left(P + \frac{n^2 a}{V^2}\right)(V - nb) = nRT$$

  - The correction terms generate the **van der Waals equation**:
    - Where a and b are empirical constants.
- To understand the effect of intermolecular forces on pressure, consider a molecule that is about to strike the wall of the container.
  - The striking molecule is attracted by neighboring molecules.
  - Therefore, the impact on the wall is lessened.

U-Tube Manometer, Problem 4



Assuming that the valve is open, what pressure is the gas exerting? (in atmospheres)