Dr. Zellmer
Time: 7 PM Sun.
40 min

Chemistry 1210
Autumn Semester 2022
Quiz XI

All Sections
November 20, 2022

Name ______________KEY________________ Rec. Instr. ___________________________________

1. (4 pts) Mercury has a density of 13.6 g/mL. The pressure measured using a mercury barometer is 745.0 mm Hg. How high would a column of a new high tech oil (density 5.50 g/mL) be at this same pressure **AND** would this be a reasonably useful barometer? (**EXPLAIN**!). (1 in = 2.54 cm, 1 ft = 0.3048 m)

$P = g \bullet d \bullet h$ (See homework problem 10.16 solutions, 10.14 in 10th ed.)

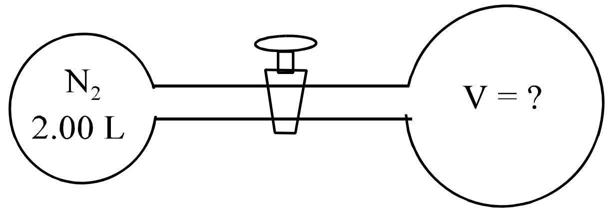
Since the P has to be the same no matter what liquid you have you get,

$d_2 \bullet h_2 = d_1 \bullet h_1$ $(P_2 = g \bullet d_2 \bullet h_2,\ P_1 = g \bullet d_1 \bullet h_1,\ P_2 = P_1)$

$$\therefore \quad h_2 = \left(\frac{d_1}{d_2}\right) h_1 = \left(\frac{13.6 \text{ g/mL}}{5.50 \text{ g/mL}}\right) * 745.0 \text{ mm} = 18\underline{4}2.1 \text{ mm oil} = 1.8\underline{4} \text{ m } (6.0\underline{4}3 \text{ ft})$$

So a barometer made using this oil would have to be just over 6 ft high to measure this pressure. In order to measure pressures we normally encounter, around 760.0 mm Hg, the barometer would have to be over 6.2 ft high. This would **NOT** be a very **practical** (useful) barometer. You need to use a liquid with a fairly large density to make a useful barometer or make a fancier barometer.

2. (5 pts) Two flasks are connected by a stopcock. Both flasks are held at the same temperature. The 2.00 L flask is filled with N_2 at a pressure of 1456 mm Hg. The flask with an unknown volume, V, was evacuated (contains no gas). The stopcock is opened and the N_2 fills both flasks. The resulting pressure after the N_2 fills both flasks is 416 mm Hg? What is the volume, V, of the flask on the right (in liters, L)?



There's a flask with unknown volume, V, connected to the 2.0 L flask with N_2 at a pressure of 1456 mm Hg. The stopcock is opened and the N_2 expands to fill both flasks. The final pressure is 416 mm Hg. What is the volume of the flask on the right? Use Boyle's Law to calculate the TOTAL volume of both flasks and then the volume of the flask on the right.

P &V change at constant T - Use **Boyle's** Law

$$P_2V_2 = P_1V_1 \quad \text{or} \quad P_fV_f = P_iV_i$$

P_1 = 1456 mm Hg $\qquad$ P_2 = 416 mm Hg

V_1 = 2.00 L $\qquad$ V_2 = ? L

$$V_2 = \frac{P_1V_1}{P_2} = \left(\frac{P_1}{P_2}\right) V_1 = \left(\frac{1456 \text{ mm Hg}}{416 \text{ mm Hg}}\right) 2.00 \text{ L} = 7.00 \text{ L}$$

Check: If P dec, V inc and it did. Hopefully this makes sense, the gas fills the whole thing.

Determine the volume of the container,

$V_2 = V_1 + V$ $\qquad$ $V = V_2 - V_1$ = 7.00 - 2.00 = **5.00 L**

3. (5 pts) What **volume** will 1.60 g of O_2 occupy at STP? (atomic weights: O = 16.00)

There are several methods for solving this problem. In all of them you need to know what STP is.

STP => Standard Temp. & Pressure: T = 0°C (273.15K) & P = 1 atm

a) Method 1

To get the volume determine the moles of $\mathbf{O_2}$ and then use IGL (PV=nRT) to get the volume.

$$? \text{ mol } O_2 = 1.60 \text{ g } O_2 \times \frac{1 \text{ mol } O_2}{32.00 \text{ g } O_2} = 0.050\underline{\mathbf{0}} \text{ mol } O_2$$

$$V = \frac{nRT}{P} = \frac{(0.050\underline{\mathbf{0}} \text{ mol})(0.08206 \text{ L•atm/mol•K})(273.15 \text{ K})}{(1 \text{ atm})}$$

$$= 1.1\underline{\mathbf{2}}073 \text{ L } O_2$$

$$= 1.12 \text{ L } O_2$$

b) Method 2

$$\mathcal{M} = \left(\frac{RT}{P}\right) D \quad \Rightarrow \quad D = \left(\frac{P}{RT}\right) \mathcal{M} \quad \text{(in g/L)}$$

$$D = \frac{(1 \text{ atm})(32.00 \text{ g/mol})}{(0.08206 \text{ L•atm/mol•K})(273.15\text{K})} = 1.42\underline{\mathbf{7}}6 \text{ g/L}$$

$$V = \frac{m}{D} = \frac{1.60 \text{ g}}{1.42\underline{7}6 \text{ g/L}} = 1.12 \text{ L}$$

Using dimensional analysis:

$$? \text{ L} = 1.60 \text{ g x } \frac{1\text{L}}{1.42\underline{7}6 \text{ g}} = 1.12 \text{ L}$$

c) Method 3 (only usable at STP)

Molar Volume at STP: V/n = 22.41 L/mol (1 mol gas = 22.41 L)

$$? \text{ L} = 0.050\underline{\mathbf{0}} \text{ mol} \times \frac{22.41 \text{ L}}{1 \text{ mol}} = 1.12 \text{ L}$$

4. (5 pts) Consider a cylinder fitted with a movable piston that can expand against the atmosphere. The initial pressure, volume and absolute temperature inside the cylinder are P_i, V_i and T_i. What is the new **temperature** of the system when the **pressure** is **tripled** and the **volume** is **decreased** to **one half** of the **original** volume?

Use **Combined Gas Law**:

$$\frac{P_f V_f}{T_f} = \frac{P_i V_i}{T_i}$$

$P_f = 3P_i$ $\quad$ $V_f = 1/2\ V_i$ (Vol dec to 1/2 its initial value)

$$T_f = \frac{P_f V_f}{P_i V_i} T_i = \left(\frac{P_f}{P_i}\right)\left(\frac{V_f}{V_i}\right) T_i$$

$$T_f = \left(\frac{3\ P_i}{P_i}\right)\left(\frac{1/2\ V_i}{V_i}\right) T_i = (3/2)T_i$$

5. (4 pts) Consider three one-liter flasks labeled A, B, and C filled with the gases NO, NO_2, and N_2O, respectively, each at STP. What can be said about the number of molecules of each gas? (atomic weights: N = 14.01, O = 16.00)

We can calculate the moles for each of the gases given using the P, V and T data given, (STP means 1 atm and 0 °C (273.15 K)) and a volume of 1 L:

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

However, you do NOT have to do this calculation. Since all have the same P, V and T they have the **same number of moles** (**same # of molecules**).

6. (5 pts) A cylinder of containing H_2 gas with a volume of 18.0 L and a pressure of 35.0 atm at a temperature of 23.0 °C is used to fill smaller cylinders with a volume of 2.00 L. How many smaller cylinders can be filled at a temperature of 25.0 °C to a volume of 2.00 L and a pressure of 745.0 torr (assuming all the gas can be transferred to the smaller cylinders).
(1 atm = 760 mmHg = 760 torr = 101.325 kPa = 14.7 lb/in^2)

Since P, V & T all change use the "Super" Combined Gas Law (with n, moles, constant):

$$\frac{P_2V_2}{n_2T_2} = \frac{P_1V_1}{n_1T_1} \qquad\qquad \frac{P_2V_2}{T_2} = \frac{P_1V_1}{T_1}$$

P_1 = 35.0 atm $\qquad$ V_1 = 18.00 L $\qquad$ T_1 = 23°C + 273.15 = 29**6**.15 K

$P_2 = 745.0 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 0.980\underline{\mathbf{2}}6$ atm $\qquad$ V_2 = ? $\qquad$ T_2 = 25°C + 273.15 = 29**8**.15 K

$$V_2 = \frac{P_1V_1T_2}{P_2T_1} = \left(\frac{P_1}{P_2}\right)\left(\frac{T_2}{T_1}\right)V_1 = \left(\frac{35.0 \text{ atm}}{0.980\underline{\mathbf{2}} \text{ atm}}\right)\left(\frac{29\underline{\mathbf{8}}.15 \text{ K}}{29\underline{\mathbf{6}}.15 \text{ K}}\right) 18.00 \text{ L} = 64\underline{7}.92 \text{ L}$$

This is the total volume (64<u>7</u>.92 L) that could be filled at a pressure of 745.0 torr.

cylinders = 64<u>7</u>.92 L x (1 cylinder/2.00 L) = 32**<u>3</u>**.51 cylinders = **324 cylinders** (3 s.f.)

Could also determine the moles of gas present in the larger cylinder and how many moles could be held in each smaller cylinder and then divide the total you started with by the number of moles which could be put in a smaller cylinder to get the number of small cylinders needed.

7. (4 pts) Five identical 1.0-L flasks contain the following gases each at 0°C and 1 atm pressure. The densities (g/L) of the gases are listed below. Which gas has the **LARGEST molar mass**?

Gas 1	Gas 2	Gas 3	Gas 4	Gas 5
0.178	3.16	0.715	0.760	1.25

The **density** is **directly proportional** to the **molar mass** (molecular or atomic weight). Therefore, the gas with the largest density should have the largest molar mass (MW or AW) (since P, V & T same for all). The densities are listed below, as are the MW (which you didn't need to calculate).

$$\mathcal{M} = \left(\frac{RT}{P}\right) D \qquad => \qquad D = \left(\frac{P}{RT}\right)\mathcal{M} \qquad \text{(in g/L)}$$

0.178	**3.16**	0.715	0.760	1.25
3.99	**70.8**	16.0	17.0	28.0

8. (5 pts) The empirical formula of a volatile liquid is C_2H_4O. A 0.345-gram sample of its vapor occupied 85.0 mL at 100.0°C and 0.942 atm. What is the **molecular formula** for the compound? (Atomic weights: H = 1.008, C = 12.01, O = 16.00)

Determine molecular formula of a gas given the empirical formula, C_2H_4O, mass & P, V, T data

MF = $(EF)_{\boldsymbol{n}}$ & MW = $\boldsymbol{n}$ * (EFW) (where MW = molecular wt and EFW = empirical formula wt)

$$\boldsymbol{n} = \frac{MW}{EFW} \quad \& \quad \mathcal{M} = \frac{\text{mass}}{\text{moles}} \quad \text{(molar mass or molecular wt)}$$

Have mass of 0.345 g, need moles. (Remember, molar mass = g/mol, same as MW.)

P = 0.942 atm; V = 85.0 mL = 0.0850 L; T = 100°C + 273.15 = 37**3**.15 K

$$n_{CxHyOz} = \frac{PV}{RT} = \frac{(0.942 \text{ atm})(0.0850 \text{ L})}{(0.08206 \text{ L•atm/mol•K})(37\underline{\mathbf{3}}.15 \text{ K})} = 2.6\underline{\mathbf{1}}48 \times 10^{-3} \text{ mol}$$

$$\mathcal{M} = \frac{0.345 \text{ g}}{2.6\underline{\mathbf{1}}48 \times 10^{-3} \text{ mol}} = 13\underline{\mathbf{1}}.93 \text{ g/mol or } 13\underline{\mathbf{1}}.93 \text{ amu}$$

EFW_{C2H4O} = 44.05 amu

$$\boldsymbol{n} = \frac{13\underline{\mathbf{1}}.93 \text{ amu}}{44.05 \text{ amu}} = 2.9951 = 3$$

MF = $(C_2H_4O)_3 = C_6H_{12}O_3$

9. (5 pts) What volume (L) of NO at 500°C and 0.5 atm will be produced in the following reaction if 10.0 L of oxygen reacts with excess NH_3 and the volume of NO is measured under the same conditions of temperature and pressure? (atomic weights: N = 14.01, H = 1.008, O = 16.00)

$$4\ NH_3(g) + 5\ O_2(g) \longrightarrow 4\ NO(g) + 6\ H_2O(g)$$

Can calculate the volume of desired from volume of given by using the volume ratio which is the same as the mole ratio from the balanced equation. You can see this from **Avogadro's Law**:

$$\frac{V_2}{n_2} = \frac{V_1}{n_1} \quad \text{or} \quad \frac{n_2}{n_1} = \frac{V_2}{V_1} \qquad \text{vol ratio = mole ratio}$$

$$? \text{ L NO} = 10.0 \text{ L } O_2 \times \frac{4 \text{ L NO}}{5 \text{ L } O_2} = 8.00 \text{ L}$$

numbers in this ratio are
coefficients in bal. eqn.

Could use the P, T, V data given to first calculate the moles of O_2 that reacts. Then do a mole-to-mole stoichiometry problem to determine the moles of NO produced from the moles of O_2. Then you could use the moles of NO with the P and T data given to calculate the volume of NO. This is not complicated but certainly more complicated and much more work than above.