
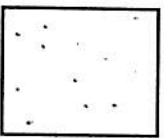
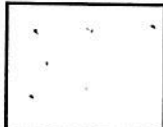


GASES AP Chemistry

1. Properties of the states of matter:

	SOLID	LIQUID	GAS
			
Shape/Volume:	definite shape and volume	definite volume, no definite shape	no definite shape or volume
Compressibility:	only slightly compressible	only slightly compressible	highly compressible <i>syringes</i>
Density:	med/high density	med/high density	low density
Movement, speed, distance:	close together, only vibrational movement	particles can slide past each other, still limited movement	particles are FAR apart from each other

2. What factors decide what phase - liquid, gas or solid - something will be?

a. Temperature - is a measure of the average kinetic energy in a substance.

1. Kinetic Energy - dependent on mass and velocity ($KE = 1/2 mv^2$)

* The greater the temperature, the greater the KE (the faster particles are moving)

* Any two gases at the same temperature will have the same ave. KE.

** IF temp was the ONLY factor, EVERYTHING would be in the same phase at room temp....

b. physical phases - are determined by the strength of attractive forces.

1. Solids = particles are held together by Strong attractive forces

2. Liquids = particles are held together by weaker attractive forces

3. Gases = particles are held together by weakest attractive forces

3. Boiling points are an indication of the strength of attractive forces.

- The higher the boiling point, the greater the attractive forces.

Q: Why are "salts" solids at room temp? ionic bonds (lattice energy) atoms are ^{very} close!

Q: Why are noble gases gases monatomic? Don't need to bond, low mass, no ^{no polar or non-polar} dipoles, ^{no} charges, etc.

1. Measuring Gases:

a. Amount of Gas (n): measured in moles

b. Volume (V): Measured in Cubic linear measurements ($1L = 1000mL$ or $10^3 \text{ centimeters}^3$)

c. Temperature (T): measure in Kelvin $T(K) = 273.15 + \text{temp } ^\circ C = T(K)$

d. Pressure (P): measured in atmospheres

P, V, n, T.

2. PRESSURE

a. Define as Force per unit area.

****b.** is created as a result of COLLISIONS of particles/ unit of area.

- Gas particles are in constant random motion. They exert pressure on any surface with which they come in contact with.

3. ATMOSPHERIC PRESSURE

a. pressure exerted by Earth's atmosphere.

b. "The air is thinner in Denver." Why? What does this mean?

c. Use a barometer to measure atmospheric pressure.

*Standard atmospheric pressure = 760 mmHg = 1 atm = 760 torr = 101,325 Pa =

101.325 kPa = 14.71 lb/in² = 1.01325 bar = 1013.25 millibar (one pascal = 1.0 N/m²)

***** Be sure you can do conversions from one pressure unit to another!!!!**

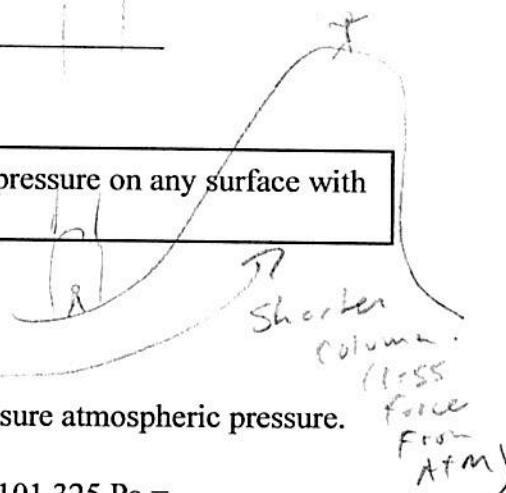
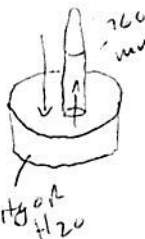
d. DEMO's.....illustrations of atmospheric pressure!

4. MANOMETERS

a. A device used to measure the pressures of gases other than the atmospheric.

1. Which vessel contains gas w/P > than atmospheric? _____
2. Which vessel contains gas w/P < than atmospheric? _____
3. What would happen in B if the apparatus were carried to the top of a high mountain?

4. Calculate the P_{gas} in C? Assume P_{atm} = 760



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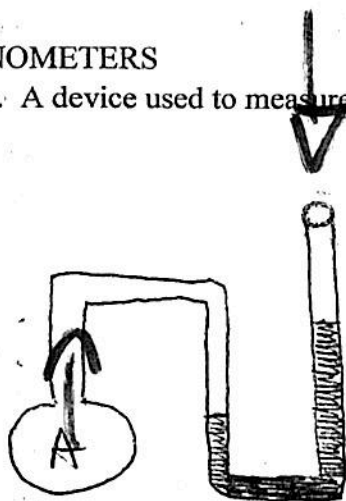
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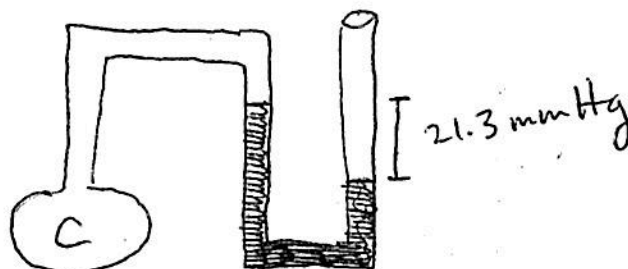
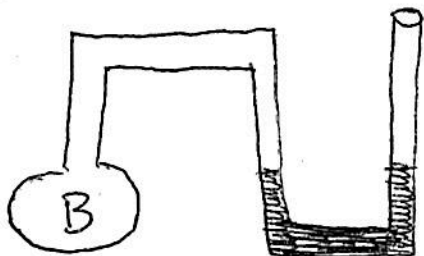
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NOTES #36 THE GAS LAWS:

Draw Experimental apparatus here↓:

I. BOYLE'S LAW:

The Pressure - Volume Relationship

Show piston on demo syringe

- a. The volume of a given amount of gas at constant temperature is inversely proportional to the pressure applied to the gas. In other words, as pressure INCREASES, volume decreases. $V_1 = \frac{k}{P_1}$
- b. Mathematically, we can write this law two ways: $V \propto \frac{1}{P}$ OR $V = \frac{k}{P}$. $V_1 P_1 = V_2 P_2$
- c. Rearranging yields a mathematical statement of Boyle's Law. $VP = k$ (a constant) ✓
The product of $P \times V$ is a constant for a fixed amount of gas at constant temperature. Take a look...

Trial	#1	#2	#3	#4	#5	#6
P (mmHg)	724	869	951	998	1230	1893
V (liters)	1.50	1.33	1.22	1.16	0.94	0.61
P x V	1090	1160	1160	1200	1200	1100

- d. For changes in pressure from P_1 to P_2 , we can apply Boyle's Law. Since the product $P \times V$ is a constant, its value at the initial pressure, P_1 , and its volume, V_1 , will EQUAL $P \times V$ at the final pressure, P_2 , and the final volume, V_2 .

$P_1 V_1 = P_2 V_2$ when moles (n) & Temp (K) are held constant. $\frac{P_1 V_1}{n T_1} = \frac{P_2 V_2}{n T_2}$

II. CHARLES'S LAW: The Temperature - Volume Relationship

$V \propto T$ OR $V = kT$

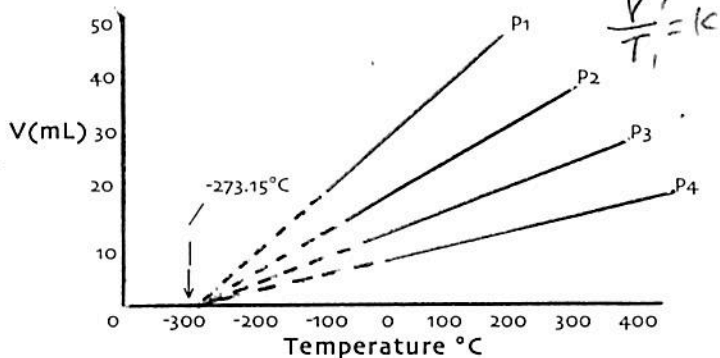
- a. At constant pressure and moles, the volume of a gas is directly proportional to its absolute temperature in units of Kelvin. $\frac{V_1}{T_1} = \frac{V_2}{T_2} = k$

- b. Mathematically, we can express this as:

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

- c. Rearranging, we get another form of Charles's Law,

$T_2 V_1 = T_1 V_2$. Since V/T is equal to a constant, then V_1/T_1 at an initial temperature, T_1 , and volume, V_1 , will be equal to V_2/T_2 at a final temperature, T_2 , and volume, V_2 .



- d. Charles's Law allowed for the prediction of ABSOLUTE ZERO, the temperature at which ALL Molecular Motion stops. Look above! ↑

@ OK $KE = 0$

III. AVOGADRO'S LAW: The Volume - Amount Relationship

$V \propto n$ OR $V = kn$

- a. At the same temperature and pressure, equal volumes of ANY gas contain the same number of molecules (or moles). In other words, the volume of gas is proportional to the number of moles present. $\frac{V_1}{n_1} = \frac{V_2}{n_2}$ OR $V_1 n_2 = V_2 n_1$

- b. It's been found that 1 mole of ANY gas at STP (standard temperature and pressure) will occupy a volume of 22.41 L. This is a very useful conversion. If you know the volume of a gas (x), after correcting for the temperature and pressure (STP), the moles can be determined: ↓

- c. STP = Standard Temperature and Pressure: 1.00 atm, $\frac{x \text{ liters}}{22.41 \text{ L/mol}} = \text{moles}$
273.15 K or 0 °C

IV. GAY-LUSSAC'S LAW: The Temperature - Pressure Relationship

$$T \propto P \quad T = kP$$

$$\frac{T_1}{P_1} = k \quad \frac{T_1}{P_1} = \frac{T_2}{P_2}$$

a. As the absolute temperature of a gas is increased its pressure will increase in a manner that is directly proportional to the temperature. This is true as long as the volume and amount of moles are held constant.

$$T_1 P_2 = T_2 P_1$$

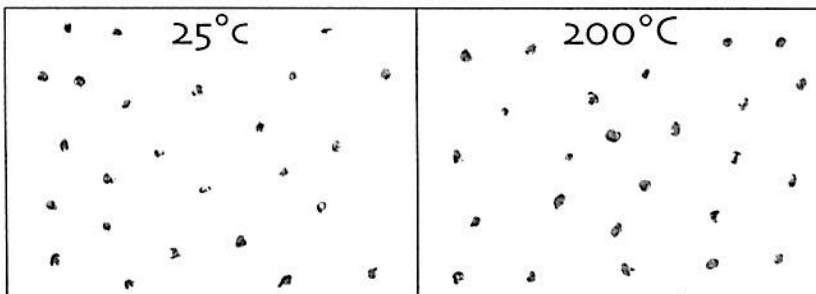
b. Upon heating, the kinetic energy of the surroundings is transferred the gas in the container. Remember, $KE(ave) \propto \text{Temperature}$, so if $T \uparrow$, $KE \uparrow$. If $KE \uparrow$, then velocity \uparrow , due to $KE = \frac{1}{2}mv^2$. As the molecules travel with greater velocities, they will exert greater collision forces on the inner surface of the container. Since $\rightarrow \text{Pressure} = \text{Force}/\text{Area}$, $T \uparrow \propto P \uparrow$.

Gas Laws - a molecular level.

GAY LUSSAC

Scenario #1 Temperature - Pressure Relationships

[Must keep N & V constant]



A sample of gas was heated from 25°C to 200°C.

1. What will happen to the pressure exerted by this gas?

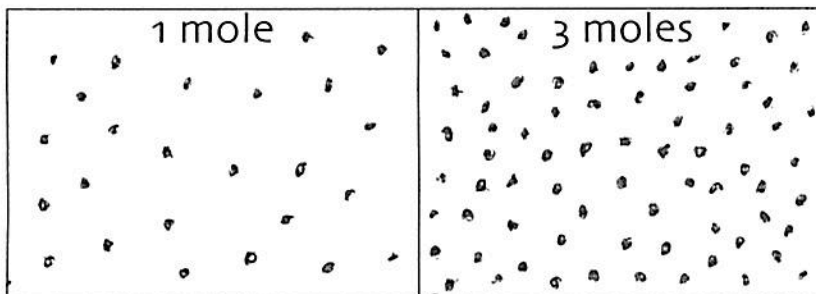
increase

2. Explain how each of the following factors does or does not contribute to the pressure change that occurs when the temperature increases.

- a) Collisions per unit time $\uparrow P$ (increase) $\uparrow P$
 b) Energy per collision increase $\uparrow P$ $m \times v = \text{momentum}$
 c) Number of particles per unit volume. SAME

Scenario #2 Pressure & # of Particles

[Must keep V & T constant]



A container initially holds 1 mole of gas and then 2 additional moles are added.

1. What will happen to the pressure exerted by this gas?

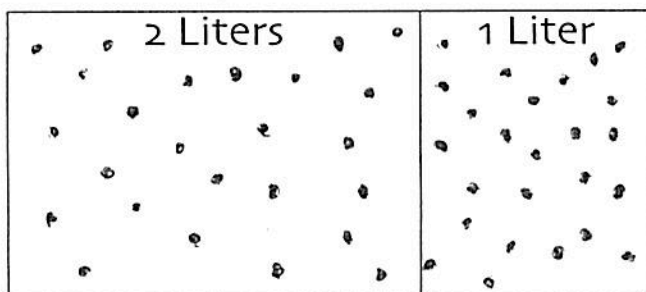
increase

2. Explain how each of the following factors does or does not contribute to the pressure change that occurs when the temperature increases.

- a) Collisions per unit time $\uparrow P$
 b) Energy per collision stays the same (no ΔKE)
 c) Number of particles per unit volume. $\uparrow P$, \uparrow density (more)

Scenario #3 Pressure - Volume Relationships

[Must keep N & T constant]



A sample of gas was initially in a 2 liter container and then the volume was reduced to 1 liter.

1. What will happen to the pressure exerted by this gas?

increase

2. Explain how each of the following factors does or does not contribute to the pressure change that occurs when the temperature increases.

- a) Collisions per unit time $\uparrow P$
 b) Energy per collision Same Temp $\propto KE$ Same temp and speed, but less distance
 c) Number of particles per unit volume. increases, $\uparrow P$

THE IDEAL GAS EQUATION, R, AND APPLICATIONS

Historically, the research of three scientists led to the relationships between four variables, P, V, T and n. By holding two variables constant these scientists displayed how the remaining two variables affect each other.

BOYLE'S LAW:	$V \propto \frac{1}{P}$	(constant n, T)
CHARLE'S LAW:	$V \propto T$	(constant n, P)
AVOGADRO'S LAW:	$V \propto n$	(constant P, T)

Gay-Lussac
 $P \propto T$

We can combine these relationships to make a more general gas law:

$$V \propto nT/P$$

R, the proportionality constant
it relates these units together:

$$V = R(nT/P)$$

The values for R are
0.0821 L-atm/mol-K or
8.314 J/mol-K

$$R = \frac{PV}{nT} = \frac{1.00 \text{ atm} \cdot 22.4 \text{ L}}{1.00 \text{ mol} \cdot 273 \text{ K}}$$

Rearranging the equation leads us to the familiar ideal gas equation:

$$PV = nRT$$

or
 $8.314 \text{ kg m}^2 / \text{s}^2 / \text{mol} \cdot \text{K}$

$$0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{10^{-3} \text{ m}^3}{2} \times \frac{101325 \text{ Pa}}{1 \text{ atm}} \times \frac{1 \text{ N/m}^2}{1 \text{ Pa}}$$

$$= 8.314 \frac{\text{N} \cdot \text{m}}{\text{mol} \cdot \text{K}} = 8.314 \frac{\text{J}}{\text{mol} \cdot \text{K}}$$

APPLICATIONS:

AS DENSITY AND MOLAR MASS.

Gas density has the units of mass per unit volume. Rearranging the ideal gas equation leads to a new equation revealing mol/vol.

$$n/V = P/RT \quad (\text{units are mol/L})$$

We can multiply both sides of the equation by M, molar mass (the number of grams in 1 mol of substance).

$$nM/V = PM/RT$$

What are the units for the left side of the equation? $\frac{\text{mol} \cdot \text{g}}{\text{mol}} \cdot \frac{1}{\text{L}} = \frac{\text{g}}{\text{L}}$ That's density!

$$\frac{8.314 \text{ kg m}^2}{\text{mol K s}^2}$$

A new equation could be:

$$d = PM/RT$$

- To find the molar mass if the density is known, simply rearrange the above equation.

$$M = dRT/P$$

Ex 1: What is the density of carbon tetrachloride vapor at 714 torr and 125°C?

$$d = \frac{(714 \text{ torr})(1 \text{ atm} / 760 \text{ torr})(153.8 \text{ g/mol})}{0.0821 \text{ L atm/mol K} (398 \text{ K})} = \boxed{4.42 \frac{\text{g}}{\text{L}}}$$

Ex 2: Calculate the average molar mass of dry air if it has a density of 1.17 g/L at 21°C and 740.0 torr.

$$M = \frac{(1.17 \text{ g/L})(0.0821 \text{ L atm/mol K})(294 \text{ K})}{(740 \text{ torr} \times 1 \text{ atm} / 760 \text{ torr})} = \boxed{29.0 \text{ g/mol}}$$

$$\frac{1}{2} m u^2 = \frac{3}{2} R T / \text{mol}$$

$$u^2 = \frac{\frac{3}{2} R T}{\text{mol}(\frac{1}{2}) \text{ mass}}$$

$$u^2 = \frac{\frac{3}{2} R T \times 2}{\text{mol mass}} = \frac{3 R T}{m}$$

$$u = \sqrt{\frac{3 R T}{m}}$$

Applications cont.

GAS MIXTURES, PARTIAL PRESSURES and MOLE FRACTIONS

Gas mixtures and partial pressures:

John Dalton (remember him from atomic theory?) observed that *the total pressure exerted by a mixture of gases is equal to the sum of the pressures that each gas would exert if it were present alone.*

This is referred to as **Dalton's Law of Partial Pressures**:

$$P_t = P_1 + P_2 + P_3 + \dots$$

Each of these gases is subject to the gas laws and behaves independently.

So, $P_1 = n_1(RT/V)$. This applies to all of the gases in a mixture. ... $P_t = (n_1 + n_2 + n_3 + \dots)RT/V$ or $= n_t(RT/V)$

▪ Mole Fractions:

It is then reasonable to assume that a gas behaves as a portion of the whole based on its presence in moles.

$$\frac{P_1}{P_t} = \frac{n_1 RT/V}{n_t RT/V} = \frac{n_1}{n_t}$$

n_1/n_t is called the *mole fraction* of the gas. The symbol X_1 will be used to designate this. The mole fraction is a dimensionless number used to represent a percentage as a decimal.

$$P_1 = (n_1/n_t)P_t = X_1 P_t$$

Ex 3: Air is 78% nitrogen gas (the mole fraction of nitrogen is 0.78). If the pressure is 1.00 atm, what is the partial pressure of nitrogen in torr?

$$P_{N_2} = X_{N_2} P_T$$

$$P_{N_2} = 0.78(1.00 \text{ atm}) \left(\frac{760 \text{ torr}}{1.00 \text{ atm}} \right) = \boxed{593 \text{ torr}}$$

Ex 4: When mountain climbing, the thin air soon reveals a lack of oxygen. What is the partial pressure of oxygen at the summit of Mount Everest (mm of Hg) when the total pressure is 338.6 millibars. Assume the oxygen content is 21%.

$$P_{O_2} = X_{O_2}(P_T) = (0.21)(338.6 \text{ mb}) \frac{760 \text{ mm Hg}}{1013.25 \text{ mb}} = \boxed{53.33 \text{ mm Hg}}$$

Ex 5: What is the partial pressure of oxygen (mm of Hg) at STP? Still assume 21% oxygen.

$$P_{O_2} = X_{O_2}(P_T) = (0.21)(1.00 \text{ atm}) \frac{760 \text{ mm Hg}}{1.00 \text{ atm}} = \boxed{160 \text{ mm Hg}}$$

DIAGRAM FOR COLLECTING GASSES OVER WATER:

$$\begin{aligned}
 \rho &= \frac{0.062 \text{ atm} \cdot \text{L}}{\text{mol} \cdot \text{K}} \times \frac{1 \times 10^{-3} \text{ m}^3}{1 \text{ L}} \times \frac{101325 \text{ N/m}^2}{\text{atm}} = \frac{8.314 \text{ N/m}}{\text{mol} \cdot \text{K}} \\
 &\quad \downarrow \text{ Pascals} \qquad \qquad \qquad \downarrow \text{ J}
 \end{aligned}$$

$$\text{So } \rho = 8.314 \frac{\text{J}}{\text{mol} \cdot \text{K}}$$

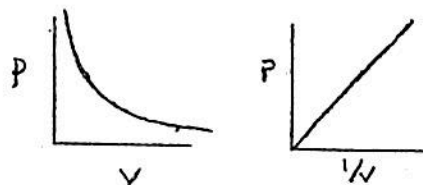
$$\left. \begin{array}{l} 22.414 \\ 1.00 (1.013) \\ 1.00 (273) \end{array} \right\} = \frac{0.062 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}$$

GAS LAWS SUMMARY

Ap Chemistry

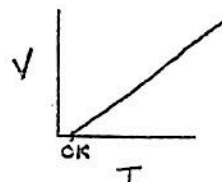
1. BOYLE'S LAW

- $P_1V_1 = P_2V_2$ * when n and T are constant
- As P , V (indirectly proportional)



2. CHARLES'S LAW

- $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ * when n and P are constant
- As T , V (directly proportional)
- Temperature ALWAYS has to be in Kelvin!! $T(K) = T(^{\circ}C) + 273$



3. AVOGADRO'S LAW

- 1 mole of *any* gas at STP takes up a volume of 22.4 L. $n \propto V$ (when $T + P$ are constant)
- STP = standard temperature and pressure = 1 atm and 273 K or $0^{\circ}C$

4. IDEAL GAS LAW EQUATION (combines Boyle's and Charles's law)

- $PV = nRT$
- Temperature has to be in Kelvin
- $R = 0.0821 \text{ atm}\cdot\text{L}/\text{K}\cdot\text{mol}$

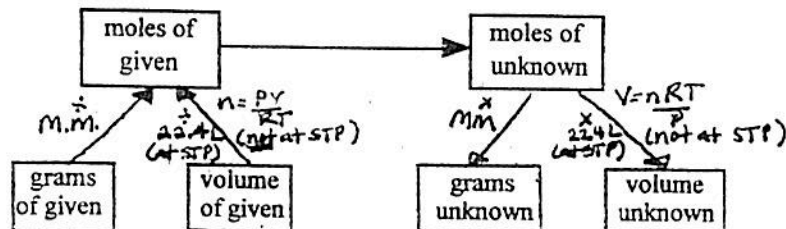
5. COMBINED GAS LAW

- $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$ * when n is constant
- Temperature has to be in Kelvin

6. GAS LAW EXTENSIONS (Incorporating molar mass and density of a gas)

- $\text{Density} = \frac{m}{V} = \frac{PM}{RT}$
- M = molar mass of gas (grams/mol), m = mass of gas (grams)
- Also remember, $D_{\text{of gas}} = \frac{\text{mass of gas}}{\text{Volume of gas}}$ and MM (molar mass of gas) = $\frac{\text{mass of gas}}{\text{moles of gas}}$

7. GAS STOICH



8. DALTON'S LAW OF PARTIAL PRESSURE

- $P_{\text{Total}} = P_A + P_B + P_C + \dots$ and $\text{Mole Fraction of gas A} = X_A = \frac{n_A}{n_{\text{total}}} = \frac{P_A}{P_{\text{total}}}$

$$\frac{P_1 V_1}{n R T_1} = \frac{P_2 V_2}{n R T_2}$$

Gas Law Calculations

Use the combined gas law

1. A 5.00 L container is filled with $N_2(g)$ to a pressure of 3.00 atm at 250 °C. What would be the resultant volume of a container that is used to store the same gas at STP?

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \quad \frac{3.00 \text{ atm}(5.00 \text{ L})}{523 \text{ K}} = \frac{1.00 \text{ atm}(V_2)}{273 \text{ K}}$$

$$\frac{P_1 V_1 T_2}{P_2 T_1} = V_2$$

$$V_2 = \boxed{7.83 \text{ Liters}}$$

2. Calculate the volume (in liters) occupied by 7.40 g of CO_2 @ STP. Use Avogadro's Law

$$7.40 \text{ g } CO_2 \times \frac{1 \text{ mol } CO_2}{44.01 \text{ g } CO_2} = 0.168 \text{ mol } CO_2 \times \frac{22.4 \text{ L}}{1 \text{ mol gas (at STP)}} = \boxed{3.77 \text{ L}}$$

$$\text{OR: } PV = nRT \\ v = \frac{nRT}{P}$$

3. **Molar Mass/Density Calculations!** Cyclopropane is used as a general anesthetic. It has a molar mass of 42.0 grams. What is the density of cyclopropane gas at 25.0 °C and 1.02 atm?

$$MM = 42.0 \text{ g/mol}$$

$$d = ?$$

$$T = 25^\circ \text{C} \rightarrow 298 \text{ K}$$

$$P = 1.02 \text{ atm}$$

$$n = 1 \text{ mole}$$

$$V = ?$$

$$PV = nRT$$

$$\times$$

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

$$\text{OR } d = \frac{PM}{RT}$$

$$V = \frac{(1 \text{ mol})(0.0821 \text{ L} \cdot \text{atm}) (298 \text{ K})}{(1 \text{ mol})(1.02 \text{ atm})}$$

$$= 23.99 \text{ L}$$

$$\text{gas } d = \frac{\text{mass}}{V} = \frac{42.0 \text{ g}}{23.99 \text{ L}}$$

$$\boxed{d = 1.75 \text{ g/L}}$$

4. A compound contains 11.79% C, 69.57% Cl, and 18.64% F

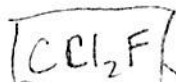
A) Find the empirical formula

B) If 0.107 g of the compound fills a 458 mL flask at 25 °C with a pressure of 21.33 mmHg, what is the molecular formula?

$$11.79\% \text{ C}$$

$$69.57\% \text{ Cl}$$

$$18.64\% \text{ F}$$



if emp. formula is correct

$$MM = 101.9 \text{ g/mol}$$

$$\frac{11.79}{12.01} = \frac{.9817}{.9811} \approx \frac{1.0 \text{ mol}}{C}$$

$$\frac{69.57}{35.45} = \frac{1.962}{.9811} \approx \frac{2.0 \text{ mol}}{Cl}$$

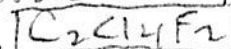
$$\frac{18.64}{19.00} = \frac{.9811}{.9811} = \frac{1.0 \text{ mol}}{F}$$

$$\text{B) } PV = nRT \quad n = \frac{PV}{RT}$$

$$n = \frac{(0.028 \text{ atm})(.458 \text{ L})}{0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} (298 \text{ K})} = 5.25 \times 10^{-4} \text{ mol}$$

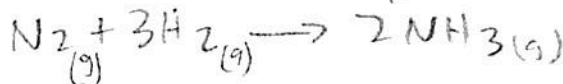
$$MM = \frac{.107 \text{ g}}{5.25 \times 10^{-4} \text{ mol}} = 204 \text{ g/mol}$$

Double the emp



5. What volume of N_2 gas at 720 torr and 23 °C is required to react with 7.35 L of H_2 gas at the same temperature and pressure?

trick question



$$7.35 \text{ L } H_2 \times \frac{1 \text{ L } N_2}{3 \text{ L } H_2} = 2.45 \text{ L } N_2$$

Kinetic Molecular Theory

Ap Chemistry

I. Kinetic Molecular Theory of Gases:

a. What are the 4 assumptions made by this theory?

1. Volume: gas particles have negligible volume, b/c gases are separate by some distance.
2. Motion: constant random motion - collisions are elastic
3. Attractive forces: no attractive/repulsive forces.
- * 4. Kinetic Energy: The average KE of gas particles is proportional to the absolute temperature.

- ANY TWO GASES AT THE SAME TEMPERATURE WILL HAVE THE SAME AVE. KE.

b. Be able to explain each of the gas laws in terms of Kinetic Molecular Theory (refer to pages 181-182).

Ex:

Boyle's Law
 $P \propto \frac{1}{V}$
 As P ↑ molecules push back more force
 Δ in V rec'd
 1896

II. Kinetic Energy: Let's look at equation for KE.

$$KE = \frac{1}{2} mv^2 \dots \text{or (more accurately)} \dots \overline{KE} = \frac{1}{2} m \overline{u^2}$$

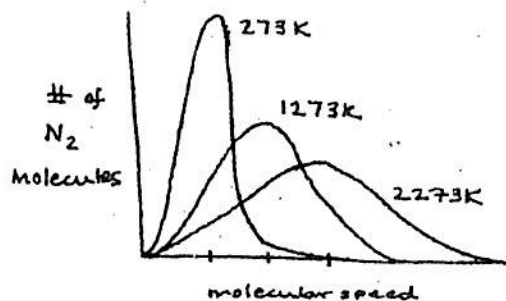
a. $\overline{u^2}$ = mean squared speed - average of the square of the speeds of all the molecules.

** b. KE depends on both the mass and the average speed of particles.

1. Just because the average KE and speed of each molecule is the same at constant T, at any one instant, are all the molecules moving at that speed? Why or why not??

NO, because after any one collision - molecules may speed up or slow down (even if all the same gas)

- We can use a Maxwell speed distribution curve to show the # of gas molecules moving at certain speeds.



** As T increases, molecular speed increases.

** As T increases, molecular speed becomes more variable. Why???

different collisions are possible @ higher speeds.

2. Let's say we are talking about TWO gases at the same temperature. Because they are at the same T, they should have the same ave. K. energy but do their molecules travel at the same average speed?? Why or why not???

NO, KE depends on both molar mass & speed. A heavier molecule is going to move slower (if masses are different)

** Thinking back to the diffusion demo, which gas, NH_3 or HCl had molecules moving faster??? NH_3

III. Determining the speed of gas particles: We can quantitatively estimate the speed of particles by calculating an average molecular speed or a root-mean-square (rms) speed (u_{rms}).

a. How do we derive u_{rms} ?? Equalize and manipulate the following two equations: (refer to pg 183-184)

$$KE \text{ per molecule} = \frac{1}{2} m u^2$$

$$KE \text{ per mole} = \frac{3}{2} RT$$

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

M = molar mass (kg/mol)
 must remember
 M is in kg/mol

$$R = 8.314 \frac{J}{mol \cdot K} = \frac{kg \cdot m^2}{s^2 \cdot mol \cdot K}$$

$1 \text{ J} = 1 \text{ N m}$
 $\text{OR } 1 \text{ J} = \frac{\text{kg m}^2}{\text{s}^2}$
 $1 \text{ ATM} = 101325 \frac{\text{N}}{\text{m}^2} (\text{Pa})$
 $R = 0.0821 \frac{\text{atm L}}{\text{mol K}} \times \frac{1 \times 10^{-3} \text{ m}^3}{1 \text{ L}} \times \frac{101325 \text{ N/m}^2}{1 \text{ atm}} = 8.314 \frac{\text{J}}{\text{mol K}}$
 $= 8.314 \frac{\text{J}}{\text{mol K}}$

b. You want u_{rms} in m/s. In order to do this R has to be in units of J/mol·K and molar mass has to be in units of kg/mol.

$R = 8.314 \text{ J/mol·K}$ (derivation in Appendix 2 of book)

$1 \text{ J} = \frac{1 \text{ kg m}^2}{\text{s}^2}$

ex. Compare quantitatively the root-mean-square speeds (in m/s) of NH_3 and HCl gas molecules at 25 C.

$\text{NH}_3: u_{\text{rms}} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3(8.314 \frac{\text{J}}{\text{mol K}})(298 \text{ K})}{0.017 \text{ kg/mol}}} = 661 \approx 660 \frac{\text{m}}{\text{s}}$
 $\text{HCl: } u_{\text{rms}} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3(8.314 \frac{\text{J}}{\text{mol K}})(298 \text{ K})}{0.03646 \text{ kg/mol}}} \approx 450 \frac{\text{m}}{\text{s}}$

IV. Gas Diffusion: Refers to the gradual mixing of gas molecules. Diffusion always proceeds from a region of high concentration to a region of low concentration. Looking at the molecular speeds from above, one would expect molecules to mix quickly. However, this is NOT the case....why?

constant random motion & interference w/ all molecules

a. The relative rates of diffusion can be calculated by comparing u_{rms} values:

$\frac{r_1}{r_2} = \frac{\text{Rate of Diffusion for gas 1}}{\text{Rate of Diffusion for gas 2}} = \frac{\sqrt{\frac{3RT}{M_1}}}{\sqrt{\frac{3RT}{M_2}}} = \frac{\sqrt{M_2}}{\sqrt{M_1}}$ Simplify!
 $\frac{r_1}{r_2} = \frac{\sqrt{M_2}}{\sqrt{M_1}}$

b. This is called Graham's Law

EX: The diffusion rate of an unknown gas is measured and found to be 31.50 mL/min. Under identical experimental conditions, the diffusion rate of O_2 is found to be 30.50 mL/min. If the choices are CH_4 , CO , NO , and CO_2 , what is the identity of the unknown gas?

$\frac{r_x}{r_{\text{O}_2}} = \frac{\sqrt{M_{\text{O}_2}}}{\sqrt{M_x}}$
 $\frac{31.50 \text{ mL/min}}{30.50 \text{ mL/min}} = \frac{\sqrt{32 \text{ g/mol}}}{\sqrt{M_x}}$
 $\sqrt{M_x} = \frac{\sqrt{32 \text{ g/mol}}}{\frac{31.50}{30.50}} = \frac{5.66}{1.049} = 5.40$
 $M_x = 29.16 \text{ g/mol} \approx 29 \text{ g/mol} = \text{NO}$

** With the diffusion, the above calculation is only a prediction. Because of all the collisions with air, the actual rates of diffusion are much more complex to determine.

V. Gas Effusion: refers to the passage of a gas through a tiny hole in an evacuated chamber. The same relationship to molar mass applies as above.

** Unlike diffusion, determination of the rates of effusion are much more precise as this transfer of gas occurs in a vacuum where there are no air molecules to interfere.

EX: Calculate the ratio of the effusion rates of hydrogen gas and uranium hexafluoride (UF_6), a gas used in the enrichment process to produce fuel for nuclear reactors. (M of UF_6 is 352.02 g/mol)

$\frac{r_{\text{H}_2}}{r_{\text{UF}_6}} = \frac{\sqrt{M_{\text{UF}_6}}}{\sqrt{M_{\text{H}_2}}} = \frac{\sqrt{352.02 \text{ g/mol}}}{\sqrt{2.02 \text{ g/mol}}} \approx 13.2$
 hydrogen released 13 x faster

$$KE = \frac{1}{2} m u^2$$

$$KE \text{ per mole} = \frac{3}{2} RT$$

So,
(mole) $\frac{1}{2} m u^2 = \frac{3}{2} RT$

$$u^2 = \frac{\frac{3}{2} RT \times 2}{m \text{ mole}}$$

$$u = \sqrt{\frac{3RT}{m}}$$

Remember $R = 8.314 \frac{\text{J}}{\text{mole K}}$

b/c R with $\text{J} = \frac{\text{kg m}^2}{\text{s}^2 \text{ mole K}}$

Deviations from Ideal Behavior Ap Chemistry

A. What is IDEAL gas behavior? Think of assumptions from KMT....

negligible volume, constant ^{motion} random collisions,
no attractive/repulsive forces, ~~loss of KE~~ / $KE \propto \text{temp}$

* Most real gases behave "ideally" but not under ALL conditions. Under what conditions would gases most likely exhibit NON-IDEAL BEHAVIOR???

1. Low temp Lose E, slow down, close together, attractive forces ^{Kinetic} volume becomes significant
2. High Pressure increased interactions, more collisions,

B. Dutch physicist, J.D. van der Waal, designed a new $PV=nRT$ equation which takes into account non-ideal behavior:

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

correction factors

- How does this "van der Waal's" Eq correct for non-ideal behavior?

- takes into account attractive forces & volume of actual gas particles
1. **PRESSURE CORRECTION:** When attractive forces are present, pressure will be LOWER than expected. Why? Attractive forces pull on gas particles, lessening KE that would be present in w/ a collision w/ the container.

So, to increase the pressure to what it *should* be ideally, use the an^2/V^2 correction factor!!

a = measure of Attractive Force; experimentally determined.

* The lower the " a " value, the lower the attractive forces.

Gas	a (atm·L ² /mol ²)
He	0.034
H ₂	0.244
N ₂	1.39
H ₂ O	5.46

n^2/V^2 = related to Concentration. Says that the attractive force goes up with concentration. The more particles per volume, the greater the likely hood for less forceful collisions.

2. **VOLUME CORRECTION:** V represents volume of container! Gas particles DO have a finite volume. We must minus off the actual volume of gas particles to get true volume of container.

Use volume correction factor, $(V - nb)$

n = # of moles

b = a constant that is unique for every gas; it is proportional to molar mass.

The more massive, usually, the greater the volume.

Gas	b (L/mol)
He	0.024
N ₂	0.040
CCl ₄	0.138

1. consider the gas
2. consider the conditions

C. Problems....

1. Which of the following would deviate the MOST from ideal gas behavior? Explain your reasoning.

- a. O_2 at $0^\circ C$ and 5 atm
 - b. He at $0^\circ C$ and 5 atm
 - c. O_2 at $400^\circ C$ and 1 atm
 - d. He at $400^\circ C$ and 1 atm
- @ low temps & high pressure ideal gases are deviated from most (attractive forces kick in)
 $O_2 > He$ (mass) b factor (O_2 is farther from negligible vol)
 deviations are greatest at $\uparrow P$ & $\downarrow Temp$. these have neither.

2. The molar volume of isopentane, C_5H_{12} , is 1.0 L at 503 K and 30.0 atm.

- a. Does Isopentane behave as an ideal gas? NO! (high P)
- b. Given that $a = 17.0 \frac{atm \cdot L^2}{mol^2}$ and $b = 0.136 L/mol$, calculate the pressure of isopentane as predicted by the van der Waal's equation and compare it to P predicted by $PV=nRT$.

(A) $P = 30.0 atm$ (experimentally determined)

2 formulas: $PV = nRT$
 $(P + \frac{a n^2}{V^2})(V - nb) = nRT$

close to experimental measured P

$$V = \frac{nRT}{P} = \frac{(1^{00} mol)(0.0821 atm \cdot L / mol \cdot K)(503 K)}{1^{00} atm}$$

$$V = 41.296 L \approx 41.3 L \text{ conventional.}$$

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

$$\frac{41.3}{0.864} = \frac{(1^{00} mol)(0.0821 atm \cdot K / mol \cdot K)(503 K)}{1^{00} L - (1^{00} mol)(0.136 \frac{L}{mol})} - \frac{17 atm \cdot L^2 (1^{00} mol)}{(1^{00})^2}$$

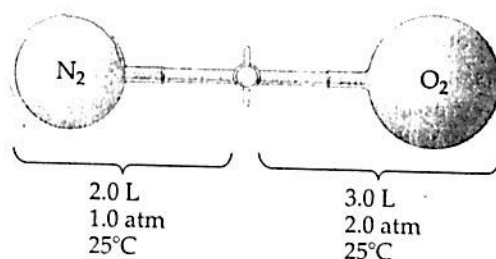
$$47.8 - 17.0 = 30.8 atm$$

The experiment supports the fact that a large deviation! this gas is not behaving ideally

For those who dare to enter the two chambers!

1. Consider the apparatus below. When the stopcock (valve) is opened between the two chambers and the gases are allowed to mix:

- A) What is the resulting pressure in the vessel?
- B) What is the mole fraction of the resulting gas mixture?
- C) What is the partial pressure of the N_2 gas in mmHg?
- D) If we inject 5.0 moles of an unknown inert gas. What is the resulting partial pressure of N_2 in bars? (of course the temperature didn't change).



2. What is meant by the term *free mean path*?

A given amount of ideal gas is placed in a piston and subjected to the following changes (restrictions are noted). How is fmp affected by the following scenarios?

- A) Increase density
- B) Increase temperature at constant volume
- C) Increase pressure at constant temperature
- D) Increase volume at constant temperature
- E) Size of the atoms

for those who don't

④ $PV = nRT$

$$n_{N_2} = \frac{PV}{RT} = \frac{1.0 \text{ atm}(2.0 \text{ L})}{.0821 \frac{\text{L atm}}{\text{mol K}} \cdot 298 \text{ K}} = .082 \text{ mol } N_2$$

$$n_{O_2} = \frac{PV}{RT} = \frac{2.0 \text{ atm}(2.0 \text{ L})}{.0821 \frac{\text{L atm}}{\text{mol K}} (298 \text{ K})} = .25 \text{ mol } O_2$$

$$n_T = .082 + .25 = .33 \text{ mol}$$

③ $X_{N_2} = \frac{.082}{.33} = .25$
 $X_{O_2} = \frac{.25}{.33} = .75$

(A) cont. $V_T = 5.0 \text{ L}$
 $T = 25.0^\circ \text{C}$
 $n = .33$

$$PV = nRT$$

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

moles

$$= \frac{.33(.0821 \text{ L atm/mol K})(298 \text{ K})}{5.0 \text{ L}}$$

1.6 atm

54. potassium thiocyanate

53. oxygen difluoride

52. iron (II) phosphate

51. silver sulfide

Write the correct formula for the following:

50. P_2O_5

49. $Sn_3(PO_4)_2$

48. Br_2O_7

47. Cu_3N_2

Name the following:

$$P_T(X_{N_2}) = P_{N_2}$$

$$(C) \quad 1.6 \underset{\text{atm}}{(.25)} (760 \text{ mmHg} / 1.0 \text{ atm}) = 300 \text{ mmHg}$$

$$(D) \quad 300 \text{ mmHg} \times \frac{1.0 \text{ atm}}{760 \text{ mmHg}} \times \frac{1.01325 \text{ bars}}{1.0 \text{ atm}}$$

0.40 bars

54. sodium cyanide

53. iron (III) nitrate

52. sulfur trioxide

51. silver sulfide

Write the correct formula for the following:

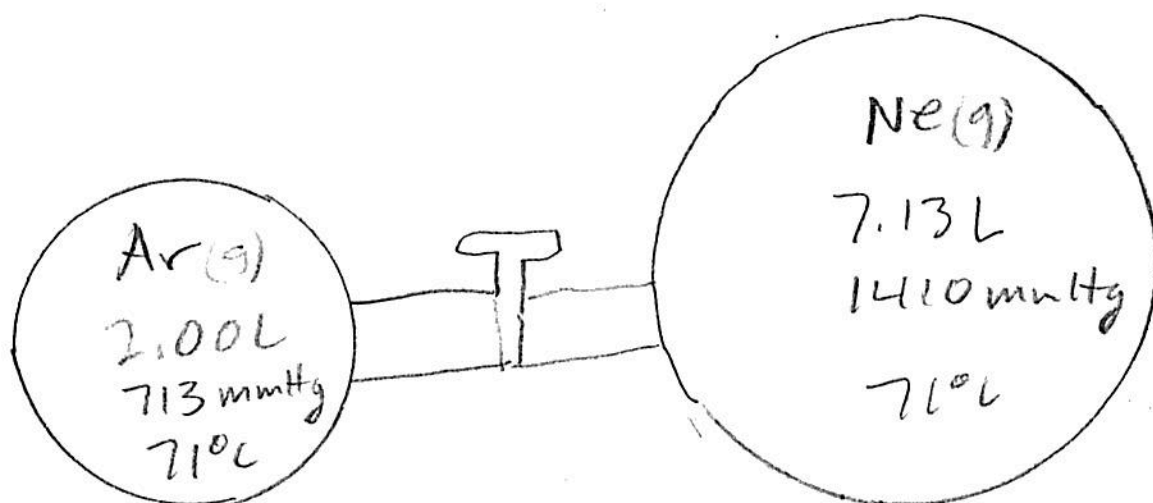
50. Ca_3N_2

49. $\text{Sn}_3(\text{PO}_4)_2$

48. Cl_2O_7

47. NaH

Name the following:



- ① What is the resultant pressure when the valve is open? Assume no temperature Δ .
- ② What is pressure if the temp increases to 110°C?
- ③ What is the mole fraction of Ar gas in the resultant mixture?

1

Problem on Back

$$n_{Ar} = \frac{713 \text{ mmHg} \left(\frac{1 \text{ atm}}{760 \text{ mmHg}} \right) (2.00 \text{ L})}{0.0821 \text{ L atm / mol K} (344 \text{ K})} = 0.0664 = n_{Ar}$$

$$n_{He} = \frac{1410 \text{ mmHg} \left(\frac{1 \text{ atm}}{760 \text{ mmHg}} \right) (7.13 \text{ L})}{0.0821 \text{ L atm / mol K} (344 \text{ K})} = 0.468 = n_{He}$$

$$P_{Total} = \frac{n_{total} R T}{V_{combined}} = \frac{0.534 (0.0821) (344)}{9.13}$$

$$P_{+} = \boxed{1.65 \text{ atm}}$$

2

$$\frac{0.534 (0.0821) (383)}{9.13} = \boxed{1.83 \text{ atm}}$$

3 $n_{Ar} + n_{He} = n_{+}$

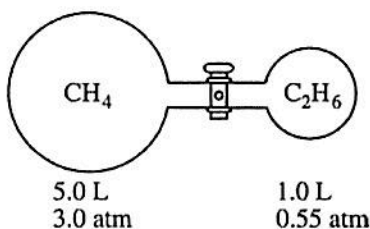
$$\frac{n_{Ar}}{n_{+}} = X_{Ar} \quad \frac{0.0664}{0.534} = 0.124 = X_{Ar}$$

2004 AP[®] CHEMISTRY FREE-RESPONSE QUESTIONS (Form B)

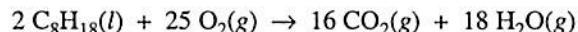
Answer EITHER Question 2 below OR Question 3 printed on page 8. Only one of these two questions will be graded. If you start both questions, be sure to cross out the question you do not want graded. The Section II score weighting for the question you choose is 20 percent.

2. Answer the following questions related to hydrocarbons.

- (a) Determine the empirical formula of a hydrocarbon that contains 85.7 percent carbon by mass.
- (b) The density of the hydrocarbon in part (a) is 2.0 g L^{-1} at 50°C and 0.948 atm .
- Calculate the molar mass of the hydrocarbon.
 - Determine the molecular formula of the hydrocarbon.
- (c) Two flasks are connected by a stopcock as shown below. The 5.0 L flask contains CH_4 at a pressure of 3.0 atm , and the 1.0 L flask contains C_2H_6 at a pressure of 0.55 atm . Calculate the total pressure of the system after the stopcock is opened. Assume that the temperature remains constant.



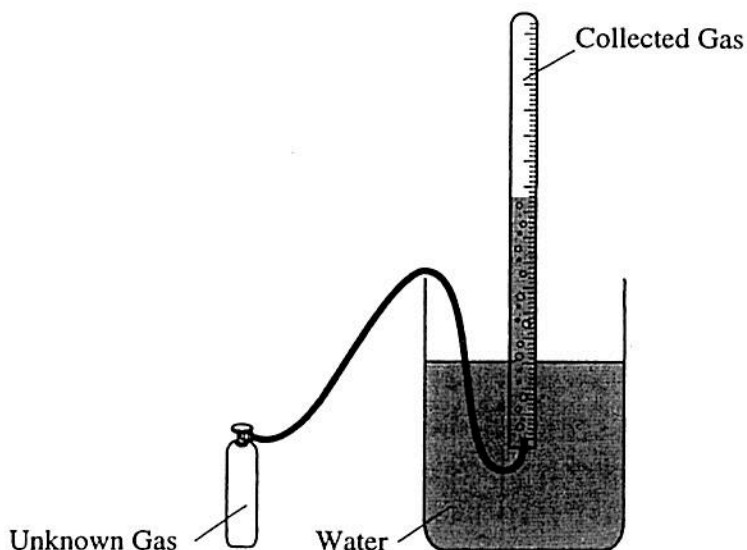
- (d) Octane, $\text{C}_8\text{H}_{18}(l)$, has a density of 0.703 g mL^{-1} at 20°C . A 255 mL sample of $\text{C}_8\text{H}_{18}(l)$ measured at 20°C reacts completely with excess oxygen as represented by the equation below.



Calculate the total number of moles of gaseous products formed.

Your responses to the rest of the questions in this part of the examination will be graded on the basis of the accuracy and relevance of the information cited. Explanations should be clear and well organized. Examples and equations may be included in your responses where appropriate. Specific answers are preferable to broad, diffuse responses.

Answer BOTH Question 5 below AND Question 6 on the next page. Both of these questions will be graded. The Section II score weighting for these questions is 30 percent (15 percent each).



5. A student performs an experiment to determine the molar mass of an unknown gas. A small amount of the pure gas is released from a pressurized container and collected in a graduated tube over water at room temperature, as shown in the diagram above. The collection tube containing the gas is allowed to stand for several minutes, and its depth is adjusted until the water levels inside and outside the tube are the same. Assume that:
- the gas is not appreciably soluble in water
 - the gas collected in the graduated tube and the water are in thermal equilibrium
 - a barometer, a thermometer, an analytical balance, and a table of the equilibrium vapor pressure of water at various temperatures are also available.
- (a) Write the equation(s) needed to calculate the molar mass of the gas.
- (b) List the measurements that must be made in order to calculate the molar mass of the gas.
- (c) Explain the purpose of equalizing the water levels inside and outside the gas collection tube.
- (d) The student determines the molar mass of the gas to be 64 g mol^{-1} . Write the expression (set-up) for calculating the percent error in the experimental value, assuming that the unknown gas is butane (molar mass 58 g mol^{-1}). Calculations are not required.
- (e) If the student fails to use information from the table of the equilibrium vapor pressures of water in the calculation, the calculated value for the molar mass of the unknown gas will be smaller than the actual value. Explain.

GO ON TO THE NEXT PAGE 

2004 Free Response Question #2

(2) hydrocarbons contain C and H only hence hydrocarbons.

(A) $\frac{85.7 \text{ g C}}{12.01 \text{ g C/mol}} = \frac{7.14 \text{ mol C}}{7.14} = 1 \text{ C}$



$\frac{14.3 \text{ g H}}{1.01 \text{ g H/mol}} = \frac{14.16 \text{ mol H}}{7.14} = 1.98 \sim 2.0 \text{ H}$

(B) Density = $\frac{m}{V} = \frac{PM}{RT}$ $M = \frac{mRT}{VP}$ $PV = nRT$

(i) $M = \frac{2.0 \text{ g} (0.0821 \text{ L atm/mol K}) (323 \text{ K})}{1.0 \text{ L} (0.948 \text{ atm})} = 55.95 \frac{\text{g}}{\text{mol}}$

(ii) $\frac{12.01 + 2.02}{14.03} = \frac{55.95}{14.03} = 3.98 \sim 4$ C_4H_8

pick any temp. It won't change

(C) $PV = nRT$ $n = \frac{PV}{RT}$ $n_{\text{CH}_4} = \frac{3.0 \text{ atm} (5.0 \text{ L})}{0.0821 \text{ L atm/mol K} (273 \text{ K})} = 0.669 \text{ mol CH}_4$

$n = \frac{PV}{RT}$ $n_{\text{C}_2\text{H}_6} = \frac{0.55 \text{ atm} (1.0 \text{ L})}{0.0821 (273)} = 0.0245 \text{ mol C}_2\text{H}_6$

$n_{\text{Total}} = 0.6935$

$T_{\text{Total}} = 6.0 \text{ L}$

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

$P_{\text{Total}} = \frac{0.6935 (0.0821) (273)}{6} = 2.59 \text{ atm}$

(D) $0.703 \frac{\text{g}}{\text{mL}} \times 255 \text{ mL} = 179.265 \text{ g C}_8\text{H}_{18}$ $\times \frac{1 \text{ mol C}_8\text{H}_{18}}{114.26 \text{ g/mol}} = 1.57 \text{ mol C}_8\text{H}_{18}$ $\times \frac{34 \text{ mol gas}}{2 \text{ mol octane}} = 26.69 \text{ mol gas}$

Scoring Guidelines for Free-Response Question 5

Question 5

(8 points)

(a) $PV = nRT$ AND $n = \frac{m}{M}$, OR molar mass = $\frac{mRT}{PV}$, OR $M = \frac{DRT}{P}$ 1 pt

- (b) temperature, atmospheric pressure, volume of the gas, and mass of gas (mass of pressurized container before and after releasing the gas) 3 pts

Note: 1 point earned for any two of the above, 2 points earned for any three of them.
 “The mass of the gas” is acceptable as a “measurement” for the 1st or 2nd point.
 Extraneous measurements (e.g., density, volume of liquid, etc.) are ignored. To earn 3rd point, “mass of pressurized container before and after releasing the gas”, or “change in mass of container” must be indicated.

- (c) to equalize internal pressure with room pressure (atmospheric pressure), or the pressure(s) will be the same. 1 pt

(d) % error = $\frac{(64 - 58) \text{ g}}{58 \text{ g}} \times 100\%$ (or $\frac{6}{58} \times 100\%$, or $\frac{6}{58}$) 1 pt

Note: No points earned for generic response (e.g., $\frac{|\text{expt.} - \text{theor.}|}{\text{theor.}} \times 100$),
 or for $\frac{6}{64} \times 100\%$. No penalty if “ $\times 100\%$ ” is absent or if value (10%) is not calculated.

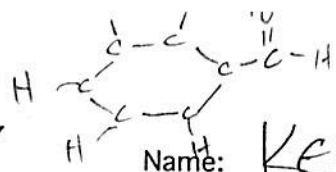
- (e) Pressure will be larger, therefore number of moles will be larger 1 pt
 molar mass = $\frac{\text{mass}}{\text{moles}}$, therefore calculated molar mass will be smaller 1 pt

OR, $M = \frac{mRT}{PV}$ (or $M = \frac{DRT}{P}$), and the denominator, PV , will be too large. 2 pts

Therefore, the value of the molar mass ($= \frac{mRT}{PV}$ or $\frac{DRT}{P}$) will be too small.

- OR, The pressure is larger, or the number of moles is larger, or since $P_{\text{total}} = (P_{\text{unknown}} - P_{\text{water}})$ we know that $P_{\text{total}} > P_{\text{unknown}}$. 1 pt only

Note: If $n = \frac{m}{M}$ is missing in part (a) but present in part (e), 1 point is earned for part (a).

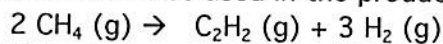


AP Chemistry Review

Class Period: ALL

Answers for which the appropriate work is not shown will receive no credit.

1. Methane can be used in the production of acetylene, according to the equation:



A 50.0-L steel vessel, filled with CH_4 to a pressure of 10.0 atm at 25°C , is heated to 1600 K to crack CH_4 and produce C_2H_2 . a. Find the mass of C_2H_2 that can be produced. $PV = nRT$

$$n_{\text{CH}_4} = \frac{PV}{RT} = \frac{10.0 \text{ atm} (50.0 \text{ L})}{0.0821 (298 \text{ K})} \times \frac{1 \text{ mol C}_2\text{H}_2}{2 \text{ mol CH}_4} \times \frac{26.04 \text{ g C}_2\text{H}_2}{1 \text{ mol C}_2\text{H}_2} = 266 \text{ g C}_2\text{H}_2$$

b. Find the pressure of the reactor at 1600 K after the reaction is complete.

2 volumes of $\text{CH}_4 \rightarrow$ 4 volumes of products.

$$\frac{266 \text{ g C}_2\text{H}_2 \times 1 \text{ mol C}_2\text{H}_2}{26.04 \text{ g C}_2\text{H}_2} \times 4 = \text{total moles} = 40.9 \text{ moles}$$

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

$$P = \frac{40.9 (0.0821) 1600}{50.0 \text{ L}} = 107 \text{ atm}$$

2. Benzaldehyde is a fragrant molecule used in artificial cherry flavoring. Combustion of 125 mg of benzaldehyde gives 363 mg of carbon dioxide and 63.7 mg of water. In another experiment, a 110-mg sample is vaporized at 150°C in a 0.100-L bulb. The vapor produces a pressure of 274 torr. Determine the molecular formula of benzaldehyde.

$$.363 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g}} \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} \times \frac{1000 \text{ mg}}{1 \text{ g}} = 99.1 \text{ mg C}$$

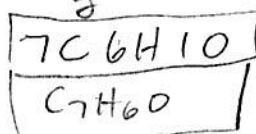
$$\begin{array}{r} 125 \text{ mg} \\ - 99.1 \text{ mg C} \\ - 7.14 \text{ mg H} \\ \hline 19 \text{ mg O} \end{array}$$

$$.0637 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g}} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} \times \frac{1.01 \text{ g H}}{1 \text{ mol H}} \times \frac{1000 \text{ mg}}{1 \text{ g}} = 7.14 \text{ mg H}$$

$$\frac{99.1 \text{ mg}}{12.01 \text{ g/mol}} = 8.25 \text{ mol C}$$

$$\frac{7.14 \text{ mg}}{1.01 \text{ g/mol}} = 7.07 \text{ mol H}$$

$$\frac{15.2 \text{ mg O}}{16.00 \text{ g/mol}} = .95 \text{ mol O}$$



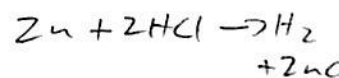
$$n = \frac{PV}{RT} = \frac{.361 (1.00)}{.0821 (423 \text{ K})} = .0104 \text{ mol}$$

$$\frac{.110 \text{ g}}{.0104 \text{ mol}} = 105.8 \text{ g/mol}$$

3. Bronze is an alloy of copper and zinc. When a 5.73-g sample of bronze was treated with excess aqueous HCl, 21.3 mg of H_2 was produced. Find the percentage by mass composition of the bronze.

HCl will react w/ Zn, not Cu! (Activity Series)

$$.0213 \text{ g H}_2 \times \frac{1 \text{ mol H}_2}{2.02 \text{ g H}_2} \times \frac{1 \text{ mol Zn}}{1 \text{ mol H}_2} \times \frac{65.38 \text{ g Zn}}{1 \text{ mol Zn}} = .689 \text{ g Zn}$$

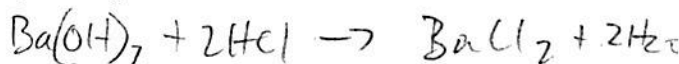


$$\frac{.689}{5.73} \times 100 = 12.0 \% \text{ Zn}$$

$$\begin{array}{r} 5.73 \\ - .689 \\ \hline 5.04 \end{array}$$

$$\frac{5.04}{5.73} \times 100 = 88.0 \% \text{ Cu}$$

4. Calculate the molarities of all ions present in a solution made by mixing 150.0 mL of $2.00 \times 10^{-2} \text{ M}$ $\text{Ba}(\text{OH})_2$ solution with 100.0 mL of $5.00 \times 10^{-2} \text{ M}$ HCl solution.



$$.150 \text{ L} \times .0200 \text{ mol Ba}(\text{OH})_2 = .00300 \text{ mol Ba}(\text{OH})_2$$

$$.100 \text{ L} \times .0500 \text{ mol HCl} = .005 \text{ mol HCl}$$

L.R!

$$[\text{H}^+] = 0$$

$$\frac{.005 \text{ mol Cl}^-}{.250 \text{ L}} = .020 = [\text{Cl}^-]$$

$$\frac{.003 \text{ mol Ba}^{2+}}{.250 \text{ L}} = [\text{Ba}^{2+}] = .012$$

$$\begin{array}{r} .00600 \text{ mol OH}^- \\ - .00500 \text{ mol H}^+ \\ \hline .00100 \text{ mol OH}^- \end{array}$$

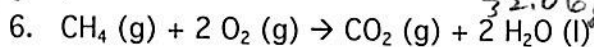
$$.00100 \text{ mol OH}^- / .250 \text{ L} = .00400 = [\text{OH}^-]$$



5. Cinnabar is an ore of mercury known to contain only Hg and S. When a 0.350-g sample is heated in oxygen, the ore decomposes completely, giving 0.302 g of pure Hg metal. Find the empirical formula of cinnabar.

$$\frac{0.302 \text{ g Hg}}{200.59 \text{ g Hg}} = 0.00151 \text{ mol Hg}$$

$$\frac{0.048 \text{ g S}}{32.06 \text{ g S}} = 0.0015 \text{ mol S}$$



$$\Delta H = -890.4 \text{ kJ/mol CH}_4$$

Calculate how many grams of water are produced if enough CH_4 is oxidized to release 4452 kJ.

$$\frac{4452}{890.4} = 5$$

$$5 \text{ mol CH}_4 \times \frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol CH}_4} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 180.2 \text{ g H}_2\text{O}$$

7. Gases are sold and shipped in metal tanks under high pressure. A typical tank of compressed air has a volume of 30.0 L and is pressurized to 15.0 atm at 298 K. What work had to be done in filling this tank?

[Hint: what volume did the air occupy before it was compressed?]

$$W = -P\Delta V$$

$$W = -15.0 \text{ atm} (\Delta V)$$

$$\Delta V = V_f - V_i = 30.0 \text{ L} - 450 \text{ L}$$

$$W = -15.0 \text{ atm} (-420 \text{ L}) \times \frac{101.3 \text{ J}}{1 \text{ atm}} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 638 \text{ kJ}$$

Compression is (+)

$$\frac{P_1 V_1}{nRT} = \frac{P_2 V_2}{nRT}$$

$$\frac{P_1 V_1}{V_1} = P_2 V_2 \Rightarrow V_1 = \frac{P_2 V_2}{P_1} = 450 \text{ L}$$

8-9. A home swimming pool contains 155 m³ of water. At the beginning of swimming season, the water must be heated from 20.0°C to 30.0°C.

a. Find the number of joules of heat energy that must be supplied.

$$\Delta H = m s \Delta T$$

$$\Delta H = 155 \text{ m}^3 \times \frac{1 \times 10^6 \text{ cm}^3}{\text{m}^3} \times \frac{1 \text{ g}}{1 \text{ cm}^3} \left(\frac{4.184 \text{ J}}{\text{g}^\circ\text{C}} \right) 10.0^\circ\text{C} = 6.49 \times 10^9 \text{ J}$$

or 6,490,000,000 J

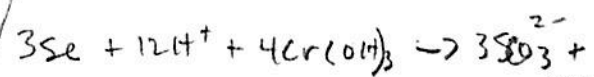
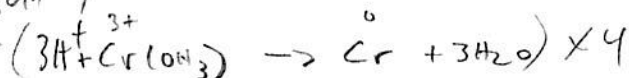
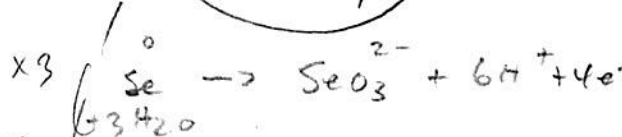
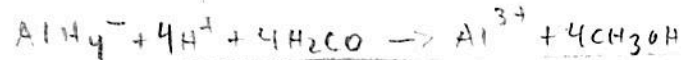
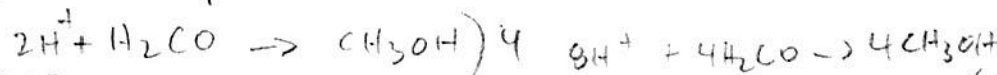
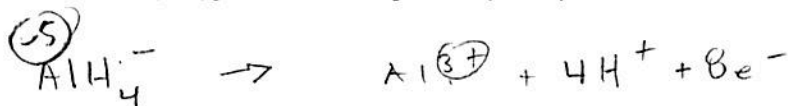
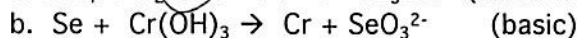
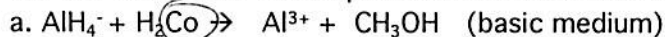
b. If this energy is supplied by a natural gas heater with an 80% heat transfer efficiency, how many grams of methane must be burned? [The heat of combustion of methane is -803 kJ/mol]

$$803 \text{ kJ/mol} \times 0.8 = 642.4 \text{ kJ/mol}$$

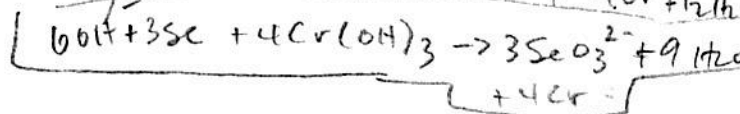
$$\frac{6,490,000,000 \text{ J}}{6,420,000 \text{ J/mol}} = \text{mol CH}_4$$

$$\times 16.05 \text{ g/mol} = 16,200 \text{ g CH}_4$$

10-11. Use redox techniques to balance the following equations:



2



12. In a series of experiments, the U.S. Navy developed an undersea habitat. In one experiment, the mole percent composition of the atmosphere in the undersea habitat was 79.0% He, 17.0% N, and 4.0% O₂. What will the partial pressure of each gas be when the habitat is 58.8 m below sea level, where the pressure is 6.91 atm? $X_{He} = .790$ $X_N = .170$ $X_O = .040$

$$P_T(X_n) = P_n \quad 6.91 \text{ atm}(.790) = P_{He} = 5.46 \text{ atm} \quad 6.91(.040) = P_{O_2} = 0.28 \text{ atm}$$

$$6.91 \text{ atm}(.170) = P_N = 1.17 \text{ atm}$$

13. A mixture of KCl (s) and KClO₃ (s) having a mass of 18.17 g is heated to convert the KClO₃ to KCl according to $2 \text{ KClO}_3 (\text{s}) \rightarrow 2 \text{ KCl} (\text{s}) + 3 \text{ O}_2 (\text{g})$. If the mixture has a mass of 12.62 g after heating, then what is the mass percentage of KClO₃ in the original mixture? $\Delta \text{ mass} \propto \text{O}_2 \text{ lost}$

$$\frac{18.17 - 12.62}{5.55 \text{ g O}_2} \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} \times \frac{2 \text{ mol KClO}_3}{3 \text{ mol O}_2} \times \frac{122.55 \text{ g KClO}_3}{1 \text{ mol KClO}_3} = 14.2 \text{ g KClO}_3$$

$$\frac{14.2}{18.17} \times 100 = \% \text{ KClO}_3 = 78.2 \%$$

14-15.  CO₂

 O₂

 He

 N₂

 CH₄

[From a previous A.P. Test] Represented above are five identical balloons, each filled to the same volume at 25°C and 1.0 atmosphere pressure with the pure gases indicated. (Then all contain the same # of moles of gas)
a. Which balloon contains the greatest mass of gas? Explain.

CO₂ has the highest M.M.

b. Compare the kinetic energies of the gas molecules in the balloons. Explain.

They are all =. $KE = \text{temp}$ Same temp = same KE

c. Which balloon contains the gas that would be expected to deviate most from the behavior of an ideal gas? Explain. all gases are non-polar, however CO₂ has the highest potential London Dispersion Forces. $\downarrow T$ and $\uparrow \text{Press}$ would liquify it!

d. Twelve hours after being filled all the balloons have decreased in size. Predict which balloon will be the smallest. Explain your reasoning. He

Through effusion, the smallest/lightest gas would escape through porous latex.

16. A reaction mixture contains 25.0 g of PCl₃ and 45.0 g of PbF₂. What mass of PbCl₂ can be obtained from the following reaction? $3 \text{ PbF}_2 + 2 \text{ PCl}_3 \rightarrow 2 \text{ PF}_3 + 3 \text{ PbCl}_2$

$$25.0 \text{ g PCl}_3 \times \frac{1 \text{ mol PCl}_3}{137.38 \text{ g PCl}_3} \times \frac{3 \text{ mol PbCl}_2}{2 \text{ mol PCl}_3} \times \frac{278.1 \text{ g PbCl}_2}{1 \text{ mol PbCl}_2} = 75.9 \text{ g PbCl}_2$$

P.B.

$$45.0 \text{ g PbF}_2 \times \frac{1 \text{ mol PbF}_2}{245.2 \text{ g PbF}_2} \times \frac{3 \text{ mol PbCl}_2}{3 \text{ mol PbF}_2} \times \frac{278.1 \text{ g PbCl}_2}{1 \text{ mol PbCl}_2} = 51.0 \text{ g PbCl}_2$$

17-19. These questions are part of a question that was on a previous AP test.



When a 2.000-g sample of pure phenol ($\text{C}_6\text{H}_5\text{OH}$) is completely burned according to the above equation, 64.98 kilojoules of heat is released. Use the information in the table below to answer the questions that follow:

Standard heat of formation

Substance	(ΔH_f°) at 25°C (kJ/mol)
C (graphite)	0.00
CO_2 (g)	-393.5
H_2 (g)	0.00
$\text{H}_2\text{O(l)}$	-285.85
O_2 (g)	0.00
$\text{C}_6\text{H}_5\text{OH (s)}$?

a. Calculate the molar heat of combustion of phenol in kJ/mol at 25°C.

$$\frac{-64.98 \text{ kJ}}{2.000 \text{ g phenol}} \times \frac{94.12 \text{ g phenol}}{1 \text{ mol phenol}} = -3,058 \frac{\text{kJ}}{\text{mol}}$$

b. Calculate the standard heat of formation, (ΔH_f°), of phenol in kJ/mol.

$$\begin{aligned} -3,058 \text{ kJ} &= \sum [6(\Delta H_f^\circ \text{CO}_2) + 3(\Delta H_f^\circ \text{H}_2\text{O})] - [\Delta H_f^\circ \text{C}_6\text{H}_5\text{OH} + 7(\Delta H_f^\circ \text{O}_2)] \\ -3,058 \text{ kJ} &= [6(-236.1) + 3(-857.55)] - [\Delta H_f^\circ \text{C}_6\text{H}_5\text{OH} + 7(0)] \\ -3,058 &= [-3,218.6 \text{ kJ}] - [\Delta H_f^\circ \text{C}_6\text{H}_5\text{OH}] \\ \Delta H_f^\circ \text{C}_6\text{H}_5\text{OH} &= -160.6 \text{ kJ/mol} \end{aligned}$$

c. If the volume of the combustion container is 10.0 L, calculate the final pressure in the container when the temperature is changed to 110.0°C. (Assume no oxygen remains unreacted and that all products are gaseous.)

$$P_V = nRT \quad P = \frac{nRT}{V} \quad n = 2.000 \text{ g phenol} \times \frac{1 \text{ mol phenol}}{94.12 \text{ g phenol}} = 0.02125 \text{ mol phenol}$$

$$P = \frac{(0.02125 \text{ mol}) (0.0821 \text{ L atm / mol K}) (383 \text{ K})}{10.0 \text{ L}} = 0.601 \text{ atm}$$

20. Oxygen is collected over water at 25°C in a 2.00 L vessel at a total barometric pressure of 765 torr. Calculate the number of moles of oxygen collected. The vapor pressure of water at 25°C is 24 torr.

$$P_T = P_{\text{O}_2} + P_{\text{H}_2\text{O}}$$

$$P_{\text{O}_2} = P_T - P_{\text{H}_2\text{O}} \rightarrow 765 \text{ torr} - 24 \text{ torr} = 741 \text{ torr}$$

$$P_V = nRT$$

$$n = \frac{P_V}{RT} = \frac{741 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} (2.00 \text{ L})}{(0.0821 \text{ L atm / mol K}) (298 \text{ K})} = 0.0797 \text{ moles O}_2$$