Chapter 10 Exam AK

1. Which of the diatomic molecules are gases?
   Hydrogen, fluorine, nitrogen, oxygen, and chlorine.

2. What is the difference between a gas and a vapor?
   Gases are a phase that is highly energetic with low attractive forces at room temperature. Solids and liquids at room temperature can sometimes form a gaseous phase. This phase is called a vapor because not all of the matter present is in one phase.

3. A liquid and a gas are moved to large containers (let's keep the volume the same for both containers). How does their behavior differ once they are moved from a small container to a larger container?
   The gas will expand to fill the new container, assuming the shape and volume of the container. The liquid will assume the shape of the container which it occupies, but does not expand to fill the volume.

4. Although liquid water and carbon tetrachloride (CCl_4) do not mix, their vapors form a homogeneous mixture. Give an explanation based on the behavior of liquids and gases to explain this phenomenon.
   Liquid molecules are close together, while gas molecules have more space between them. The vapors of these two liquids mix because they are not interacting with attractive or repulsive forces. The space between the molecules is large enough to sustain differences in attractive forces.

5. **(4 points)** An ideal gas initially at 710 torr and 30.59°C occupies 2600 mL. Calculate the final temperature in °C, if the conditions are changed to a pressure of 1.20 atm and volume of 3.25 L

1.20 atm	**760 torr**	**3.25 L**	**30.59 K + 273.15 K**	**= 487.70K**
710 torr	**1 atm**	**2.6 L**		

487.70K – 273.15 K = 214.55K , 215 °C (2SF) 2.1×10^{2}°C

6. **(6 points)** A 23.5-mL volume of hydrochloric acid reacts completely with a solid sample of $MgCO_3$. The volume of CO_2 formed is 154 mL at 25.98°C and 731.6 mmHg. What is the molarity of the acid solution?

$$2HCl_{(aq)} + MgCO_{3(s)} \rightarrow CO_{2(g)} + H_2O_{(l)} + MgCl_{2(aq)}$$

Strategy: solve for moles of carbon dioxide, equate to moles HCl, find molarity.

	HCl	CO_2
Moles		
P		731.6mmHg, O.9626 atm
V	23.6mL, 0.0236L	154mL, 0.154 L
T		25.98°C, 299.13K

$$P = 731.6\ mmHgX \frac{1atm}{760\ mmHg} = 0.9626atm$$

$$V(gas) = 154\ mLx \frac{10^{-3}L}{1mL} = 0.154\ L$$

$$V(acid) = 23.5mLx \frac{10^{-3}L}{1mL} = 0.0235\ L$$

$25.98°Cx \frac{1K}{1°C} + 276.15K = 299.13K;\ PV = nRT$, solving for moles of gas,

$$\frac{PV}{RT} = \frac{0.96263\ atmx0.154L}{\frac{0.08206Latm}{Kmol} x\ 299.13K} = 0.0060393\ mol\ CO_2$$

$$0.0060393\ mol\ CO_2 x \frac{2HCl}{1CO_2} x \frac{1}{0.0235L} = 0.514\mathcal{M}\ HCl$$

7. **(6 points)** Cyanogen, a highly toxic gas, is composed of 46.2 g% C and 53.8% N by mass. At 25°C and 750 torr, 1.05 g of cyanogen occupies 0.500 L. What is the molecular formula of cyanogen.

46.2 g C	**1 mol C**	**=3.847 mol C**
	12.01 g C	
53.8 g N	**1 mol N**	**3.840 mol N**
	14.01 g N	

C	**3.847 mol C**	**1.00 C**
N	**1.071 mol N**	**1N**

CN, formula mass is 26.02 g/mol

750 torr	**1 atm**	**0.500 L**	**K•mol**	**= 0.02017 mol cyanogen**
	760 torr	**(25K+ 273.15 K)**	**0.08206 L•atm**	

1.05 g / 0.02017 mol cyanogen = 52.06 g/mol cyanogen. There are two formula units of CN in the cyanogen molecule. The formula is C_2N_2

8. **(5 points)** The rate of effusion of oxygen gas at 0°C is 4,61X10^{2}m/sec, what is the rate of SO_2 gas at the same pressure and temperature?

$$\frac{rate\ SO_2}{rate\ O_2} = \sqrt{\frac{MMO_2}{MMSO_2}};\ \frac{rate\ SO_2}{\frac{4.61x10^2m}{s}} = \sqrt{\frac{\frac{31.9988g}{mol}}{\frac{64.064g}{mol}}}\ 326\ \text{m/s}$$

9. **(4 points)** Suppose you were marooned on a tropical island and had to make a primitive barometer using sea water (density = 1.10 g/mL). What height would the water reach in your sea water barometer when a mercury barometer would reach 77.5 cm? d (Hg) = 13.6g/mL. **h_1d_1=h_2d_2**

$$\frac{\mathbf{13.6\ g\ Hg/mL\ x77.5\ cm\ high}}{\mathbf{1.10\ g\ sea\ water/mL}} 958\ cm\ high$$

10. **(4 points)** A gas cylinder with a volume of 6.00L contains 1.00 g of Ar and 2.00 g of Ne. The temperature of the two gases is 294K.

a. Find the partial pressure of each gas.

b. Find the mole fraction of each gas.

Moles of the gases

1.00 g Ar	1mol Ar	= 0.02503 mol Ar
	39.948 g Ar	

2.00 g Ne	1 mole Ne	= 0.09911 mol Ne
	20.18 g Ne	

P_{Ar}	N_{Ar}
P_{total}	N_{total}

P_{Ne}	n_{Ne}
P_{total}	N_{total}

Mole fraction of each gas

X_{Ar}	0.02503 mol Ar	0.202
	0.124 mol total	

X_{Ne}	0.09911 mol Ne	= 0.798
	0.124	

You could also solve for the total pressure of the gases, and use Dalton's Law to solve for the individual pressures.

$$P\ Ar = \frac{0.02503\ mol\ Ar \bullet\ 0.08206 \bullet\ 294\ K}{6.00L} = [0.1006\ atm]\ \boldsymbol{ANS:}\ \ \mathbf{0.101\ atm}$$

$$P\ Ne = \frac{0.09911\ mol\ Ne \bullet 0.08206 * 294\ K}{6.00\ L} = [0.3985\ atm]\ \boldsymbol{ANS}\mathbf{: 0.399\ atm}$$

0.1006 + 0.3985 = 0.4991 atm total pressure. **Ans:** 0.499 atm

11. **(8 points)** A sample of nitrogen gas is at STP. The volume of the container is decreased while keeping the temperature constant. Use kinetic-molecular theory to explain whether each of the following would increase, decrease, or remain constant and WHY.

a) the average KE

The temperature is constant so the average KE is also constant. Changing the temperature would change the average KE.

b) the average speed

The average speed is dependent on the temperature. If the temperature increases, the molecules have more KE, but since the temperature is constant, the KE is constant, and the speed is also constant.

c) the frequency of the collisions

The frequency of the collisions is inversely proportional to the volume. There will be more collisions with the walls of the container, because the path that the molecules travel is shorter to have a successful collision with the container walls even though the KE is constant.

d) the frequency of collisions per unit area

The frequency of the collisions per unit area must go if the frequency of the collisions increases.

e) The pressure of the gas

Pressure is a measure of the number of collisions/unit area and the force/unit area. If the force stays the same (KE does not change) but the number of collisions increases, the pressure must also increase. This is the basis of Boyle's Law.

12. **(4 points)** A bicycle tires filled with air to a pressure of 100. PSI at a temperature of 19°C. Riding the bike on a hot day increases the temperature of the tire to 58°C. The tire volume increases by 4.00%. What is the new pressure in the tire?

Let the initial volume = 100 L, then the final volume is 104 L

100. PSI	100.L	(58K + 273.15 K)	= 109 PSI
	(19K+273.15K)	104 L	

13. (**6 points)** Automobiles are equipped with airbags. Many that inflate with N_2 use the rapid reaction of NaN_3 + Fe_2O_3 which is initiated by a spark. How many grams of NaN_3 sodium azide) would be required to provide 75.0L of N_2 at 25.0°C and 748 mmHg?
$6NaN_{3(s)} + Fe_2O_{3(s)} \rightarrow 3Na_2O_{(s)} + 2Fe_{(s)} + 9N_{2(g)}$
moles of nitrogen gas

$$n = \frac{P \bullet V}{RT_K}$$

748 mmHg	**1 atm**	**75.0L**			**= 3.017 mol N_2**
	760 mmHg		$\mathbf{\frac{0.08206\ L \bullet atm}{K \bullet mol}}$	**298.15K**	

3.017 mol N_2	6 mol NaN_3	**65.028 g NaN_3**	**130.79 g NaN_3**	**131 g NaN_3**
	9 mol $N_{2(g)}$	**1 mol NaN_3**		

14. (6 points) Two flasks of equal volume are filled with a gas. Flask A contains H_2 at 0°C and 1 atm pressure. Flask B contains CO_2 at 25°C and 2 atm pressure. Compare these two gases, using the postulates of KMT, with respect to each of the following: [hint: start with a definition]

a. The average kinetic energy per molecule

The hydrogen gas has a lower average KE because the temperature of the gas is lower than that of the carbon dioxide. Average KE is ∝ to the temperature in kelvin.

b. The average molecular velocity.

Since the average KE is lower for the hydrogen, the average molecular velocity should be lower too. The velocity is ∝ **the temperature as well; colder gases have lower average KE and therefore velocity because there are more molecules moving at a slower speed than at a faster speed. However, velocity is also dependent on molar mass. To determine the gas with the lower (or higher), use Grahams Law of Effusion. Let hydrogen be gas 1, and carbon dioxide, gas 2**

$$\frac{rate_1}{rate_2} = \sqrt{\frac{T1}{MM1} \bullet \frac{MM2}{T2}} = \sqrt{\frac{273.15K}{2.016\ g/mol} \bullet \frac{44.01g/mol}{298.15K}} = \frac{4.47}{1}$$

because the Hydrogen Is lighter, it has a faster rate than the warmer carbon dioxide

problem cont'd on next page

c. The number of molecules

PV= nRT, We don't know the moles of gas and we don't know the volume. We do know that the volumes are equal. We can rearrange the ideal gas law to reflect this:

$$\frac{moles}{} = \frac{PV}{R\,T}$$

thusly, the volume is a unit volume, and we are comparing the pressure to the temperature.

$$\frac{moles}{} = \frac{1L \bullet 1\ atm}{R \bullet 273.15K} = 0.0446\ mols\ H_2$$

$$\frac{moles}{} = \frac{1L \bullet 2\ atm}{R \bullet 298.15K} = 0.00817\ mols\ CO_2$$

The two gases are at different pressures and temperatures for the same volume. For this to be true, the number of moles must be different. CO_2 has a larger number of moles. The size of the pressure is determined by the number of collisions with the walls of the container and the force of the collisions. Increasing the temperature causes more forceful collisions with the wall. The pressure increases.

d. The mass of the gas.

Heavier gases move slower than lighter gases. The hydrogen has a smaller molar mass, so even though it has a lower pressure, and a lower temperature, it will move faster than the heavier carbon dioxide molecules at a higher temperature and pressure.

15. (4 points) An incandescent light bulb is filled with 6.00 X 10^{-5} mol of argon. The bulb has a volume of 800.0 mL. What is the pressure of the argon in the light bulb at 75°C?

$$P\ Ar = \frac{6.00 \bullet 10^{-5}\ mol\ Ar \bullet\ 0.08206 \bullet\ (75 + 273.15)\ K}{0.8000\ L} = \boldsymbol{ANS:\ \ 0.00214\ atm}$$

16. (12 points) A quantity of Neon gas originally at 5.25 atm in a 2.00-L container at 26.0°C is transferred to a 12.5 L container at 20°C. A quantity of He originally at 5.25 atm and 26.0°C in a 5.00-L container is transferred to the same container (12.5L) containing the neon.

a. (3 points) What is the pressure of the neon in the new container?
b. (3 points)What new pressure of the He gas?
c. (2 points) What is the total pressure of the new container?
d. (4 points) What are the mole fractions of He and Ne in the new container?

P_{Ne}	5.25 atm	2.00L	293.15K	0.8232 atm, to 3 SF, 0.823 atm
		299.15K	12.5 L	

P_{He}	5.25 atm	5.00L	293.15K	2.058 atm, to 3 SF, 2.06 atm
		299.15K	12.5 L	

P_{total}= 2.058 +0.8232= 2.881atm. 2.88 atm

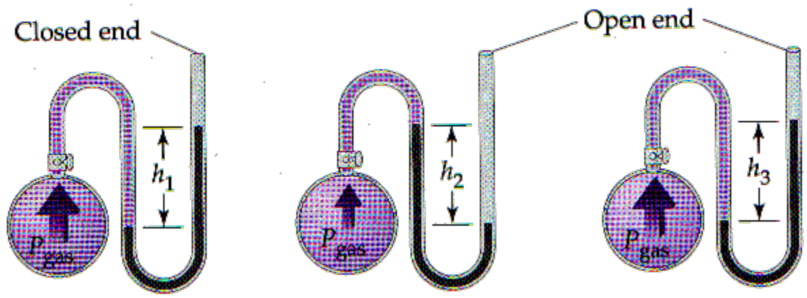
P_{Ar}	N_{Ar}		P_{Ne}	n_{Ne}
P_{total}	N_{total}		P_{total}	N_{total}

Mole fraction of each gas

X_{Ne}	0.8232 atm	0.2858, **0.286**	X_{He}	2.058 atm	= 0.7144, **0.714**
	2.881atm			2.881 atm	

17. (6 points) If the atmospheric pressure is 0.995 atm, what is the pressure of the enclosed gas in each of the two open ended manometers. Assume the grey material is mercury. [1]

H_1 52 cm H_2 = 67 cm H_3=10.3 cm



In a closed end manometer, the pressure exert is the height of the column of mercury, the P is 520 mmHg. The problem, though, is asking about the open end manometers.
In the second manometer, the atmospheric pressure is exerting a force on the gas. The pressure of the gas is
P_{gas}= P_{atm} —P_{height}, 0.995 atm —670 mmHg/760 mmHg = 0.113 atm, to the correct SF, 0.11 atm
In the third manometer, the gas is exerting a force on the atmosphere. The pressure of the gas is
P_{gas}= P_{atm}+P_{height}, 0.995 atm+103 mmHg/760 mmHg 1.13 atm

18. (4 points) What is the molar mass of a compound that takes 2.0 times longer to effuse through a porous plug than it did for the same amount of XeF_2 at the same temperature and pressure?
The rate of effusion for the XeF_2 = ½ the rate of the unknown compound, or the rate of the unknown is twice that of the xenon difluoride. This tells me that the unknown is heavier than the xenon difluoride.

$$\frac{\mu_{XeF_2}}{\mu_x} = \sqrt{\frac{MM_x}{MM_{XeF_2}}}$$

[1] http://www2.chemistry.msu.edu/courses/cem152/snl_cem152_SS12/_images/manometer.gif

$$\frac{\mu_{XeF_2}}{\mu_x} = \frac{2}{1} = \sqrt{\frac{MM_x}{169.3\ g/mol}}$$

molar mass is 677.2 g/mol

19. (4 points) An ideal gas initially at 1,209 mmHg and 30.00°C occupies 2,600. mL. Calculate the final temperature in **°C**, if the conditions are changed to a pressure of 1.50 atm and volume of 5.32 L

1.50 atm	**760 mmHg**	**5,320 mL**	**30.00 K + 273.15 K**	**= 582.87K**
1,209 mmHg	**1 atm**	**2,609 mL**		

582.87K – 273.15 K = 309°C (3SF)

20. (6 points) A self-contained breathing apparatus uses canisters containing potassium superoxide, KO_2. The superoxide consumes the CO_2 exhaled by a person and replaces it with oxygen. What mass of potassium superoxide is required to react with 8.00 L of carbon dioxide at 22.0°C and 767 mmHg?

$$4KO_{2(s)} + 2CO_{2(g)} \rightarrow 2K_2CO_{3(s)} + 3O_{2(g)}$$

$$n = \frac{P \bullet V}{RT_K}$$

767mmHg	**1 atm**	**8.00L**			**= 0.3333 mol** CO_2
	760 mmHg		$\frac{\mathbf{0.08206\ L \bullet atm}}{\mathbf{K \bullet mol}}$	**295.15K**	

0.3333 mol CO_2	$4KO_{2(s)}$	**71.1 g** $KO_{2(s)}$	**47.39 g** $\mathbf{KO_2}$	**47.4 g** $\mathbf{KO_2}$
	$2CO_2$	**1 mol** $KO_{2(s)}$		