

P.S #15

10.1 In the gas phase molecules are far apart, while in the liquid they are touching.

- (a) A gas is much less dense than a liquid because most of the volume of a gas is empty space.
- (b) A gas is much more compressible because of the distance between molecules.
- (c) Gaseous molecules are so far apart that there is no barrier to mixing, regardless of the identity of the molecule. All mixtures of gases are homogeneous. Liquid molecules are touching. In order to mix, they must displace one another. Similar molecules displace each other and form homogeneous mixtures. Very dissimilar molecules form heterogeneous mixtures.

- 10.2
- (a) Because gas molecules are far apart and in constant motion, the gas expands to fill the container. Attractive forces hold liquid molecules together and the volume of the liquid does not change.
  - (b)  $\text{H}_2\text{O}$  and  $\text{CCl}_4$  molecules are too dissimilar to displace each other and mix in the liquid state. All mixtures of gases are homogeneous. (See Solution 10.1 (c)).
  - (c) Because gas molecules are far apart, the mass present in 1 mL of a gas is very small. The mass of a gas present in 1 L is on the same order of magnitude as the mass of a liquid present in 1 mL.

10.6 Using the relationship derived in Solution 10.5 for two liquids under the influence of gravity,  $(d \times h)_{\text{lid}} = (d \times h)_{\text{Hg}}$ . At 752 torr, the height of a Hg barometer is 752 mm.

$$\frac{1.20 \text{ g}}{1 \text{ mL}} \times h_{\text{lid}} = \frac{13.6 \text{ g}}{1 \text{ mL}} \times 760 \text{ mm}; h_{\text{lid}} = \frac{13.6 \text{ g/mL} \times 752 \text{ mm}}{1.20 \text{ g/mL}} = 8.52 \times 10^3 \text{ mm} = 8.52 \text{ m}$$

# Pressure conversions / manometer problems KEY

1)  $109 \text{ kPa} \frac{1.0 \text{ atm}}{101.325 \text{ kPa}} = \underline{1.08 \text{ atm}}$

2)  $0.62 \text{ atm} \frac{760 \text{ mmHg}}{1.0 \text{ atm}} = \underline{471 \text{ mmHg}}$

3)  $1.07 \text{ atm} \frac{760 \text{ mmHg}}{1.0 \text{ atm}} = \underline{813 \text{ mmHg}}$

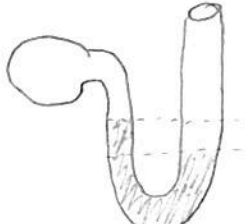
4)  $758.7 \text{ mmHg} \frac{1.0 \text{ atm}}{760 \text{ mmHg}} = \underline{0.998 \text{ atm}}$

5)  $761.3 \text{ mmHg} \frac{101,325 \text{ Pa}}{760 \text{ mmHg}} = \underline{101,498 \text{ Pa}}$

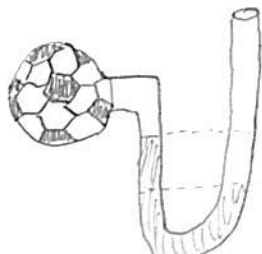
6)  $101,498 \text{ Pa} - \left( 99.82 \text{ kPa} \times \frac{1000 \text{ Pa}}{1.0 \text{ kPa}} \right) = \underline{1678 \text{ Pa}}$

7)  $43.2 \text{ lb/in}^2 \frac{1 \text{ atm}}{14.7 \text{ lb/in}^2} = \underline{2.94 \text{ atm}}$

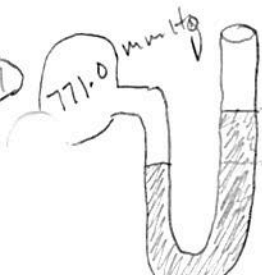
8)  $34.8 \text{ lb/in}^2 \frac{101.325 \text{ kPa}}{14.7 \text{ lb/in}^2} = \underline{239.7 \text{ kPa}}$

9) 

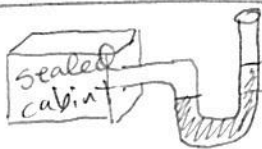
$$\begin{array}{r} 750 \text{ mmHg} \\ + 15 \text{ mmHg} \\ \hline \boxed{765 \text{ mmHg}} \end{array}$$

10) 

$$\begin{array}{r} 770 \text{ mmHg} \\ - 10 \text{ mmHg} \\ \hline \boxed{760 \text{ mmHg}} \end{array}$$

11) 

$$\begin{array}{r} 771.0 \text{ mmHg} \\ - 11.2 \text{ mmHg} \\ \hline \boxed{759.8 \text{ mmHg}} \end{array}$$

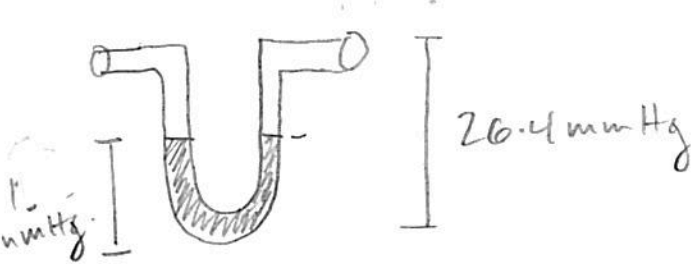
12) 

$$\begin{array}{r} 762.4 \text{ mmHg} \\ + 3.6 \\ \hline 766.0 \text{ mmHg} \end{array}$$

$$766.0 \text{ mmHg} \frac{101.325 \text{ kPa}}{760.0 \text{ mmHg}} = \underline{\boxed{102.1 \text{ kPa}}}$$

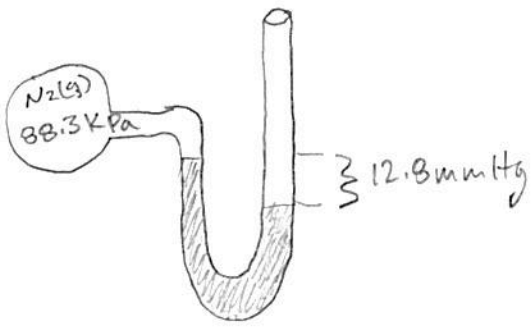
# manometer PROBLEMS (CONT)

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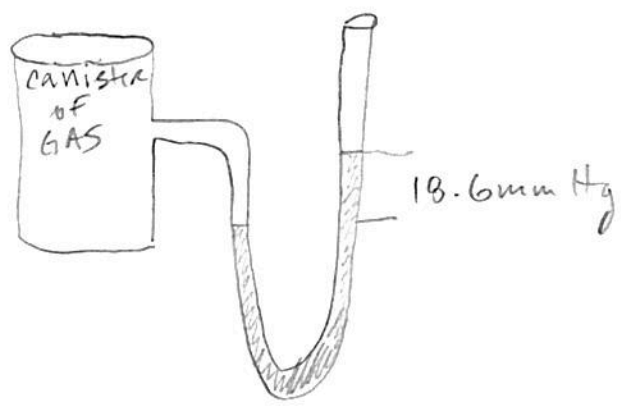
$$\begin{array}{r} 26.4 \text{ mmHg} \\ - 13.2 \text{ mmHg} \\ \hline 13.2 \text{ mmHg} \end{array} \quad \begin{array}{l} \text{in one direction} \\ \text{OR} \\ \text{26.4 mmHg total} \end{array}$$

14



$$\begin{array}{r} 12.8 \text{ mmHg} \quad \frac{101.325 \text{ KPa}}{760 \text{ mmHg}} = 1.707 \text{ KPa} \\ 88.300 \text{ KPa} \\ + 1.707 \text{ KPa} \\ \hline 90.0 \text{ KPa} \end{array}$$

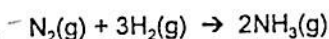
15



$$\begin{array}{r} 1.03 \text{ atm} \quad \frac{760 \text{ mmHg}}{1.00 \text{ atm}} = 782.8 \text{ mmHg} \\ 782.8 \text{ mmHg} \\ + 18.6 \text{ mmHg} \\ \hline 801.4 \text{ mmHg} \end{array}$$

- 10.18 (a) P and V are inversely proportional at constant T. If the volume decreases by a factor of 3, the pressure increases by a factor of 3.
- (b) P and T are directly proportional at constant V. If T decreases by a factor of 2, P also decreases by a factor of 2.
- (c) P and n are directly proportional at constant V. If n decreases by a factor of 2, P also decreases by a factor of 2.

10.22 According to Avogadro's hypothesis, the mole ratios in the chemical equation will be volume ratios for the gases if they are at the same temperature and pressure.



The volumes of  $\text{H}_2$  and  $\text{N}_2$  are in a stoichiometric  $\frac{3.6 \text{ L}}{1.2 \text{ L}}$  or  $\frac{3 \text{ vol H}_2}{1 \text{ vol N}_2}$  ratio, so either can be used to determine the volume of  $\text{NH}_3(\text{g})$  produced.

$$1.2 \text{ L N}_2 \times \frac{2 \text{ mol NH}_3}{1 \text{ mol N}_2} = 2.4 \text{ L NH}_3(\text{g}) \text{ produced}$$

10.26  $n = g/M$ ;  $PV = nRT = gRT/M$ ;  $M = gRT/PV$ .

2-L flask:  $M = 4.8 \text{ T}/2.0(\text{X}) = 2.4 \text{ T/X}$

3-L flask:  $M = 0.36 \text{ T}/3.0 (0.1 \text{ X}) = 1.2 \text{ T/X}$

The molar masses of the two gases are not equal. The gas in the 2-L flask has a molar mass that is twice as large as the gas in the 3-L flask.

10.30 Find the volume of the tube in  $\text{cm}^3$ ;  $1 \text{ cm}^3 = 1 \text{ mL}$ .

$r = d/2 = 4.5 \text{ cm}/2 = 2.25 = 2.3 \text{ cm}$ ;  $h = 5.3 \text{ m} = 5.3 \times 10^2 \text{ cm}$

$V = \pi r^2 h = 3.14159 \times (2.25 \text{ cm})^2 \times (5.3 \times 10^2 \text{ cm}) = 8.429 \times 10^3 \text{ cm}^3 = 8.4 \text{ L}$

$PV = \frac{g}{M} RT$ ;  $g = \frac{MPV}{RT}$ ;  $P = 2.03 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 2.621 \times 10^{-3} = 2.62 \times 10^{-3} \text{ atm}$

$g = \frac{20.18 \text{ g Ne}}{1 \text{ mol Ne}} \times \frac{\text{K} \cdot \text{mol}}{0.08206 \text{ L} \cdot \text{atm}} \times \frac{2.67 \times 10^{-3} \text{ atm} \times 8.429 \text{ L}}{308 \text{ K}} = 0.018 \text{ g Ne}$

10.43 Analyze/Plan. Given: mass, conditions (P, V, T) of unknown gas. Find: molar mass.  $M = gRT/PV$ . Solve:

$M = \frac{gRT}{PV} = \frac{1.012 \text{ g}}{0.354 \text{ L}} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{K} \cdot \text{mol}} \times \frac{372 \text{ K}}{742 \text{ torr}} \times \frac{760 \text{ torr}}{1 \text{ atm}} = 89.4 \text{ g/mol}$

10.46  $n_{\text{H}_2} = \frac{P_{\text{H}_2} V}{RT}$ ;  $P = 814 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 1.071 = 1.07 \text{ atm}$

$n_{\text{H}_2} = 1.071 \text{ atm} \times \frac{\text{K} \cdot \text{mol}}{0.08206 \text{ L} \cdot \text{atm}} \times \frac{64.5 \text{ L}}{305 \text{ K}} = 2.760 = 2.76 \text{ mol H}_2$

$2.760 \text{ mol H}_2 \times \frac{1 \text{ mol CaH}_2}{2 \text{ mol H}_2} \times \frac{42.10 \text{ g CaH}_2}{1 \text{ mol CaH}_2} = 58.1 \text{ g CaH}_2$

1.38 mol  $\text{CaH}_2$

$PV = nRT$   
 $PV = \frac{g}{M} RT$   
 $M = \frac{gRT}{PV}$

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

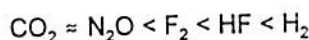
- 10.52 (a) The partial pressure of gas A is **not affected** by the addition of gas C. The partial pressure of A depends only on moles of A, volume of container and conditions; none of these factors change when gas C is added.
- (b) The total pressure in the vessel **increases** when gas C is added, because the total number of moles of gas increases.
- (c) The mole fraction of gas B **decreases** when gas C is added. The moles of gas B stay the same, but the total moles increase, so the mole fraction of B ( $n_B/n_t$ ) decreases.

10.57 *Analyze/Plan.* Mole fraction = pressure fraction. Find the desired mole fraction of  $O_2$  and change to mole percent. *Solve:*

$$x_{O_2} = \frac{P_{O_2}}{P_t} = \frac{0.21 \text{ atm}}{8.38 \text{ atm}} = 0.025; \text{ mole \%} = 0.025 \times 100 = 2.5\%$$



- 10.67 (a) *Plan.* The greater the molecular (and molar) mass, the smaller the rms speed of the molecules. *Solve:* In order of increasing speed (and decreasing molar mass):



- (b) *Plan.* Follow the logic of Sample Exercise 10.14. *Solve:*

$$u_{\text{H}_2} = \sqrt{\frac{3RT}{M}} = \left( \frac{3 \times 8.314 \text{ kg} \cdot \text{m}^2/\text{s}^2 \cdot \text{K} \cdot \text{mol} \times 300 \text{ K}}{2.02 \times 10^{-3} \text{ kg/mol}} \right)^{1/2} = 1.92 \times 10^3 \text{ m/s}$$

$$u_{\text{CO}_2} = \left( \frac{3 \times 8.314 \text{ kg} \cdot \text{m}^2/\text{s}^2 \cdot \text{K} \cdot \text{mol} \times 300 \text{ K}}{44.0 \times 10^{-3} \text{ kg/mol}} \right)^{1/2} = 4.12 \times 10^2 \text{ m/s}$$

As expected, the lighter molecule moves at the greater speed.

- 10.63 The fact that gases are readily compressible supports the assumption that most of the volume of a gas sample is empty space.

- 10.71 *Analyze.* Given: relative effusion rates of two gases at same temperature. Find: molecular formula of one of the gases. *Plan.* Use Graham's law to calculate the formula weight of arsenic (III) sulfide, and thus the molecular formula. *Solve:*

$$\frac{\text{rate (sulfide)}}{\text{rate (Ar)}} = \left[ \frac{39.9}{M(\text{sulfide})} \right]^{1/2} = 0.28$$

$$M(\text{sulfide}) = (39.9 / 0.28)^2 = 510 \text{ g/mol (two significant figures)}$$

The empirical formula of arsenic(III) sulfide is  $\text{As}_2\text{S}_3$ , which has a formula mass of 246.1. Twice this is 490 g/mol, close to the value estimated from the effusion experiment. Thus, the formula of the vapor phase molecule is  $\text{As}_4\text{S}_6$ .

- 10.72 The time required is proportional to the reciprocal of the effusion rate.

$$\frac{\text{rate (X)}}{\text{rate (O}_2)} = \frac{105 \text{ s}}{31 \text{ s}} = \left[ \frac{32 \text{ g O}_2}{M_x} \right]^{1/2}; M_x = 32 \text{ g O}_2 \times \left[ \frac{105}{31} \right]^2 = 370 \text{ g/mol (two sig figs)}$$

- 10.79 *Analyze.* Conditions and amount of  $\text{CCl}_4(\text{g})$  are given. *Plan.* Use ideal-gas equation and van der Waals equation to calculate pressure of gas at these conditions. *Solve:*

$$(a) P = 1.00 \text{ mol} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{K} \cdot \text{mol}} \times \frac{313 \text{ K}}{28.0 \text{ L}} = 0.917 \text{ atm}$$

$$(b) P = \frac{nRT}{V - nb} - \frac{an^2}{V^2} = \frac{1.00 \times 0.08206 \times 313}{28.0 - (1.00 \times 0.1383)} - \frac{20.4(1.00)^2}{(28.0)^2} = 0.896 \text{ atm}$$

*Check.* The van der Waals result indicates that the real pressure will be less than the ideal pressure. That is, intermolecular forces reduce the effective number of particles and the real pressure. This is reasonable for 1 mole of gas at relatively low temperature and pressure.

- 10.86 If the air in the room is at STP, the partial pressure of  $\text{O}_2$  is  $0.2095 \times 1 \text{ atm} = 0.2095 \text{ atm}$ . Since the gases in air are perfectly mixed, the volume of  $\text{O}_2$  is the volume of the room.

$$V = 10.0 \text{ ft} \times 8.0 \text{ ft} \times 8.0 \text{ ft} \times \frac{(12)^3 \text{ in}^3}{\text{ft}^3} \times \frac{(2.54)^3 \text{ cm}^3}{\text{in}^3} \times \frac{1 \text{ L}}{1000 \text{ cm}^3} = 1.812 \times 10^4 \text{ L}$$

$$g = \frac{PV}{RT} = \frac{32.00 \text{ g O}_2}{\text{mol O}_2} \times \frac{\text{K} \cdot \text{mol}}{0.08206 \text{ L} \cdot \text{atm}} \times \frac{0.2095 \text{ atm} \times 1.812 \times 10^4 \text{ L}}{273 \text{ K}} = 5.4 \times 10^3 \text{ g O}_2$$

P.S  
#17  
BK  
PROB.

## Advanced Gas Laws (Key)

$$PV = nRT \quad PV = \frac{g}{M} RT \quad \frac{P}{M} = \frac{g}{RT} \quad M = \frac{1}{PV/gRT} = \frac{1}{1.013 \times 10^5 \times 22.4 \times 10^{-3} \times 273.15}$$

$$M = 24.35 \text{ g/mol} \quad \therefore \text{CH}_3 \text{ has a mass of } 15.06 \text{ g} \quad \text{or approximately } \frac{1}{2}$$

Molecular formula is  $\boxed{\text{C}_2\text{H}_6} = \text{ethane}$

$$2) \text{STP } \frac{32}{T_1} = 15 \text{ g NaHCO}_3 \times \frac{1 \text{ mol NaHCO}_3}{84.01 \text{ g}} = 0.1786 \text{ mol NaHCO}_3$$

$$0.1786 \text{ mol NaHCO}_3 \times \frac{1 \text{ mol CO}_2}{1 \text{ mol NaHCO}_3} = 0.1786 \text{ mol CO}_2 \times 22.4 \text{ L @ STP}$$

$$= 3.99 \text{ Liters} \approx 4.0 \text{ Liters @ STP}$$

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \quad T_2 \frac{P_1 V_1}{T_1 V_2} = P_2 = \frac{310 \text{ K} (1.0 \times 10^5 \text{ Pa})}{(273 \text{ K}) (1.5 \text{ L})} = \boxed{3.03 \text{ atm}}$$

$$3) \frac{V_1}{T_1} = \frac{V_2}{T_2} \quad V_2 = \frac{V_1 T_2}{T_1} \quad V_2 = \frac{0.5 \text{ L} (273 \text{ K})}{310 \text{ K}} \quad V_2 = 0.429 \text{ L}$$

$$0.429 \text{ L} \times 0.20 = 0.0858 \text{ L of O}_2 \text{ @ STP} \times \frac{1 \text{ mol}}{22.4 \text{ L}} = 0.00383 \text{ mol O}_2$$

$$0.00383 \text{ mol O}_2 \times \frac{2 \text{ mol O}_2}{25 \text{ mol O}_2} = 0.003064 \text{ mol O}_2$$

$$\times 114.26 \text{ g O}_2 \text{ before} = \boxed{0.035 \text{ grams O}_2}$$

$$1) 1345 \text{ kPa} \times \frac{1.0 \text{ atm}}{101.325 \text{ kPa}} = 13.27 \text{ atm} \quad d = \frac{PM}{RT}$$

$$d = \frac{13.27 \text{ atm} \times 28 \frac{\text{g}}{\text{mol}}}{0.0821 \text{ L atm / mol K} \times 298 \text{ K}} = \boxed{15.19 \frac{\text{g}}{\text{L}}}$$

$$d = \frac{1.0 \text{ atm} \times 28 \frac{\text{g}}{\text{mol}}}{0.0821 \text{ L atm / mol K} \times 298 \text{ K}} = \boxed{1.14 \frac{\text{g}}{\text{L}}}$$

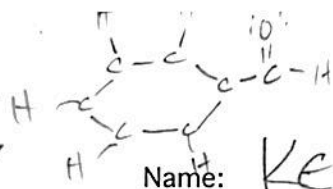
5) At STP hydrogen gas and Helium gas both occupy the same volume per mole; however, the density in g/mols/liter will determine buoyancy. Less buoyant gases will "rise" above more dense gases.

Example 1 mol of H<sub>2</sub> gas occupies 22.4 L @ STP

$$\boxed{0.046 \text{ g/Liter}}$$

1 mol of He gas occupies 22.4 L @ STP.

$$\boxed{0.179 \text{ g/Liter}}$$



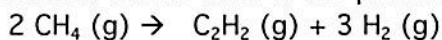
# AP Chemistry Review

Name: Key/Collins

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  $\text{CH}_4$  to a pressure of 10.0 atm at  $25^\circ\text{C}$ . is heated to 1600 K to crack  $\text{CH}_4$  and produce  $\text{C}_2\text{H}_2$ . a. Find the mass of  $\text{C}_2\text{H}_2$  that can be produced.  $PV=nRT$

$$n_{\text{CH}_4} = \frac{PV}{RT} = \frac{10.0 \text{ atm} (50.0 \text{ L})}{.0821 (298 \text{ K})} \times \frac{1 \text{ mol C}_2\text{H}_2}{2 \text{ mol CH}_4} \times \frac{26.04 \text{ g C}_2\text{H}_2}{1 \text{ mol C}_2\text{H}_2} = 266 \text{ g C}_2\text{H}_2$$

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.

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

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

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  $150^\circ\text{C}$  in a 0.100-L bulb. The vapor produces a pressure of 274 torr. Determine the molecular formula of benzaldehyde.

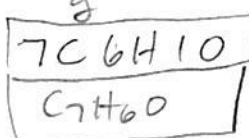
$$.363 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g}} \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} \times \frac{1000 \text{ mg}}{\text{g}} = 99.1 \text{ mg C}$$

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

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

$$\frac{79.2 \text{ g}}{12.01 \text{ g}} = 6.6 \text{ mol C}$$

$$\frac{5.6 \text{ g H}}{1.01} = 5.5 \text{ mol H} \quad \frac{15.2 \text{ g O}}{16.00} = .95 \text{ mol O}$$

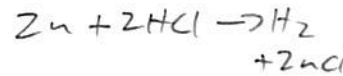


$$n = \frac{PV}{RT} = \frac{.361 (0.100)}{.0821 (423 \text{ K})} = .00104 \text{ mol} \quad \frac{.110 \text{ g}}{.00104 \text{ mol}} = 105.8 \text{ g/mol}$$

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  $\text{H}_2$  was produced. Find the percentage by mass composition of the bronze.

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

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



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

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

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!

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

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

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

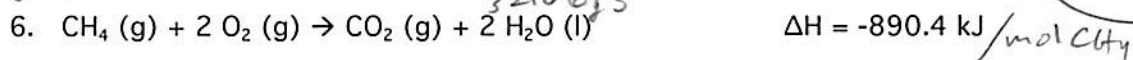
$$1 - \frac{.00600 \text{ mol OH}^-}{.250 \text{ L}} = .00400 = [\text{OH}^-]$$





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.

$$\begin{array}{r} 350 \\ - 302 \\ \hline 48 \end{array} \quad \begin{array}{r} 0.302 \text{ g Hg} \\ 200.59 \text{ g Hg} \\ \hline 0.00151 \text{ mol Hg} \end{array} \quad \begin{array}{r} 0.648 \text{ g S} \\ 32.06 \text{ g S} \\ \hline 0.0202 \text{ mol S} \end{array}$$



Calculate how many grams of water are produced if enough  $\text{CH}_4$  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?]

$$W = -P\Delta V$$

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

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

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

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

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

Compression is (+)

8-9. A home swimming pool contains 155 m<sup>3</sup> 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.

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

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

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]

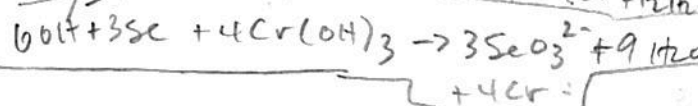
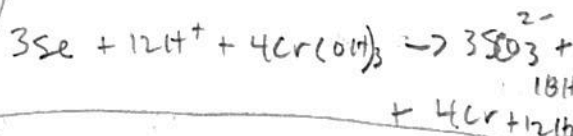
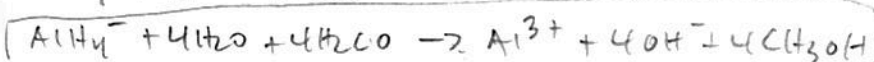
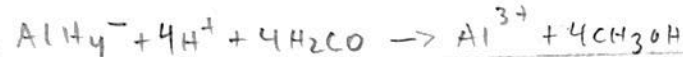
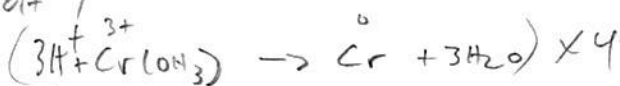
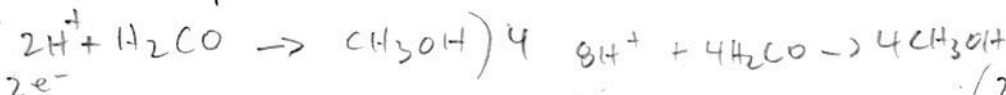
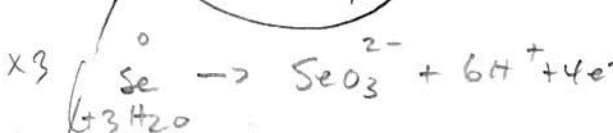
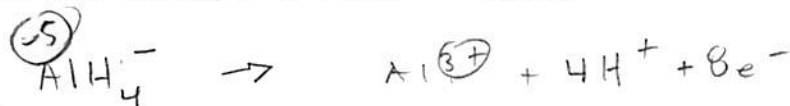
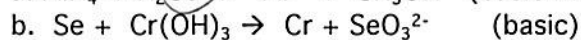
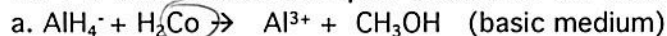
$$803 \text{ kJ} \times (0.8) \times \frac{1000 \text{ J}}{\text{kJ}} = 6,420,000 \text{ J}$$

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

$$\times 16.05 \text{ g/mol}$$

$$= 16,200 \text{ g CH}_4$$

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



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% O<sub>2</sub>. 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?  $X_{He} = .790$   $X_N = .170$   $X_{O_2} = .040$

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

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

13. A mixture of KCl (s) and KClO<sub>3</sub> (s) having a mass of 18.17 g is heated to convert the KClO<sub>3</sub> 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 KClO<sub>3</sub> in the original mixture?  $\Delta \text{ in mass } \propto \text{ O}_2 \text{ lost}$

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

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

14-15.  CO<sub>2</sub>

 O<sub>2</sub>

 He

 N<sub>2</sub>

 CH<sub>4</sub>

[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.

CO<sub>2</sub> 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 CO<sub>2</sub> 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 PCl<sub>3</sub> and 45.0 g of PbF<sub>2</sub>. What mass of PbCl<sub>2</sub> can be obtained from the following reaction?  $3 \text{ PbF}_2 + 2 \text{ PCl}_3 \rightarrow 2 \text{ PF}_3 + 3 \text{ PbCl}_2$

$$25.0 \text{ g PCl}_3 \times \frac{1 \text{ mol PCl}_3}{137.33 \text{ g PCl}_3} \times \frac{3 \text{ mol PbCl}_2}{2 \text{ mol PCl}_3} \times \frac{278.1 \text{ g PbCl}_2}{1 \text{ mol PbCl}_2} = 75.9 \text{ g PbCl}_2$$

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

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	( $\Delta H_f^\circ$ ) at 25°C (kJ/mol)
C (graphite)	0.00
$\text{CO}_2$ (g)	-393.5
$\text{H}_2$ (g)	0.00
$\text{H}_2\text{O}$ (l)	-285.85
$\text{O}_2$ (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, ( $\Delta H_f^\circ$ ), 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.)

$$\begin{aligned} P_V &= nRT & P &= \frac{nRT}{V} & n &= 2.000 \text{ g phenol} \times \frac{1 \text{ mol phenol}}{94.12 \text{ g phenol}} = 0.02125 \text{ mol phenol} \\ P &= \frac{0.02125 \text{ mol} (0.0821 \text{ L atm / mol K}) (383 \text{ K})}{10.0 \text{ L}} = 0.601 \text{ atm} \end{aligned}$$

$0.02125 \text{ mol} \times 9 \text{ (moles of products)} = 0.1912 \text{ moles of gaseous products}$

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.

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

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

$$P_V = nRT$$

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