Practice Test Unit 3 - Gases

Part I: *Circle the letter of the best annswer(s). No calculators are allowed on this portion of the practice test!*

Goal: I can describe the general characteristics of gases as compared to the other states of matter, and list the ways in which gases are distinct.

A

1. Circle all of the true statements below.
   a. Gases are more easily compressed than either liquids or solids.
   b. There is more attractive forces between gaseous particles than liquid particles.
   c. Gases tend to be more dense than solids.
   d. Gases have less space between particles than liquids.
   e. The properties of liquids are more affected by pressure than gases.

Goal: I can define atmosphere, torr, mmHg and pascal, the most common units in which pressure is expressed. I can also describe how a barometer and a manometer work.

B

2. Which of the following is/are TRUE statements?
   a. 1 atm = 1 torr
   b. 1 torr = 1 mmHg
   c. 1 mmHg = 1 mmH_2O
   d. The higher the atmospheric pressure, the lower the column of Hg in a barometer.
   e. Manometers measure gas pressure only whereas barometers measure ga pressure & temp.

Goal: I can describe how a gas responds to changes in pressure, volume, temperature, and quantity of gas.

D

3. Consider a sample of gas confined at constant temperature and volume in the closed system. If more of this same gas is added at constant temperature, what effect is observed on pressure and average molecular velocity?
   a. Both pressure and average molecular velocity increase.
   b. Pressure decreases and average molecular velocity remains the same.
   c. Pressure remains the same and average molecular velocity increases.
   d. Pressure increases and average molecular velocity remains the same. Same Temp = Same Velocity
   e. Pressure remains the same and average molecular velocity decreases.

C

4. Which describes a change that occurs when a sample of nitrogen is sealed in a metal tank then heated from 250 K to 300 K?
   a. The density of the sample decreases.
   b. The volume of the sample increases. metal tank means constant volume
   c. The pressure of the sample increases.
   d. The mean distance between molecules increases. Gases fill volume no matter V
   e. The number of molecules in the container increases.

B 5. When a sample of oxygen gas in a closed container of constant volume is heated until its absolute temperature is doubled, which of the following is also doubled?

a. The density of the gas.
b. The pressure of the gas.
c. The average velocity of the gas molecules.
d. The number of molecules per cm^3.
e. The potential energy of the molecules.

$n, R, V =$ constant
$PV = nRT$
$P = T$ 2T (double)
$P = 2T$

Goal: I can use the gas laws, including the combined gas law, to calculate how one variable of a gas (P, V, n, or T) responds to changes in the one or more of the other variables.

E 6. A sample of ideal gas was heated at constant volume from 25°C to a temperature sufficient to exactly double the pressure. What was the volume of the gas sample?

a. 22.4 liters
b. 20.5 liters
c. 24.4 liters
d. 44.8 liters
e. cannot be determined

A 7. A gas occupies a 1.5 liter container at 25°C and 2.0 atm. If the gas is transferred to a 3.0 liter container at the same temperature, what will be the new pressure?

a. 1.0 atm
b. 2.0 atm
c. 3.0 atm
d. 4.0 atm
e. 5.0 atm

$V_1 = 1.5L$
$V_2 = 3.0L$
$P_1 = 2.0 Atm$
$P_2 = ?$

$PV = nRT$
$P_1 V_1 = P_2 V_2$
$P_2 = \frac{P_1 V_1}{V_2} = \frac{(2.0Atm)(1.5L)}{2\ 3.0L}$
$P_2 = 1.0 Atm$

B 8. Of the following, ________ is a valid statement of Charles' law.

a. P/T = constant
b. V/T = constant
c. PV = constant
d. V = constant * n
e. V = constant * P

$\cancel{P}V = nRT$
$\frac{V_1}{T_1} = \frac{V_2}{T_1}$
Prince = P constant
~constant

Goal: I can use the ideal-gas equation to solve for one variable (P, V, n, or T) given the other three variables or information from which they can be determined.

A 9. A sample of pure gas at 27°C and 380 mm Hg occupied a volume of 492 mL. What is the number of moles of gas in this sample?

a. 0.010 moles
b. 7.6 moles
c. 10 moles
d. 6×10^{21} moles
e. none of these values

$T = 27 + 273 = 300$
$P = 380mm\left(\frac{1Atm}{760}\right) = .5Atm$
$V = 492ml = .492L$
$n = ?$

$PV = nRT$
$n = \frac{PV}{RT} = \frac{(.5Atm)(.492L)}{(.08206)(300K)}$

Less than 1
Small value

Goal: I can calculate the molar mass of a gas, given gas density under specified conditions of temperature and pressure. I can also calculate gas density under stated conditions, knowing molar mass.

B

10. A 22.4 liter sample of gas at STP weighs 16.0 grams. What is the molecular weight of the gas?
    a. 22.4 g/mol
    b. 16.0 g/mol
    c. 29.4 g/mol
    d. 12.0 g/mol
    e. 32.0 g/mol

$MM = ?\ (\frac{g}{mole})$ $\quad MM = \frac{dRT}{P}$ @ STP

B

11. Which of the following would express the approximate density of carbon dioxide gas at 0°C and 2.00 atm pressure (in grams per liter)?
    a. 2 g/L
    b. 4 g/L
    c. 6 g/L
    d. 8 g/L
    e. None of the above

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

$d = \frac{MM\,P}{RT} = \frac{(44.0\frac{g}{mole})(2.00\,atm)}{(.08206)(273\,K)} \approx 4$

$T = 0° = 273\,K$
$P = 2.00\,atm$
$CO_2 = 44.0\,g/mole$

C

12. When 2.00 grams of a certain volatile liquid is heated, the volume of the resulting vapor is 821 mL at a temperature of 127°C at standard pressure. The molecular weight of this substance is ________.
    a. 20.0 g/mole
    b. 40.0 g/mole
    c. 80.0 g/mole
    d. 120.0 g/mole
    e. 160.0 g/mole

$MM = dRT$ $\quad mm = \frac{g}{mole}$

$MM = \frac{mass}{V} RT$

$= \frac{(\frac{2.0g}{.821L})(.08206)(400.K)}{1\,atm}$

$\approx 80\ g/mole$

$V_1 = 821\,mL = .821\,L$
$T_1 = 127°C = 400.\,K$
$P_1 = 1.0\,atm$
$D = \frac{mass}{V}$

B

13. Of the following, ________ correctly relates pressure, volume, temperature, molar mass (M), density (d), and mass (g).

a) $M = \frac{dRT}{PV}$

b) $M = \frac{gRT}{PV}$

c) $M = \frac{PT}{gRV}$

d) $M = \frac{gV}{RT}$

e) $M = \frac{RT}{gV}$

$MM = \frac{dRT}{P}$ $\quad d = \frac{g}{V}$

$MM = \frac{g\,RT}{VP}$

B

14. Which **one** of the following statements about the density of a gas is correct?
   a. It is independent of temperature.
   b. It decreases with increasing temperature at constant pressure.
   c. It is independent of pressure.
   d. It decreases with increasing pressure at constant temperature.
   e. It doubles when the volume of a container doubles without a change in pressure or temperature.

E

15. The density of an unknown gas is found to be 1.65 $g \cdot L^{-1}$. Under the same conditions, the density of oxygen gas is found to be 1.10 $g \cdot L^{-1}$. The molecular mass of the unknown gas is closest to:

a. 14 b. 24 c. 28 d. 32 e. 48

$MM = \frac{dRT}{P}$ R, T, P constant

$\frac{MM_{un}}{d_{un}} = \frac{MM_{O_2}}{d_O}$ MM x

≈ 32 ⟹ $MM_u = \frac{MM_O \, d_u}{d_O} = \frac{(32.0)(1.65)}{1.10}$

only E is larger

Goal: I can solve gas stoichiometry problems at standard conditions and at non-standard conditions.

B

16. The combustion of carbon monoxide yields carbon dioxide. The volume of oxygen gas needed to produce 22 grams of carbon dioxide at STP is: $2CO + O_2 \rightarrow 2CO_2$?L 22g

a. 4.0 liters b. 5.6 liters c. 11 liters d. 22 liters e. 32 liters

$\left(\frac{22g\ CO_2}{1}\right)\left(\frac{1\ mole\ CO_2}{44.0g\ CO_2}\right)\left(\frac{1\ mole\ O_2}{2\ mole\ CO_2}\right)\left(\frac{22.4L\ O_2}{1mole\ O_2}\right) = \frac{22.4}{4} ≈ 5...$

B

17. A 27.0 g sample of an unknown hydrocarbon is burned in excess oxygen to form 88.0 grams of carbon dioxide and 27.0 grams of water. What is a possible molecular formula of the hydrocarbon?
   a. CH_4
   b. C_2H_2
   c. C_4H_3
   d. C_4H_6
   e. C_4H_{10}

H Has to come

$C_xH_x + O_2 \rightarrow H_2O + CO_2$
excess 27.0g 88.0g

$C_2H_3 = C_4H_6$

H $\left(\frac{27.0g H_2O}{1}\right)\left(\frac{1 mole H_2O}{18.0g H_2O}\right)\left(\frac{2 mole H}{1 mole H_2O}\right) = 3.0$ mole H /2 = 1.5 x2 3

C $\left(\frac{88.0g CO_2}{1}\right)\left(\frac{1 mole CO_2}{44.0g CO_2}\right)\left(\frac{1 mole C}{1 mole CO_2}\right) = 2$ mole C /2 ≠ 1, 2 2

B

18. A sample of 9.00 g of aluminum metal is added to an excess of HCl. The volume of hydrogen gas produced at standard temperature and pressure is:
   a. 22.4 liters
   b. 11.2 liters
   c. 7.46 liters
   d. 5.60 liters
   e. 3.74 liters

$2Al + 6HCl \rightarrow 3H_2 + 2AlCl_3$
9.0g excess ?L STP

$\left(\frac{9.0g Al}{1}\right)\left(\frac{1 mole Al}{26.98g Al}\right)\left(\frac{3 mole H_2}{2 mole Al}\right)\left(\frac{22.4L}{1 mole}\right) ≈ 11$ L

A

19. $3\,Ag_{(s)} + 4\,HNO_3 \rightarrow 3\,AgNO_3 + NO_{(g)} + 2\,H_2O$

.1 m 6.0M 10.0L ? moles

The reaction of silver metal and dilute nitric acid proceeds according to the equation above. If 0.10 moles of powdered silver is added to 10.0 mL of 6.0 molar nitric acid, the number of moles of NO gas that can be formed is:

$\left(\frac{6.0 moles}{L}\right)\left(\frac{.01 L}{1}\right) = .06$ moles

a. 0.015 mole b. 0.020 mole c. 0.030 mole
d. 0.045 mole e. 0.090 mole

3 Ag	limiting 4 HNO3	NO
.1	.06	0
3x	4x	x

x = .015

.015 mole produced

Goal: I can calculate the partial pressure of any gas in a mixture, given the composition of that mixture. (Dalton's Law of Partial Pressures)

E

20. What is the partial pressure of helium when 8.0 grams of helium and 16 grams of oxygen are in a container with a total pressure of 5.00 atm?

$\left(\frac{8.0g\ He}{1}\right)\left(\frac{1\ mole\ He}{4.0g\ He}\right) = 2\ mole\ He$

$\left(\frac{16g\ O_2}{1}\right)\left(\frac{1\ mole\ O_2}{32.00g\ O_2}\right) = .5\ mole\ O_2$

a. 0.25 atm b. 1.00 atm c. 1.50 atm d. 2.00 atm e. 4.00 atm

$P_p = \frac{He}{He+O_2} P_T = \frac{2}{2+.5}(5.0) = 4.00\ Atm$

Goal: I can calculate the mole fraction of a gas in a mixture, given its partial pressure and the total pressure of the system.

B

21. One mole of nitrogen, two moles of neon and four moles of argon are sealed in a cylinder. The combined pressure of the gases is 1400 mm Hg. What is the partial pressure of nitrogen in the cylinder?
   a. 100 mm Hg
   b. 200 mm Hg
   c. 400 mm Hg
   d. 500 mm Hg
   e. 1400 mm Hg

$P_N = \frac{N}{N+Ne+Ar} P_T$

$= \frac{1}{1+2+4}\ 1400\ mm$

$= \frac{1}{7}\ 1400 = 200$

Goal: I can explain the technique of collecting a gas "over water" and make the associated calculations.

22. A sample of zinc metal reacts completely with excess hydrochloric acid according to the following equation:

T = 40°C = 313 K

$P_1 = 720\ mm$

$P_V = 55\ mm\ Hg$

$$Zn(s) + \overset{excess}{2HCl(aq)} \rightarrow ZnCl_2(aq) + H_2(g)$$

8.00 liters of hydrogen gas at 720. mm Hg is collected over water at 40.°C (vapor pressure of water at 40.°C = 55 mm Hg). How much zinc was consumed by the reaction?

a. $\frac{(720/760)\bullet 8.00}{(0.0821)\bullet 313}$

b. $\frac{(760/720)\bullet 313}{(0.0821)\bullet 2}$

c. $\frac{(665/760)\bullet 8.00\bullet (65.39)}{0.0821\bullet 313}$

d. $\frac{(665/760)\bullet 8.00}{(65.39)\bullet (0.0821)\bullet 313}$

e. $\frac{8.00\bullet 313\bullet 65.39}{(665/760)\bullet (0.0821)}$

Need partial Pressure

720 − 55

= 665 mm / 760

PV = nRT To make Atm

$n = \frac{PV}{RT} = \frac{\left(\frac{665}{760}\right)(8.0\ L)}{(.08206)(313)}$

$\left(\frac{8.0L\ H_2}{1}\right)$

$\left(\frac{n\ H_2}{1}\right)\left(\frac{1\ mole\ Zn}{1\ mole\ H_2}\right)\left(\frac{65.39}{1\ mole}\right)$

$\frac{(665)(8.0)\ 65.39}{(.08206)(313)}$

23. When a gas is collected over water, the pressure is corrected by

C

a. adding the vapor pressure of water.
b. multiplying by the vapor pressure of water.
c. subtracting the vapor pressure of water at that temperature.
d. subtracting the temperature of the water from the vapor.

Goal : I can use the Kinetic Molecular Theory to explain the behavior of gases.

$(100g\ O_2)\left(\frac{1\ mole\ O_2}{32.0g\ O_2}\right) \approx 3$ mole O_2 $(100g\ He)\left(\frac{1\ mole\ He}{4.0g\ He}\right) = 25$ mole He

24. 100 grams of $O_2(g)$ and 100 grams of He(g) are in separate containers of equal volume. Both gases are at 100°C. Which **one** of the following statements is true?

E

a. Both gases would have the same pressure. NO
b. The average kinetic energy of the O_2 molecules is greater than that of the He molecules. Same Vol
c. The average kinetic energy of the He molecules is greater than that of the O_2 molecules. & Same temp
d. There are equal numbers of He molecules and O_2 molecules. NO 3 mole vs 25 mole
e. The pressure of the He(g)would be greater than that of the $O_2(g)$.

25

$V_N = 1$
$V_H = 3$

25. Which **2** of the following are **TRUE**?

$\frac{V_1}{n_1} = \frac{V_2}{n_2}$ ← # moles
$\frac{1}{} = 3$ NO 1+3=4

a. According to Avogadro's hypothesis, when one volume of nitrogen reacts with three volumes of hydrogen to form ammonia, four volumes of ammonia should form.
b. According to Graham's law, you should expect NH3(g) to effuse faster through a tiny hole than CO2(g). 44g 17g Lower mass faster
c. You would expect a gas at ~~high~~ Low pressures to behave like an ideal gas.
d. The vapor pressure of ether is greater than the vapor pressure of ethanol. Therefore you should expect more vapor above ether than ethanol in a closed system. P_{ether} ↑ $P_{ethanol}$
e. When you measure the rate of NH3 spreading throughout a long tubee you are measuring the rate of effusion of NH3. NO into a chamber flow Rate!

Goal: Explain the concepts of effusion and diffusion and make associated calculations using Graham's Law.

B

26. Graham's Law refers to:
a. boiling points of gases.
b. gaseous effusion.
c. gas compression problems.
d. volume dependence upon temperature.

27. The diffusion time for carbon dioxide gas was 105 sec. For gas X, 126 sec was required for the same number of moles of gas to diffuse under the same conditions. What is the approximate molecular weight of the unknown gas?

$R = \frac{D}{T}$

a. 12 g/mol $CO_2 = 44.0g/mole$
b. 24 g/mol
c. ~~37~~ g/mol 30
d. 44 g/mol
e. 63 g/mol

$$\frac{T_{CO_2}}{T_{un}} = \sqrt{\frac{mm_{un}}{mm_{CO_2}}}$$

$mm_{un} = 30$

$$\frac{105}{126} = \sqrt{\frac{?}{44}}$$

$$44\left(\frac{105}{126}\right)^2 = mm_{un}$$

C

28. If the average velocity of a methane molecule, CH_4, is 5.00×10^4 cm/sec at 0°C, what is the average velocity of helium molecules at the same temperature and pressure conditions?

a. 2.50×10^4 cm/sec
b. 5.00×10^4 cm/sec
c. 1.00×10^5 cm/sec
d. 2.00×10^5 cm/sec
e. 5.00×10^5 cm/sec

$CH_4 = 16.0 g/mole$ $He = 4.0 g/mole$ T, P Constant $V_{CH_4} = 5.00 \times 10^4 cm/sec$ $V_{He} = ?$

$$\frac{V_{He}}{V_{CH_4}} = \sqrt{\frac{mm_{CH_4}}{mm_{He}}}$$

$$V_{He} = \sqrt{\frac{16}{4}}(5.00 \times 10^4 cm/sec) = 2(5.00 \times 10^4) = 1.0 \times 10^5$$

C

29. According to the kinetic molecular theory, in which of the following gases will the root-mean-square speed of the molecules be the highest at 200°C ?

a. HCl 36g   b. Cl_2 70g   c. H_2O 18g   d. SF_6 200g heavy
e. None . .. the rms speeds are the same!

$$u_{rms} = \sqrt{\frac{3RT}{mm}}$$

∴ fastest has smallest mm

C

30. Which change increases the mean free path of molecules in a sample of gas?

a. Increase in pressure at constant temperature.
b. Increase in density at constant temperature.
c. Increase in temperature at constant pressure.
d. Increase in temperature at constant volume.
e. Increase in pressure at constant volume.

Heat them up → Speed them up

D

31. Under certain conditions, methane gas, CH_4, diffuses at a rate of 12 cm/s. Under the same conditions, an unknown gas diffuses at a rate of 8.0 cm/s. The molecular mass of the unknown gas is closest to:

CH_4: 16.0g, 12 cm/s   Un: ?, 8.0 cm/s

a. 6   b. 20   c. 24   d. 36   e. 72

$$\frac{12}{8} = \sqrt{\frac{mm_{Un}}{16.0g}} \qquad \left(\frac{3}{2}\right)^2 16 = M.M_{Un} \qquad \frac{9}{4}16 = 36$$

B

32. When a sample of ideal gas is heated from 20° C to 40° C, the average kinetic energy of the system changes. Which factor describes this change?

T_1 20 + 273 = 293
T_2 40 + 273 = 313

a. $\frac{1}{2}$
b. $\frac{313}{293}$
c. $\sqrt{\frac{313}{293}}$
d. $\frac{293}{313}$
e. 2

Goal: I can cite the general conditions of T and P under which real gases most closely approximate ideal-gas behavior. Also, I can explain the origin of the correction terms P and V that appear in the van der Waals equation.

33. For a substance that remains a gas under the conditions listed, deviation from the ideal gas law would be most pronounced at:

Ideal gas = gas behaves more like at ↑T ↓P

∴ deviation would happen @ opposite ↓T ↑P

A

a. -100°C and 5.0 atm
b. -100°C and 1.0 atm
c. 0°C and 1.0 atm
d. 100°C and 1.0 atm
e. 100°C and 5.0 atm

34. An ideal gas differs from a real gas in that the molecules of an ideal gas . . .

A

a. have no attraction for one another.
b. have appreciable molecular volumes.
c. have a molecular weight of zero.
d. have no kinetic energy.
e. has an average molecular mass.

Key

Part II: Solve each of the following problems. Show a set-up for each. Label your final answers with appropriate units and **box** your answers.

1. A sample of gas at 15°C and 1 atm has a volume of 2.58 L. What volume (in L) will this gas occupy at 38°C and 1 atm?

Given

$T_1 = 15°C = 288K$ $T_2 = 38°C = 311K$

$P =$ constant $V_2 = ?$

$V_1 = 2.58L$

Soln.

$PV = nRT$

$\frac{V}{T} = \frac{nR}{P}$

$\frac{V_1}{T_1} = \frac{V_2}{T_2}$

$V_2 = \frac{V_1}{T_1} T_2$

$= \frac{(2.58L)(311K)}{288K}$

$\boxed{V_2 = 2.79L}$

2. Suppose you have a 12.2 L sample containing 0.50 mol of oxygen gas (O_2) at a pressure of 1 atm and a temperature of 25°C. If all of this O_2 were converted to ozone (O_3) at the same temperature and pressure, what would be he volume, in liters, of the ozone?

$3O_{2\,(g)} \longrightarrow 2O_{3\,(g)}$

12.2L ?L

.50 mol

1 Atm 1Atm

25°C = 298K 298K

$\left(\frac{.50 \text{ mole } O_2}{1}\right)\left(\frac{2 \text{ mole } O_3}{3 \text{ mole } O_2}\right)$

$= .33 \text{ mole } O_3$

$PV = nRT$

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

$\frac{V_1}{n_1} = \frac{V_2}{n_2}$

$V_2 = \frac{V_1 n_2}{n_1}$

$= \frac{(12.2L)(.33 \text{ mole } O_3)}{.50 \text{ mole } O_2}$

$\boxed{V_2 = 8.1 \text{ L } O_2}$

3. A sample of diborane gas (B_2H_6), a substance that bursts into flame when exposed to air, has a pressure of 345 torr at a temperature of -15°C and a volume of 3.48 L. If the conditions are changed so that the temperature is 36°C and the pressure is 468 torr, what will the volume (in L) of the sample be?

Given: B_2H_6

$P_1 = 345$ Torr $P_2 = 468$ Torr

$T_1 = -15°C = 258K$ $T_2 = 36°C = 309K$

$V_1 = 3.48L$ $V_2 = ?$

Soln.

$PV = nRT$

$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

$V_2 = \frac{P_1V_1}{T_1}\frac{T_2}{P_2}$

$= \frac{(345 \text{ Torr})(3.48L)(309K)}{(258K)(468 \text{ Torr})}$

$\boxed{V_2 = 3.07L}$

4. A sample of methane gas, CH_4, having a volume of 2.80 L at 25°C and 1.68 atm was mixed with a sample of oxygen gas having a volume of 35.0 L at 31°C and 1.25 atm. The mixture was then ignited to form carbon dioxide and water. Calculate the volume of CO_2 formed (in L) at a pressure of 2.50 atm and a temperature of 125°C.

$$CH_{4(g)} + 2O_{2(g)} \longrightarrow CO_{2(g)} + 2H_2O_{(g)}$$

CH_4	O_2	CO_2
2.80L	35.0L	? L
25°C	31°C	125°C
1.68 Atm	1.25 Atm	2.50 Atm

① CH_4

$PV = nRT$

$n = \frac{PV}{RT} = \frac{(1.68\,Atm)(2.80L)}{(0.08206)(298K)}$

$n = .192$ mole CH_4

O_2

$n = \frac{(1.25\,Atm)(35.0L)}{(0.08206)(304K)} = 1.75$ mole O_2

②

	CH_4	$2O_2$	→	CO_2
	.192	1.75		0
	$-1X$	$-2X$		X
	$-.192$			.192
	0			.192 moles

$X = .192$

③ CO_2

$PV = nRT$

$V = \frac{nRT}{P}$ (398K)

$= \frac{(.192\,moles)(.08206)}{2.50\,Atm}$

$\boxed{V_{CO_2} = 2.51L}$

5. The density of a gas was measured at 1.50 atm and 27°C and found to be 1.95 g/L. Calculate the molar mass of the gas.

$D = 1.95\,g/L$

$P = 1.50\,Atm$

$T = 27 = 300.\,K$

$MM = \frac{DRT}{P} = \frac{(1.95\,g/L)(0.08206)(300.K)}{1.50\,Atm}$

$\boxed{MM = 32.0\,g/mole}$

6. Mixtures of helium and oxygen are used in scuba diving tanks to help prevent the "bends". For a particular dive, 46.0 L O_2 at 25°C and 1.0 atm and 12.0 L of He at 25°C and 1.0 atm were pumped into a tank with a volume of 5.0 L. Calculate the partial pressure of each gas (in atm) and the total pressure (in atm) in the tank at 25°C.

Given.

$V_{Final} = 5.0L$ $P_{Partial} = ?$ $P_T = ?$ @ 25°C

O_2

$V = 46.0L$

$T = 25°C = 298K$

$P = 1.0\,Atm$

He

$V = 12.0L$

$T = 25°C = 298K$

$P = 1.0\,Atm$

Soln.

$PV = nRT$

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

$n_{O_2} = \frac{(1.0\,Atm)(46.0L)}{(.08206)(298K)}$

$n_{O_2} = 1.9$[redacted] moles

$n_{He} = \frac{(1.0\,Atm)(12.0L)}{(.08206)(298K)}$

$n_{He} = .49$[redacted] moles

Now $V_{final} = 5.0L$

$PV = nRT \Rightarrow P = \frac{nRT}{V}$

$P_{O_2} = \frac{(1.9\,mole)(.08206)(298K)}{5.0L}$

$\boxed{P_{O_2} = 9.3\,Atm}$

$P_{He} = \frac{(.49\,mole)(.08206)(298K)}{5.0L}$

$\boxed{P_{He} = 2.4\,Atm}$

$P_T = P_{He} + P_{O_2} = 2.4\,Atm + 9.3\,Atm$

$\boxed{P_T = 11.7\,Atm}$

7. A sample of solid potassium chlorate ($KClO_3$) was heated in a test tube and decomposed to form solid potassium chloride and oxygen gas. The oxygen was collected by the displacement of water at 22°C at a total pressure of 754 torr. The volume of gas collected was 0.650 L, and the vapor pressure of water at 22°C is 21 torr. Calculate the mass of $KClO_3$ sample, in grams, that was decomposed.

$$2KClO_{3(s)} \xrightarrow{\Delta} 2KCl + 3O_{2(g)}$$

? g

T = 22°C = 295 K
P = 754 Torr
V = .650 L

@22 Vp = 21 Torr

$$P_{O_2} = P_T - P_{vapor}$$
$$= 754 - 21$$
$$= 733 \text{ Torr}$$

PV = nRT

$$n = \frac{PV}{RT} = \frac{(733 \text{ Torr})(.650 L)}{(62.36)(295 K)}$$

n = .0259 mole O_2

$$\left(\frac{.0259 \text{ mole } O_2}{1}\right)\left(\frac{2 KClO_3}{3 \text{ mole } O_2}\right)\left(\frac{122.6 g\ KClO_3}{1 \text{ mole } KClO_3}\right)$$

$$= \boxed{2.12 g\ KClO_3}$$

8. A student tries to determine the volume of a glass bulb. These are her results:

mass of bulb filled with dry air	91.6843 g
mass of evacuated bulb	91.4715 g
temperature of air	23°C
pressure of air inside of bulb	744 mmHg

= 0.2128 g dry Air

Assume the composition of air is 78% N_2, 21% O_2, and 1.0 % Ar. What is the volume (in mL) of the bulb. (Hint: First calculate the molar mass of air!)

$$MM_{air} = (.78)(28.02 g/mole) + (.21)(32.00 g/mole\ O_2) + (.001)(39.95 g/mole\ Ar)$$
$$= 29 \text{ g/mole}$$

$$\text{Moles Air} = \frac{\text{mass}}{MM} = \frac{0.2128 g \text{ dry Air}}{29 \text{ g/mole}} = 0.0073 \text{ moles Air}$$

$$P = (744 mm)\left(\frac{1 Atm}{760 mm}\right) = .979 \text{ Atm}$$

V = ? ml

T = 23°C = 296 K

n = 0.0073 moles

PV = nRT

$$V = \frac{nRT}{P} = \frac{(0.0073 \text{ moles})(0.08206)(296 K)}{.979 \text{ Atm}}$$

V = .181 L

$$\boxed{V = 181 \text{ ml}}$$

9. Answer each of the following:

a) Under the same conditions of temperature and pressure, which of the following gases would behave most ideally? Explain your choice.

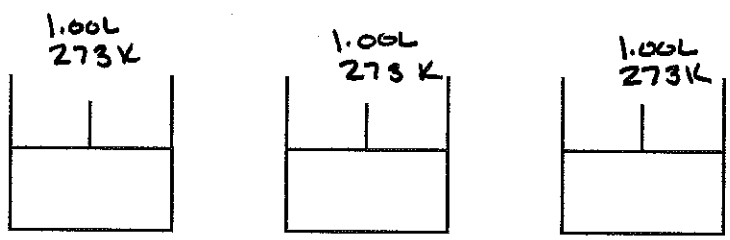
Ne, N_2, or CH_4

Ne would Behave more ideally Low MM + Size, less Attractive forces

b) Would you expect a real gas to have a higher or lower pressure than an ideal gas under the same conditions? Why?

A real gas would have a high pressure under the same conditions as an ideal gas due to the Ideal gas makes the Assumption that an Ideal gas molecules have No volume or Intremolecular Forces

10. Consider three pistons each containing 1.00 g of the gas specified in 1.00 liters measures at 273 K. The pressure is not specified.



Temp = 273° K

1.00g He 1.00 g Xe 1.00 g Kr

$\left(\frac{1.00g\ He}{1}\right)\left(\frac{1\ mole\ He}{4.00g\ He}\right) = .25\ mole$ $\left(\frac{1.00g\ Xe}{1}\right)\left(\frac{1\ mole\ Xe}{131.29}\right) = .00762$ $\left(\frac{1.00g\ Kr}{1}\right)\left(\frac{1\ mole\ Kr}{83.80g}\right) = .0119$

Comment on the following:

A. The pressure of the gases in each piston.

He will have the highest pressure due to the largest (.250 mole) # of moles. Kr will have 2nd highest P + Xe lowest

B. The average velocity of the gases in each piston.

He molecules will have the greater avg speed, Because He has smallest molar Mass $V = \sqrt{\frac{mm_{He}}{mm_x}}$

C. The density of the gases in each piston.

$D = \frac{mass}{V}$ ∴ all pistons have same density

D. The number of gas molecules in each piston.

He will have most molecules # mole $\times \frac{6.02\times10^{23}\ molecules}{1\ mole}$

∴ Largest mole will greatest # of gas molecules

E. The average kinetic energy of the gases in each piston.

The Avg KE is the same Because the same temp for each piston + KE is Depedant upon temp

11. Use the following information to answer the questions below. A student measured the mass of a sealed 700. mL glass flask that contained dry air. The student then flushed the flask with the unknown gas, resealed it, and measured the mass again. Both the air and the unknown gas were at 22.0°C and 740.0 mm . The data for the experiment are shown in the table below.

A. Volume of sealed flask	700.0 mL
B. Mass of sealed flask and dry air	140. 0g
C. Mass of sealed flask and unknown gas	141.6 g

V = 700.0 mL = .7000 L

$P = 740.0\ mm\left(\frac{1\ Atm}{760\ mm}\right) = .974\ Atm$

T = 22.0°C = 295.0 K

Find the mass, in grams, of the following:

A. Of the dry air that was in the sealed flask. (The density of dry air is 1.20 g/L at 22.0°C and 1.00 atm.)

mass dry Air ?g

$D = \frac{m}{V}$

$m = DV$

$\left(\frac{1.20 g}{L}\right)\left(\frac{.7000 L}{1}\right) = .840 g$ dry Air

B. Of the sealed flask itself with no air in it.

140.0 g Flask + Air − .840 g Air = 139.2 Flask Empty

C. Of the unknown gas that was added to the sealed flask.

Mass unknown Gas = Flask − Empty Flask

141.6 g − 139.2 g = 2.4 g unknown GAS

D. Knowing all the info above find the molar mass of the unknown gas.

$MM = \frac{DRT}{P}$ $D = \frac{m}{V}$ $MM = \frac{mass\ RT}{PV} = \frac{(2.4 g)(0.08206)(295 K)}{(.974\ Atm)(.7000 L)} = 85.2\ g/mole$

E. If the gas is Cl_2, then what is the percent error?

Cl_2 = 70.9 g/mole

$\left|\frac{85.2 - 70.9}{70.9}\right| \times 100 = 1.69\ \%$

F. What error would cause an increase in MM of the gas? Weight of collected Gas to high or Record temp to high,

G. What error would cause an decrease in MM of the gas? Not Subtract Air in Part B