Dr. Zellmer
Time: 7 PM Sun.
40 min

Chemistry 1210
Autumn Semester 2022
Quiz XII

All Sections
December 4, 2022

Name ______KEY______ Rec. Instr. ____________________

1. (5 pts) A mixture of gases contains 0.25 mol He and 0.65 mol N_2. If the total pressure of the mixture is 5.50 atm at 25°C, what are the **partial** **pressures** of He and N_2 at 25 °C? (At. Wts: He = 4.00, H = 14.01)

Method 1 (easier): Use $P_i = \chi_i P_T$ & $\chi_i = n_i/n_{Total}$

$$\chi_{He} = \frac{0.25 \text{ mol He}}{(0.25 + 0.65) \text{ mol}} = 0.2\underline{7}7 \qquad \chi_{N2} = \frac{0.65 \text{ mol } N_2}{(0.90 \text{ mol})} = 0.7\underline{2}2$$

$$P_{He} = \chi_{He} P_T = (0.2\underline{7}7)(5.50 \text{ atm}) = 1.\underline{5}2 \text{ atm} = 1.5 \text{ atm}$$

$$P_{N2} = \chi_{N2} P_T = (0.7\underline{2}2)(5.50 \text{ atm}) = 3.\underline{9}7 \text{ atm} = 4.0 \text{ atm}$$

Method 2 (harder, longer): use IGL & find vol using P_{Tot} & n_{Tot}

$$V = \frac{n_{Tot}RT}{P_{Tot}} = \frac{(0.90 \text{ mol})(0.08206 \text{ L•atm/mol•K})(298.15 \text{ K})}{(5.50 \text{ atm})} = 4.\underline{0}03 \text{ L}$$

$$P_{He} = \frac{n_{He}RT}{V} = \frac{(0.25 \text{ mol})(0.08206 \text{ L•atm/mol•K})(298.15 \text{ K})}{(4.\underline{0}03 \text{ L})} = 1.\underline{5}2 \text{ atm}$$

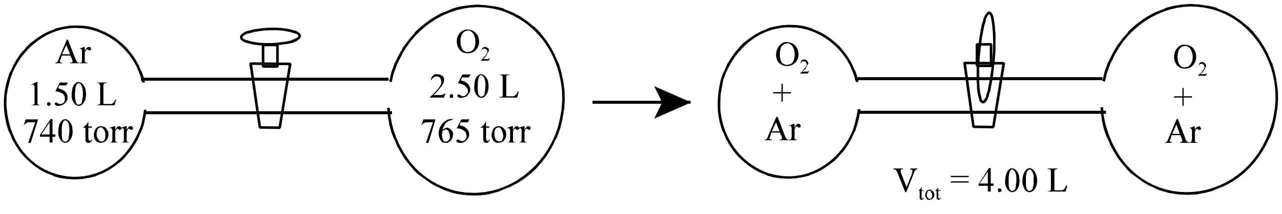
$$P_{N2} = \frac{n_{N2}RT}{V} = \frac{(0.65 \text{ mol})(0.08206 \text{ L•atm/mol•K})(298.15 \text{ K})}{(4.\underline{0}03 \text{ L})} = 3.\underline{9}7 \text{ atm}$$

2. (6 pts) A 1.50 L container of Ar at 740.0 torr and 25.0 °C is connected to a 2.50 L container of O_2 at 765.0 torr and 25.0 °C. What is the **<u>total</u> <u>pressure</u>** (torr) after the gases have mixed if the temperature remains at 25.0 °C? (Atomic weights: O = 16.00, Ar = 39.95)

Use Dalton's Law of Partial Pressures,

$$P_{Total} = P_1 + P_2 + P_3 + \bullet\bullet\bullet$$

Each gas acts independently of the other gases and each follows the IGL as does P_{Tot}.



Open the valve and allow the gases to mix.

Now each gas occupies the <u>**ENTIRE**</u> **volume** of both flasks, 4.0 L.

Need to find new Pressure of each gas in "new" 4.0 L container - Use Boyle's Law

$$\text{Ar} \qquad P_{Ar,2} = \frac{P_{Ar,1}\,V_1}{V_2} = \frac{(740\text{ torr})\,(1.5\text{ L})}{4.0\text{ L}} = 27\underline{7}.5\text{ torr}$$

$$O_2 \qquad P_{O2,2} = \frac{P_{O2,1}\,V_1}{V_2} = \frac{(765\text{ torr})\,(2.5\text{ L})}{4\,.0\text{ L}} = 47\underline{\mathbf{8}}.125\text{ torr}$$

$P_T = 27\underline{7}.5 + 47\underline{\mathbf{8}}.125 = 75\underline{\mathbf{5}}.625\text{ torr} = 756\text{ torr}$ (using DLPP)

3. (3 pts) Which gases, N_2O, C_2H_2, NO, diffuse **<u>faster</u>** than O_2 under identical experimental conditions? (atomic weights: N = 14.01, O = 16.00, C = 12.01, H = 1.008)

The rate of effusion is directly proportional to the molecular speed which is inversely proportional to the square root of the molar mass. Thus, lighter particles have a faster average velocity and faster rate of effusion.

$$u = \left(\frac{3\,RT}{\mathcal{M}}\right)^{1/2} \qquad\qquad r_{eff} \propto u$$

So $r_{eff} \propto 1/(\mathcal{M})^{1/2}$ at contant T ($\mathcal{M}$ = molar mass, i.e. molec. wt.)

So rate of effusion is proportional to $1/(\mathcal{M})^{1/2}$ (inversely prop to $\mathcal{M}^{1/2}$).

$\therefore$ smaller $\mathcal{M}$ => greater r_{eff}

Thus, a lighter (smaller) gas effuses and diffuses at a faster rate.

O_2 (32.00) N_2O (44.02) C_2H_2 (26.04) NO (30.01)

Both **NO** and **C_2H_2** are **lighter** than the **O_2** so they **effuse and diffuse <u>faster</u>**.

4. (3 pts) Which one of the statements below is **FALSE**?

a) A real gas behaves more nearly as an ideal gas at high temperatures and low pressures.

b) In the van der Waals equation, one would expect a larger value of "a" for HF than for He.

c) Gases approach their liquefaction points as temperature decreases and as pressure increases.

d)* The "b" factor in the van der Waals equation should be larger for He than for Cl_2.

e) It is not possible for the “a” factor in the van der Waals equation to be negative.

In the van der Waals equation (below), the “b” factor corrects for molecular volume and the “a” factor corrects for intermolecular attractive forces.

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

causes **(+)** deviation	causes **(−)** deviation
$P_m > P_i$	$P_m < P_i$
due to vol of molecules	due to IAF

P_m = measured (real) P; P_i = ideal P

Thus, larger molecule => larger molecular volume => larger b

b: corrects for finite molecular volume of real gases (b=0 for an ideal gas)

a: corrects for attractive forces between real molecules (a=0 for an ideal gas)

D is false: The “**b**” factor for **He** is <u>**smaller**</u> than that for $\mathbf{Cl_2}$, since **He** is smaller.

a) real gases behave nearly ideally at high T (where KE is great enough to overcome IAF) and low P (larger container volume, in which case the molecular volume is less important). As the temp dec. and pressure inc. a gas behaves less ideally and approaches its liquefication point.

b) It would be expected the “a” value would be larger for HF than He. HF is larger and has stronger attractive forces. It has London Dispersion Forces, Dipole-Dipole and Hydrogen Bonding while He has only LF. In the gas phase, due to the movements of a gas the more important factor is likely to be the LF, which are larger in HF since it’s larger than He (although just above the b.p. for HF it still forms dimers, two HF molecules held together via HB).

c) As temp dec (cools and molecules slow down) and pressure inc. (forces molecules closer together) it will eventually liquefy.

e) The “a” and “b” factors for an ideal gas are both zero. Ideal gas molecules experience no att. forces and take up now space (like points). The “a” can’t be negative. That would mean the gases are repelling each other. We’re not talking about plasmas, like charged particles in the gas phase (which would repel each other).

5. (5 pts) The pressure of a 2.855 moles sample of O_2 contained in a 1.00 L flask is140.0 atm at 300°C. One can conclude (to 3 significant figures) that

a) The gas is behaving ideally.
b)* The gas exhibits a positive deviation from ideality.
c) The gas exhibits a negative deviation from ideality.
d) The collision of the molecules with the walls of the container are no longer elastic
e) no statement can be made without the Van der Waals constants.

Want to determine if the gas is behaving ideally or if not is the deviation from ideality a negative or positive deviation since this determines whether the a or b factor in the the **van der Waals** equation is more important (attractive forces or molecular volume):

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

or

$$P = \frac{nRT}{V - nb} - \frac{n^2a}{V^2}$$

causes **(+)** deviation	causes **(−)** deviation
$P_m > P_i$	$P_m < P_i$
due to vol of molecules	due to IAF

P_m = measured (real) P; P_i = ideal P

Need to determine if the data indicates ideal behavior or not.

Two ways to do this problem from the given information.

<u>Use data given & use IGL</u>

For ideal gas, $\frac{PV}{nRT} = 1$ for all P

Calc. PV/nRT for 1.00 L

$$\frac{PV}{nRT} = \frac{(140 \text{ atm})\,(1.00 \text{ L})}{(2.855 \text{ mol})\,(0.08206 \text{ L•atm/mol•K})\,(573 \text{ K})} = 1.0\underline{4}28 > 1 \quad \text{(to 3 s.f.)}$$

<u>Pos</u>. dev.

<u>Or</u>

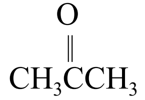
Calc. P_i and compare to P_m

$$P_i = \frac{(2.855 \text{ mol})\,(0.08206 \text{ L•atm/mol•K})\,(573 \text{ K})}{(1.00 \text{ L})} = 13\underline{4}.24 \text{ atm}$$

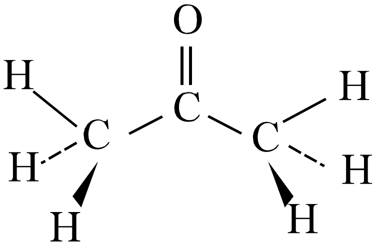
$P_m > P_i$ **<u>Pos</u>**. dev.

A positive deviation is caused by molecular vol., so the volume of the gas molecules must take up a significant fraction of the volume of the container under these conditions and is the more impt. factor in this case.

6. (3 pts) Which of the following compounds is **<u>INCORRECTLY</u>** paired with the intermolecular forces that exist between neighboring molecules?

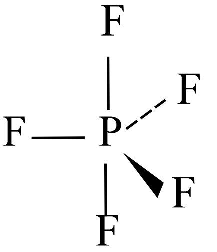
a) PF_5	London forces only
b) $AsCl_3$	London forces, dipole-dipole forces
c) $SiCl_4$	London forces only
d) CH_3CH_2F	London forces, dipole-dipole forces
e)*  $CH_3\overset{\overset{O}{\|}}{C}CH_3$	London forces, dipole-dipole forces, hydrogen bonding

Determine if molecules are nonpolar or polar (Draw Lewis Structure and determine shape). Also, remember that in general if there is one or more lone pair of electrons (lpe−) on the "central" atom the molecule will be polar. The exceptions to this rule is a linear molecule resulting from 5 e− pairs (3 atoms and 2 lpe−) and a square planar molecule resulting from 6 e− pairs (4 atoms and 2 lpe−), which can be nonpolar.



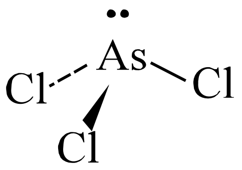
Tetrahedral around CH_3 C atoms & trigonal planar around C=O C atom.
POLAR (around C atoms, especially C=O group)
LF & **DD** (NO H bonding - since no H on the O)

a) PF_5



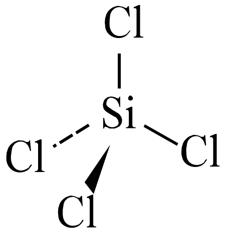
Trigonal bipyramidal
<u>Non</u>polar (All 5 atoms around P are indentical.)
LF <u>only</u>

b) $AsCl_3$



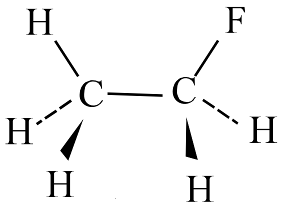
Trigonal pyramidal (like $:NH_3$,$:NF_3$,$:PH_3$, $:PF_3$, etc.)
<u>Polar</u> ; **LF** & **DD**
(Usually lpe− on central atom ⇒ polar)

c) $SiCl_4$



Tetrahedral (like CH_4, CF_4, SiH_4, etc.)
<u>Non</u>polar (all 4 atoms on Si are identical)
LF <u>only</u>

d) CH_3CH_2F



Tetrahedral around C atoms
POLAR (around the C atoms, atoms on C's are not identical)
LF & **DD** (No H bonding - since no H on the F)

7. (8 pts) Examine the following compounds and then answer the questions below which refer to these compounds, as pure substances. (**<u>Give a short explanation for your choices</u>**, i.e, types of **attractive forces**, whether **polar** or **nonpolar**, etc.)

LF ≡ London Forces D-D ≡ Dipole-Dipole IAF ≡ intermolecular attractive forces

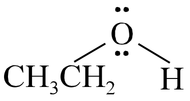
	C_3H_8	$CH_3CH_2 - \overset{..}{\underset{..}{O}} - H$	$CH_3 - \overset{..}{\underset{..}{O}} - CH_3$	CH_2F_2
	nonpolar	polar	polar	polar
MW ≈	44 amu	46 amu	46 amu	52 amu

a) Which compound(s) has(have) **<u>only</u> London forces**?

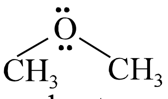
C_3H_8 nonpolar (only C and H in formula - nonpolar)

Nonpolar molecules have **<u>ONLY</u> LF**

b) Which compound(s) has(have) **<u>Dipole-Dipole</u>** forces?



bent



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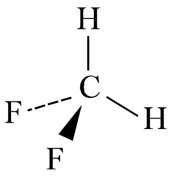
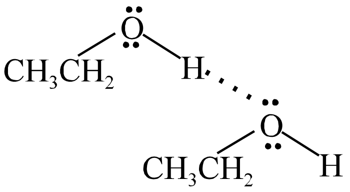


tetrahedral

<u>polar</u> molec. have D-D forces
(In addition to LF)

c) Which compound(s) can form **H-bonds** between molecules (as a pure substance)?



This molec. can form H-bonds between two such molecules. There must be a H atom covalently bonded to an N, O or F atom and a lone pair e^- on one of those atoms (This is in addition to LF and DD att. forces.)

(CH_3OCH_3 and CH_2F_2 are polar but can <u>NOT</u> form H-bonds.)

d) Which compound should have the **<u>highest</u>** heat of vaporization, ΔH_{vap}?

CH_3CH_2OH has LF, D-D & H-bonding. The first 3 molecules are about the same size so the LF are about the same. CH_2F_2 isn't that much different in size than the other three so it's LF isn't that much different than first 3 (the CH_2F_2 LF may be slightly larger). The C_3H_8 is nonpolar and has only LF. CH_3CH_2OH, CH_3OCH_3 and CH_2F_2 are polar and have DD IAF in addition to LF (CH_3CH_2OH is more polar and has larger DD IAF). Most importantly, CH_3CH_2OH can form H-bonds between CH_3CH_2OH molecules (whereas CH_3OCH_3 and CH_2F_2 can't as pure substances since the H atoms are bonded to the C atoms and not an N, O or F atom).

∴ **CH_3CH_2OH** has **<u>strongest</u> IAF** & **<u>highest</u>** ΔH_{vap}

Remember (generally): **<u>Stronger</u>** Att. Forces ⇒ **<u>Lower</u> Vapor Pressure**
<u>Slower</u> rater of **evaporation**

<u>Higher</u> b.p., m.p., sublimation pt.
ΔH_{vap}, ΔH_{fus}, ΔH_{sub},
critical temp., surface tension, viscosity

8. (3 pts) Which of the following statements is **<u>INCORRECT</u>**?

a) Br_2 has a **higher** boiling point than Cl_2.

b) H_2S has **stronger** London forces than H_2O.

c) The vapor pressure of solid He is **higher** than the vapor pressure of solid Ne at a given temperature.

d) CH_3-CH_2-OH has a **greater** viscosity than CH_3-O-CH_3.

e)* CO_2 has a **higher** critical temperature than SO_2.

a) T: b.p. (Br_2) > b.p. (Cl_2)

	b.p. (Br_2)	b.p. (Cl_2)
MW:	160	71
# e− :	70	34

Br_2		Cl_2
LF	>	LF

The Br_2 and Cl_2 molecules are both nonpolar. The only AF that nonpolar molecules have is LF. LF increase as the size of the molecule increases (it does depend on shape as well but both molecules are linear). Br_2 is bigger and has larger LF, thus, it has stronger IAF and a higher boiling pt.

b) T: LF (H_2S) > LF (H_2O)

	LF (H_2S)	LF (H_2O)
MW:	34	18
# e− :	18	10

H_2S		H_2O
LF	>	LF
DD	<	DD
		H-bonds

H_2S is larger than H_2O and thus has larger LF. In relation to the total AF and properties, H_2O has stronger DD and H-bonding that H_2S doesn't have. The H-bonding that occurs between H_2O molecules is a very strong IAF compared to the other forces and is in addition to the other IAF and is the main reason for the higher IAF of H_2O compared to other molecules of approximately same size.

c) T: VP(He) > VP(Ne)

	VP(He)	VP(Ne)
MW:	4	32
# e− :	18	18

He		Ne
LF	<	LF

Both are noble gas atoms. Since they are atoms the only AF between them is London Forces. LF depend on the size and Ne is bigger than He and thus Ne has larger LF and lover vapor pressure.

d) T: Viscosity (CH_3CH_2OH) > Viscosity (CH_3OCH_3)

	Viscosity (CH_3CH_2OH)	Viscosity (CH_3OCH_3)
MW:	32	32
# e− :	18	18

CH_3CH_2OH		CH_3OCH_3
LF	≈	LF
DD	>	DD
H-bonds		

Both molecules are polar. The molecules are the same size and have similar LF. The CH_3CH_2OH molecules are more polar than the CH_3OCH_3 molecules (because of the fairly polar OH group) so CH_3OH has greater DD forces. However, only CH_3CH_2OH can form H-bonds between molecules in the pure substance. Thus CH_3CH_2OH has stronger IAF and a greater viscosity.

e) F: critical temp (CO_2) < critical temp(SO_2)

	critical temp (CO_2)	critical temp(SO_2)
MW:	44	64
# e− :	22	32

CO_2		SO_2
LF	<	LF
		DD

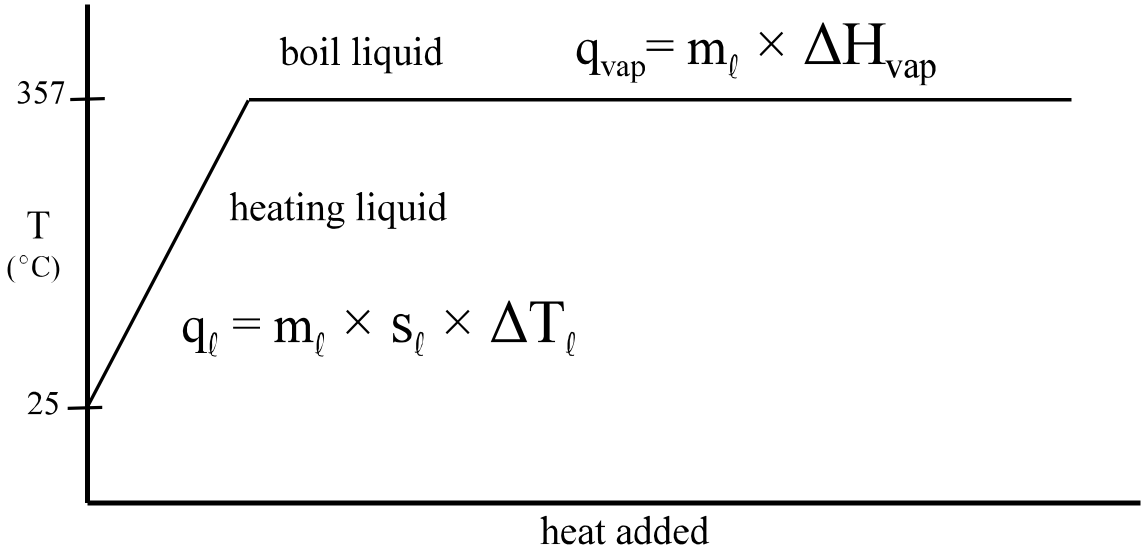
SO_2 has stronger IAF than the nonpolar CO_2 and thus a <u>higher</u> critical temperature.

Remember: **<u>Stronger</u>** Att. Forces ⇒ **<u>Lower</u> Vapor Pressure**
<u>Slower</u> rater of **evaporation**

<u>Higher</u> b.p., m.p., sublimation pt.
ΔH_{vap}, ΔH_{fus}, ΔH_{sub},
critical temp., surface tension, viscosity

9. (5 pts) Calculate the amount of heat (kJ) required to heat 125 g of mercury (Hg) from 25.0 ºC to its boiling point (357 ºC) and then vaporize it? (specific heat of liquid Hg = 0.138 J/g•ºC, ΔH_{vap} = 292 J/g)

Draw a heating curve and label it (makes it easier)



$$q = q_{\text{heat liquid}} + q_{\text{boiling}}$$

$$= (m_\ell \bullet s_\ell \bullet \Delta T_\ell) + (m_\ell \bullet \Delta H_{vap})$$

$$= [125\text{ g} \bullet 0.138\text{ J/g-°C} \bullet (357 - 25.0)\text{°C}] + (125\text{ g} \bullet 292\text{ J/g})$$

$$= 57\underline{\mathbf{2}}7\text{ J} + 36\underline{\mathbf{5}}00\text{ J} = 42\underline{\mathbf{2}}27\text{ J} = 42.2\text{ kJ}$$

10. (6 pts) A substance has a ΔH_v = 20.0 kJ/mol. It has a vapor pressure of 0.800 atm at -2.00 ºC. What is it's **normal boiling point**?

$$\ln\left(\frac{P_2}{P_1}\right) = \frac{\Delta H_v}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right) \quad \text{or} \quad \log\left(\frac{P_2}{P_1}\right) = \frac{\Delta H_v}{2.303\ R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

$P_1 = 0.800$ atm $\qquad$ $P_2 = 1.00$ atm (normal b.p. occurs at 1 atm)

$T_1 = -2.00\,°C + 273.15 = 271.15K$ $\qquad$ $T_2 = ?$ (the normal b.p.)

T **MUST** be in kelvin (K)

$$\ln\left(\frac{1.00}{0.800}\right) = \frac{20.0 \text{ x } 10^3 \text{ J/mol}}{8.314 \text{ J/mol•K}}\left(\frac{1}{271.15 \text{ K}} - \frac{1}{T_2}\right)$$

$$0.22\underline{\mathbf{3}}14 = 2.4\underline{\mathbf{0}}558 \text{ x } 10^3 \text{ K}\left(3.687\underline{\mathbf{9}}9 \text{ x } 10^{-3} \text{ K} - T_2^{-1}\right)$$

$$T_2^{-1} = 3.592\underline{\mathbf{3}}4 \text{ x } 10^{-3} \text{ K}^{-1} \qquad T_2 = 278.1\underline{\mathbf{4}}5 \text{ K} = 4.9\underline{\mathbf{9}}59\,°C = 5.00\,°C$$

higher T ⟶ higher VP (and vice versa)