CHEM 35

General Chemistry

EXAM #2

October 18, 2000

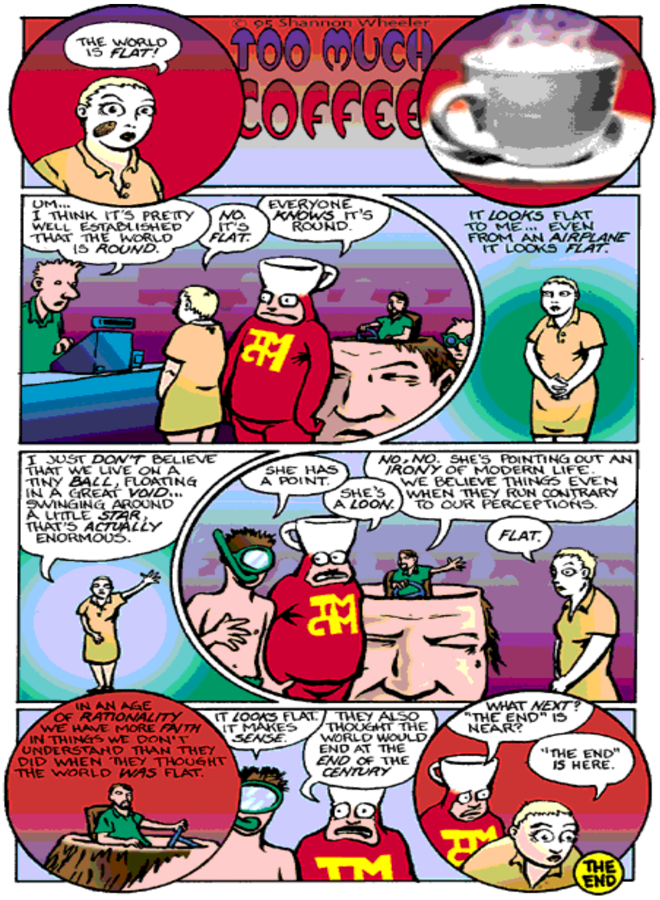
Name: Anne Serkey

SSN:_________________________

Lab T.A.:______________________

INSTRUCTIONS: Read through the entire exam before you begin. Answer all of the questions. For questions involving calculations, show **all** of your work -- **HOW** you arrived at a particular answer is **MORE** important than the answer itself! Circle your final answer to numerical questions.

The entire exam is worth a total of 150 points. Attached are a periodic table and a formula sheet jam-packed with useful stuff. Good Luck!



Page	Possible Points	*Points Earned*
2	20	20
3	30	30
4	15	15
5	20	20
6	15	15
7	20	20
8	30	30
TOTAL:	***150***	150

1. *(10 pts)* Pure acetic acid ($HC_2H_3O_2$; MW = 60.05 g/mol) is a liquid with a density of 1.049 g/mL at 25 ^oC. Calculate the molarity (mol/L) of a solution of acetic acid made by dissolving 15.00 mL of pure acetic acid at 25 ^oC in enough water to make 100.0 mL of solution.

$$\frac{15.00\text{ mL HAc}}{100.0\text{ mL Soln}} \text{ x } \frac{1.049\text{ g HAc}}{\text{mL HAc}} \text{ x } \frac{1\text{ mol HAc}}{60.05\text{ g HAc}} \text{ x } \frac{1000\text{ mL}}{\text{L}} = \begin{array}{l}2.620316\\ \text{mol HAc/L}\end{array}$$

$$= \underline{\mathit{2.630\ M\ HAc}}$$

2. *(10 pts)* How many mL of 0.155 *M* HCl are needed to react completely with 35.0 mL of a 0.101 *M* $Ba(OH)_2$ solution? The *unbalanced* reaction equation is:

$$HCl(aq) + Ba(OH)_2(aq) \rightarrow H_2O(l) + BaCl_2(aq)$$

Balanced: $2HCl(aq) + Ba(OH)_2(aq) \rightarrow 2H_2O(l) + BaCl_2(aq)$

$$35.0\text{ mL Ba(OH)}_2 \frac{\text{L}}{1000\text{ mL}} \text{ x } \frac{0.101\text{ mol Ba(OH)}_2}{\text{L}} \text{ x } \frac{2\text{ mol HCl}}{1\text{ mol Ba(OH)}_2} \text{ x } \frac{\text{L}}{0.155\text{ mol HCl}} \text{ x } \frac{1000\text{ mL}}{\text{L}} =$$

$$= 45.6129\text{ mL HCl}$$

$$= \underline{\mathit{45.6\ mL\ HCl}}$$

3. *(10 pts)* The *Hindenburg* was a famous hydrogen-filled dirigible that exploded in 1937. If the *Hindenburg* held 2.0 x 10^5 m^3 of hydrogen gas (H_2) at 25. ^oC and 1.0 atm, what mass (kg) of hydrogen was present?

$$V = 2.0 \times 10^5\ \text{m}^3 \times \frac{1000\ \text{L}}{1\ \text{m}^3} = 2.0 \times 10^8\ \text{L}$$

$$n = \frac{PV}{RT} = \frac{(1.0\ \text{atm})(2.0 \times 10^8\ \text{L})}{(0.08206\ \text{L-atm/mol-K})(298.15\ \text{K})} = 8.174547 \times 10^6\ \text{mol}\ H_2$$

$$8.1745 \times 10^6\ \text{mol}\ H_2 \times \frac{2.01588\ \text{g}\ H_2}{\text{mol}\ H_2} \times \frac{1\ \text{kg}}{1000\ \text{g}} = 1.64789 \times 10^4\ \text{kg}$$

$$= \underline{1.6 \times 10^4\ \text{kg}\ H_2}$$

4. *(10 pts)* Calculate the molar mass (g/mol) of a gas if 3.75 g occupies 0.935 L at 430. torr at 35.0 ^oC.

$$430\ \text{torr} \times \frac{1\ \text{atm}}{760\ \text{torr}} = 0.56579\ \text{atm}$$

$$n = \frac{PV}{RT} = \frac{(0.56579\ \text{atm})(0.935\ \text{L})}{(0.08206\ \text{L-atm/mol-K})(308.15\ \text{K})} = 2.09205 \times 10^{-2}\ \text{mol}$$

$$\text{Molar mass} = \frac{3.75\ \text{grams}}{2.09205 \times 10^{-2}\ \text{mol}} = 179.25\ \text{g/mol} = \underline{179.\ \text{g/mol}}$$

5. A mixture containing 0.538 mol He(*g*), 0.315 mol Ne(*g*), and 0.103 mol Ar(*g*) is confined in a 5.00-L vessel at 25.0 ^oC.

   a. *(5 pts)* Calculate the total pressure (atm) of the mixture.

$$N_{total} = 0.538 + 0.315 + 0.103 = 0.956\ \text{mol}$$

$$P_{total} = \frac{nRT}{V} = \frac{(0.956\ \text{mol})(0.08206\ \text{L-atm/mol-K})(298.15\ \text{K})}{5.00\ \text{L}}$$

$$= 4.6779\ \text{atm} = \underline{4.68\ \text{atm}}$$

   b. *(5 pts)* Calculate the partial pressure (atm) of He in the mixture.

$$P_{He} = X_{He}P_{total} \qquad X_{He} = n_{He}/n_{total} = \frac{0.538\ \text{mol}}{0.956\ \text{mol}} = 0.562762$$

$$P_{He} = (0.562762)(4.6779\ \text{atm})$$
$$= 2.63256\ \text{atm} = \underline{2.63\ \text{atm}}$$

6. *(5 pts)* A balloon made of rubber that is permeable to small molecules is filled with helium to a pressure of 1 atm. This balloon is then placed in a box that contains pure hydrogen (H_2), also at a pressure of 1 atm. Will the balloon expand or contract? Explain.

The balloon with *EXPAND*. The rate of effusion of the H_2 is greater than that of He, due its lower molar mass (recall: $r_1/r_2 = [(M_2/M_1)]^{½}$), so H_2 will enter the balloon faster than He will leave the balloon and it will expand.

7. Vessel A contains CO gas at 0 ^{o}C and 1 atm. Vessel B contains CO_2 gas at 20 ^{o}C and 0.5 atm. The two vessels have the same volume.

   a. *(3 pts)* Which vessel contains more molecules? Briefly explain.

VESSEL A – Since the absolute temperatures are almost the same (273 versus 293 K), we can say that the number of moles of gas is proportional to the pressure. Since $P_A = 2 \times P_B$, there are more moles of gas in vessel A.

   b. *(3 pts)* In which vessel is the average kinetic energy of the molecules greater? Briefly explain.

VESSEL B – Kinetic energy is proportional to *temperature*, so the gas at the greater temperature will have the greater average kinetic energy. $T_B > T_A$, so the molecules in vessel B have a greater K.E.

   c. *(4 pts)* In which vessel is the rms speed of the molecules greater? Briefly explain.

VESSEL A – We need to take into account both the *temperature* and the *molar mass*, as v_{rms} is proportional to $(T/M)^{½}$. The molar mass difference is greater than the temp difference, so the vessel with the lighter gas molecules (vessel A) will contain the faster molecules (even though they are at a slightly lower temp).

8. The ideal gas law is often used under conditions in which gases are unlikely to behave in an ideal fashion.

   a. *(5 pts)* Calculate, using the ideal gas law, the pressure (atm) of 40.0 mol of argon gas contained in a volume of 2.00 L at 300 K.

$$\mathbf{P = \frac{nRT}{V} = \frac{(40.0\ mol)(0.08206\ L\text{-}atm/mol\text{-}K)(300\ K)}{2.00\ L} = 492.36\ atm}$$

$$\mathbf{= \underline{\mathit{492.\ atm}}}$$

   b. *(8 pts)* Now, using the van der Waals equation, calculate the pressure (atm) of the same argon gas sample.

$$\mathbf{P = \frac{nRT}{V\text{-}nb} - \frac{an^2}{V^2}}$$

$$\mathbf{= \frac{(40.0\ mol)(0.08206\ L\text{-}atm/mol\text{-}K)(300\ K)}{(2.00\ L - (40.0\ mol)(0.0320\ L/mol)} - \frac{(1.36\ atm\text{-}L^2/mol^2)(40.0\ mol)^2}{(2.00\ L)^2}}$$

$$\mathbf{= 1367.67\ atm - 544.00\ atm = 823.67\ atm = \underline{\mathit{824.\ atm}}}$$

   c. *(7 pts)* If there is a greater than 10% difference in the two pressure values, explain why the ideal gas law fails. If there is less than a 10% difference in the two pressure values, indicate under what conditions you would expect to find a more significant difference between the two calculated pressures.

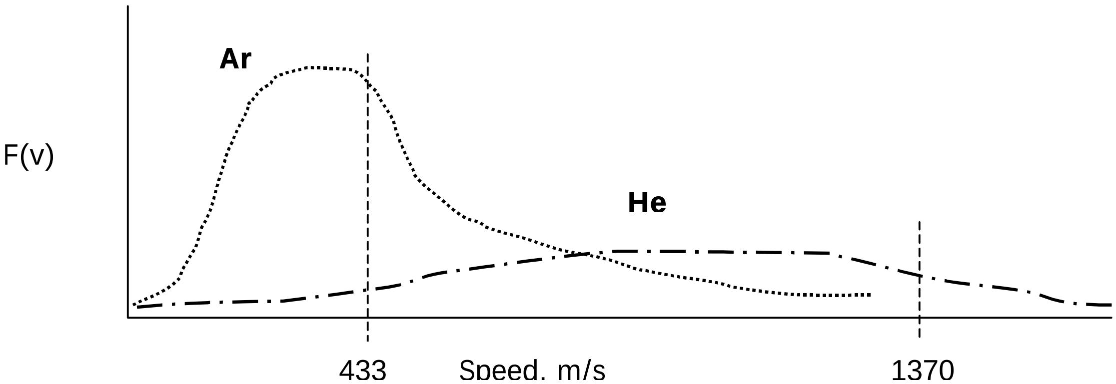
The pressure difference between the IGL and the VDW equation is much more than 10% under these conditions (it’s > 40%). The IGL fails under conditions of *high pressure* and *low temperature* and we have very high pressure here. The VDW equation corrects for the IGL’s ignorance of both the volume occupied by the gas molecules themselves as well as their intermolecular attraction. The former would tend to *elevate* the pressure by increasing intermolecular repulsion, while the latter would tend to *decrease* the pressure by decreasing the volume occupied by the gas species clustering closer together (due to attractive forces). Clearly, the *repulsive* forces dominate here as the VDW equation pressure is nearly *double* that predicted by the IGL.

9. Consider two gas samples: He and Ar at 300 K.

   a. *(10 pts)* Sketch the Maxwell-Boltzmann speed distribution profile for both gases; use the same set of axes for both gases. Calculate the rms speed for both gases and indicate the location of these values on the plot. Please note: the ONLY quantitative aspect of this plot should be these rms speeds – the rest should be strictly a qualitative sketch showing clearly any differences between the two gases.

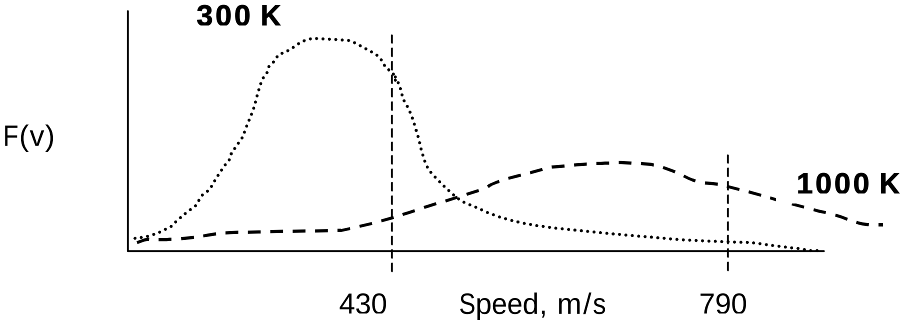
$$v_{rm}\ (\mathrm{Ar}) = (3RT/M)^{1/2} = [3(8.3145\ \mathrm{J/mol})(300\mathrm{K})/(39.948 \times 10^{-3}\ \mathrm{kg/mol})]^{1/2}$$
$$v_{rm}\ (\mathrm{Ar}) = \underline{433\ \mathrm{m/s}}$$

$$v_{rm}\ (\mathrm{He}) = (3RT/M)^{1/2} = [3(8.3145\ \mathrm{J/mol})(300\mathrm{K})/(4.0026 \times 10^{-3}\ \mathrm{kg/mol})]^{1/2}$$
$$v_{rm}\ (\mathrm{He}) = \underline{1370\ \mathrm{m/s}}$$



   b. *(5 pts)* How would the speed distribution for Ar be affected by an increase in temperature to 1000 K? Sketch speed distribution profiles for Ar (on a new set of axes, please) at the initial temperature (300 K) and at 1000 K. As before, use the calculated rms speed as the sole quantitative calibration point for these qualitative speed distribution plots.

$$v_{rm}\ (\mathrm{Ar}) = (3RT/M)^{1/2} = [3(8.3145\ \mathrm{J/mol})(1000\mathrm{K})/(39.948 \times 10^{-3}\ \mathrm{kg/mol})]^{1/2}$$
$$v_{rm}\ (\mathrm{Ar}) = \underline{790\ \mathrm{m/s}}$$



10. Each of the following reactions produces a gas as one of its products. For each reaction: write a balanced reaction equation, classify the reaction as being a *precipitation*, *acid/base*, or *oxidation/reduction* reaction, and give evidence to support your classification by either identifying the precipitate compound, the acid and the base, or by indicating which of the reactants are oxidized and which are reduced.

    a. $\mathbf{CaCO_3(\mathit{s}) + HNO_3(\mathit{aq}) \rightarrow}$ **products**

    Balanced reaction equation *(3 pts)*:

    $$CaCO_3(s) + 2\ HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + H_2O(l) + CO_2(g)$$

    Reaction type *(1 pt)*: **precipitation** or [✓ **acid/base**] or **redox**
    (circle one)

    Evidence *(1 pt)*:
    Acid: HNO_3 Base: $CaCO_3$

    b. $\mathbf{Cu(\mathit{s}) + HNO_3(\mathit{aq}) \rightarrow}$ **products**

    Balanced reaction equation *(3 pts)*:

    $$Cu(s) + 2\ HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + H_2(g)$$

    Reaction type *(1 pt)*: **precipitation** or **acid/base** or [✓ **redox**]
    (circle one)

    Evidence *(1 pt)*:
    Oxidized: Cu Reduced: H^+

11. *(2 pts each)* Determine the oxidation state of sulfur in each of the following compounds:

    a. H_2S ***-2***

    b. S_8 ***0***

    c. SCl_2 ***+2***

    d. Na_2SO_3 ***+4***

    e. SO_4^{2-} ***+6***

12. A sample of 70.5 mg of K_3PO_4 (MW = 212.266 g/mol) is added to 15.0 mL of 0.050 *M* $AgNO_3$, resulting in the formation of a precipitate.

    a. *(5 pts)* Write a balanced net ionic equation for this reaction.

    $3\ \text{Ag}^+(aq) + \text{K}_3\text{PO}_4(s) \rightarrow \text{Ag}_3\text{PO}_4(s) + 3\ \text{K}^+(aq)$

    b. *(5 pts)* What is the limiting reactant in this reaction.

$$0.0705\text{ g K}_3\text{PO}_4 \times \frac{\text{mol K}_3\text{PO}_4}{212.266\text{ g K}_3\text{PO}_4} \times \frac{3\text{ mol AgNO}_3}{1\text{ mol K}_3\text{PO}_4} \times \frac{\text{L}}{0.050\text{ mol AgNO}_3} \times \frac{1000\text{ mL}}{\text{L}} =$$

19.9 mL $AgNO_3$ (needed)

Only 15 mL $AgNO_3$ available, so
$AgNO_3$ is LIMITING REACTANT

    c. *(5 pts)* Calculate the mass (g) of precipitate formed in this reaction

$$0.0150\text{ L AgNO}_3 \times \frac{0.050\text{ mol AgNO}_3}{\text{L}} \times \frac{1\text{ mol Ag}_3\text{PO}_4}{3\text{ mol AgNO}_3} \times \frac{418.576\text{ g Ag}_3\text{PO}_4}{\text{mol Ag}_3\text{PO}_4} =$$

= 0.10464 g Ag_3PO_4
= *0.10 g Ag_3PO_4*

13. *(10 pts)* Determine the Lewis structures (yes, there is more than one) for SO_2, and compute oxidation numbers and formal charges for the sulfur and oxygen atoms.

$$16\ e^- \rightarrow \quad |\underline{O} = \underline{S} - \overline{\underline{O}}| \quad \leftrightarrow \quad |\overline{\underline{O}} - \underline{S} = \underline{O}|$$

Oxidation Numbers: ***O ® -2*** ***S ® +4***

Formal Charges: fc = #valence electrons – (#bonds + # nonbonded e^-)

S = 6 – (3 + 2) = +1
O = 6 – (2 + 4) = 0 (double-bonded oxygen)
O = 6 – (1 + 6) = -1 (single-bonded oxygen)

b.*(5 pts)* Do the oxidation numbers or the formal charges more realistically indicate the charge distribution within the molecule? Explain.

Oxidation numbers assume bonds are totally ionic, while formal charges assume bonds are mostly covalent. In SO_2, bonding is between two non-metals and, so, is mostly covalent. Thus, the formal charges are better indicators of the charge distribution.

Betcha thought that I forgot all about the

EXTRA CREDIT! – *10 pts*

A newspaper article about the danger of global warming from the accumulation of greenhouse gases such as carbon dioxide states that "reducing driving your car by 20 miles a week would prevent release of over 1000 pounds of CO_2 per year into the atmosphere." Is this a reasonable statement? Assume that gasoline is octane (C_8H_{18} - MW = 114.1502 g/mol) and that it is burned completely to CO_2 and H_2O in the engine of your car. Note that you will need to make some reasonable guesses as to the gas mileage of your car, the density of octane, etc.

$$2\ C_8H_{18} + 25\ O_2 \rightarrow 16\ CO_2 + 18\ H_2O$$

Assuming: 20 miles/gallon
4 quarts/gallon
1 L = 1 quart
density of gas = density of water = 1 g/mL

$$\frac{20\text{ miles}}{\text{week}} \times \frac{\text{gallon}}{20\text{ miles}} \times \frac{52\text{ weeks}}{\text{year}} = \mathit{52\ gallons/year}$$

$$52\text{ gallons} \times \frac{4\text{ quarts}}{\text{gallon}} \times \frac{1\text{ L}}{1\text{ qt}} \times \frac{1000\text{ mL}}{\text{L}} \times \frac{1\text{ g } C_8H_8}{\text{mL}} =$$

$$= \mathit{2.08 \times 10^5\ g\ C_8H_{18}}$$

$$2.08 \times 10^5\text{ g } C_8H_{18} \times \frac{1\text{ mol } C_8H_{18}}{114.15\text{ g } C_8H_{18}} \times \frac{16\text{ mol } CO_2}{2\text{ mol } C_8H_{18}} \times \frac{44\text{ g } CO_2}{\text{mol } CO_2} =$$

$$= \mathit{6.414 \times 10^5\ g\ CO_2}$$

$$6.414 \times 10^5\text{ g } CO_2 \times \frac{\text{kg}}{1000\text{ g}} \times \frac{2.2046\text{ lb}}{\text{kg}} =$$

$$= 1.414 \times 10^3\text{ lb } CO_2$$

$$= \underline{\mathit{1,400\ lb\ CO_2/year}}$$

Yes! The statement is Reasonable.