Gaseous Stoichiometry

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1. Diborane, B_2H_6, is a highly explosive compound formed by the reaction

$$\underset{?g}{3NaBH_4 (s)} + 4BF_3 (g) \rightarrow \underset{\substack{1.00L \\ 0°C \\ 1.00\,atm}}{2B_2H_6 (g)} + 3NaBF_4 (s)$$

a. What mass of sodium borohydride, $NaBH_4$, is required to form 1.00L of B_2H_6 at 0° C and 1.00 atm?

$$PV=nRT$$

$$n = \frac{PV}{RT} = \frac{(1.00\,atm)(1.00\,L)}{(.0821)(273\,K)} = .0446 \text{ mol } B_2H_6$$

$$\frac{.0446 \text{ mol } B_2H_6}{} \left| \frac{3 \text{ mol } NaBH_4}{2 \text{ mol } B_2H_6} \right| \frac{37.84\text{ g } NaBH_4}{1 \text{ mol } NaBH_4} = \boxed{2.53\text{ g } NaBH_4}$$

1. Diborane, B_2H_6, is a highly explosive compound formed by the reaction

$$3NaBH_4\ (s) + \underset{\substack{?L \\ 20°C \\ 742\ mmHg}}{4BF_3\ (g)} \rightarrow 2B_2H_6\ (g) + \underset{6.00\ g}{3NaBF_4\ (s)}$$

b. What volume of BF_3 at 20. ° C and 742 mmHg is required to produce 6.00 g $NaBF_4$?

$$\frac{6.00\text{ g NaBF}_4}{} \left| \frac{1\text{ mol NaBF}_4}{109.79\text{ g}} \right| \frac{4\text{ mol BF}_3}{3\text{ mol NaBF}_4} = .0729\text{ mol BF}_3$$

$$PV = nRT$$

$$V = \frac{nRT}{P} = \frac{(.0729\text{ mol})(.0821)(293\text{K})}{\left(742\text{ mmHg}/760\text{ atm}\right)} = \boxed{1.80\text{L BF}_3}$$

2. Calcium reacts with water, yielding hydrogen gas and calcium hydroxide. How many grams of water are required to produce 7.00 L of dry hydrogen at 1.05 atm and 30. °C? (you will need to write and balance the reaction)

$$Ca_{(s)} + \underset{?g}{2H_2O_{(l)}} \longrightarrow \underset{\substack{7.00L \\ 1.05\text{ atm} \\ 30°C = 303K}}{H_{2(g)}} + Ca(OH)_{2(aq)}$$

3. Hydrogen cyanide, HCN, is a poisonous gas that was used in the gas chambers of Hitler's concentration camps. It can be formed by the reaction

$$\underset{?\,g}{NaCN\,(s)} + H^{+}\,(aq) \rightarrow \underset{\substack{8.53\,L \\ 22\,°C \\ 751\,mmHg}}{HCN\,(g)} + Na^{+}\,(aq)$$

What mass of NaCN is required to make 8.53 L of HCN at 22 ° C and 751 mmHg?

$$n_{HCN} = \frac{(751/760\,atm)(8.53L)}{(.0821)(295K)} = .348\ mol\ HCN$$

$$\frac{.348\ mol\ HCN}{} \left| \frac{1\ mol\ NaCN}{1\ mol\ HCN} \right| \frac{49.00\ g\ NaCN}{1\ mol\ NaCN} = \boxed{17.1\ g\ NaCN}$$

4. A gaseous mixture contains 5.78 g of methane, 2.15 g of neon, and 6.80 g of sulfur dioxide. What pressure is exerted by the mixture inside a 75.0L cylinder at 85 ° C? Which gas contributes the greatest pressure? = greatest P = largest # n = $\boxed{CH_4}$

$$\frac{5.78\ g\ CH_4}{} \left| \frac{1\ mol\ CH_4}{16.05\ g} \right. = .360\ mol\ CH_4$$

$$\frac{2.15\ g\ Ne}{} \left| \frac{1\ mol\ Ne}{20.18\ g} \right. = .107\ mol\ Ne$$

$$\frac{6.80\ g\ S}{} \left| \frac{1\ mol\ SO_2}{64.06\ g} \right. = .106\ mol\ SO_2$$

$$n_T = \begin{array}{r} .360 \\ .107 \\ +\ .106 \\ \hline .573\ mol \end{array}$$

$$P_T = \frac{n_T RT}{V} = \frac{(.573\ mol)(.0821)(358K)}{75.0L} = \boxed{.225\ atm}$$

5. To prepare a sample of hydrogen gas a student reacts zinc with hydrochloric acid. The overall net reaction is $Zn (s)\ 2H^{+} (aq) \rightarrow Zn^{+2} (aq) + H_2 (g)$

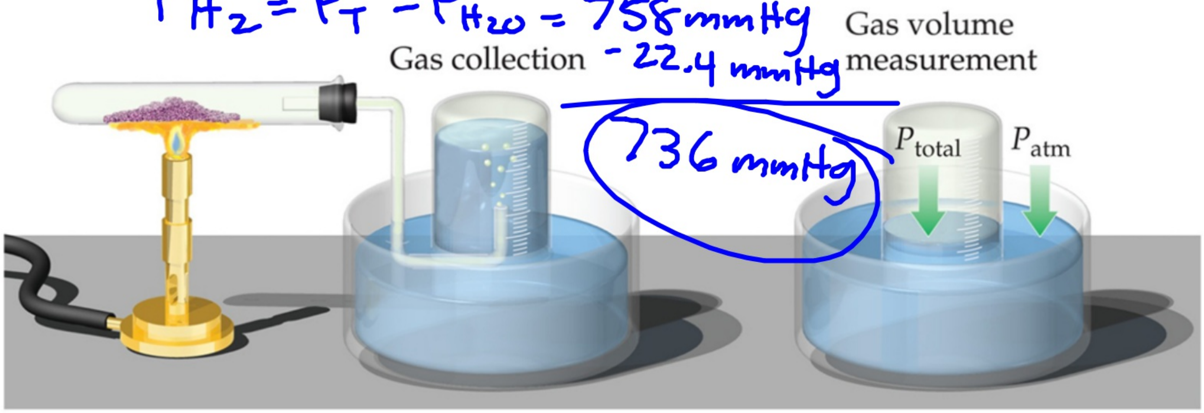
The hydrogen gas is collected over water at 24 ° C and the total pressure is 758 mmHg (vapor pressure of H_2O = 22.4 mmHg).

a. What is the partial pressure of the hydrogen?

$$P_T = P_{H_2O} + P_{H_2}$$

$$P_{H_2} = P_T - P_{H_2O} = 758\text{ mmHg} - 22.4\text{ mmHg}$$

$$\boxed{736\text{ mmHg}}$$



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b. How many grams of hydrogen are there in a 2.00 L sample of wet gas?

$P_{H_2} = 736$ mmHg

$V = 2.00L$

$m = ?$ g

$T = 24°C + 273 = 297K$

$$PV = nRT$$

$$n = \frac{PV}{RT} = \frac{(736/760\text{ atm})(2.00L)}{(.0821)(297\text{ K})}$$

$$n = .0794\text{ mol } H_2 \times \frac{2.02\text{ g } H_2}{1\text{ mol } H_2}$$

$$= \boxed{.160\text{ g } H_2}$$

6. Helium (0.56g) and hydrogen are mixed in a 1.0 L flask at room temperature (298 K). The partial pressure of He is 150mmHg and that of H_2 is 25 mmHg. How many grams of H_2 are present?

$V_T = 1.0\text{ L}$

$T = 298\text{ K}$

$P_{He} = 150\text{ mmHg}$

$P_{H_2} = 25\text{ mmHg}$

$P_T = 150 + 25 = 175\text{ mmHg}$

$m_{He} = 0.56\text{ g}$

$m_{H_2} = ?\text{ g}$

$$\frac{0.56\text{ g He} \,\Big|\, 1\text{ mol He}}{\Big|\, 4.00\text{ g}} = .14\text{ mol He}$$

$$\frac{P_{He}}{P_{H_2}} = \frac{n_{He}}{n_{H_2}}$$

*The Pressure of a gas is directly proportional to # mole @ same T+V

$\uparrow n = \uparrow P$

$$\frac{150\text{ mmHg}}{25\text{ mmHg}} = \frac{.14\text{ mol He}}{n_{H_2}}$$

$$n_{H_2} = \frac{.023\text{ mol } H_2 \,\Big|\, 2.02\text{ g } H_2}{\Big|\, 1\text{ mol } H_2} = \boxed{.047\text{ g } H_2}$$

P_T *Method 1

7. What is the total pressure in atmospheres of a gas mixture that contains 1.0 g of H_2 and 8.0 g of Ar in a 3.0L container at 27 ° C. What are the partial pressures of the two gases? P_{H_2}? P_{Ar}?

$$P_{H_2} = \frac{n_{H_2}RT}{V} = \frac{(.50\text{ mol})(.0821)(300\text{ K})}{3.0\text{ L}} = \boxed{4.1\text{ atm} = P_{H_2}}$$

$$\frac{1.0\text{ g } H_2 \,\Big|\, 1\text{ mol } H_2}{\Big|\, 2.02\text{ g } H_2} = .50\text{ mol } H_2$$

$$P_{Ar} = \frac{n_{Ar}RT}{V} = \frac{(.20\text{ mol})(.0821)(300\text{ K})}{3.0\text{ L}} = \boxed{1.6\text{ atm} = P_{Ar}}$$

$$\frac{8.0\text{ g Ar} \,\Big|\, 1\text{ mol Ar}}{\Big|\, 39.95\text{ g Ar}} = .20\text{ mol Ar}$$

$P_T = P_{H_2} + P_{Ar}$

$P_T = 4.1\text{ atm} + 1.6\text{ atm}$

$= \boxed{5.7\text{ atm}} = P_T$

* Method 2

7. What is the total pressure in atmospheres of a gas mixture that contains 1.0 g of H_2 and 8.0 g of Ar in a 3.0L container at 27 ° C. What are the partial pressures of the two gases?

$$P_T = \frac{n_T RT}{V} = \frac{(.70\text{ mol})(.0821)(300K)}{3.0L} = \boxed{5.7\text{ atm}}$$

$$\begin{array}{l|l} 1.0 g\, H_2 & 1\text{ mol } H_2 \\ \hline & 2.02 g\, H_2 \end{array} = .50\text{ mol } H_2$$

$$\begin{array}{l|l} 8.0 g\, Ar & 1\text{ mol } Ar \\ \hline & 39.95 g\, Ar \end{array} = .20\text{ mol } Ar$$

$$n_T = .50 + .20 = .70\text{ mol Total}$$

*After Finding P_T, find P_{H_2} using PV=nRT, then Subtract $P_T - P_{H_2}$ to find P_{Ar}

8. The hydrocarbon benzene (C_6H_6) burns to give CO_2 and water vapor.

$$\underset{0.095 g}{2\, C_6H_6 (g)} + 15\, O_2 (g) \rightarrow 12\, CO_2 (g) + \underset{\substack{P = ?\text{ atm} \\ 4.75L \\ 30.0°C}}{6\, H_2O (g)}$$

a) If a 0.095g sample of $\overset{=LR}{C_6H_6}$ burns completely in $\overset{\text{excess}}{O_2}$, what is the pressure of water vapor in a 4.75 L flask at 30.0 ° C?

$$\begin{array}{l|l|l} 0.095 g\, C_6H_6 & 1\text{ mol } C_6H_6 & 6\text{ mol } H_2O \\ \hline & 78.12 g & 2\text{ mol } C_6H_6 \end{array} = .0036\text{ mol } H_2O$$

$$P_{H_2O} = \frac{n_{H_2O} RT}{V} = \frac{(.0036\text{ mol})(.0821)(300K)}{4.75L} = \boxed{.019\text{ atm}}$$

8. The hydrocarbon benzene (C_6H_6) burns to give CO_2 and water vapor.

$$2\,C_6H_6\,(g) + 15\,O_2\,(g) \rightarrow 12\,CO_2\,(g) + 6\,H_2O\,(g)$$

0.095 g

P=?atm
4.75 L
22°C

b) If the O_2 gas needed for combustion is contained in a 4.75L flask at 22 °C what is its pressure? P_{O_2}=?atm

$$\frac{0.095\text{ g } C_6H_6}{} \cdot \frac{1\text{ mol } C_6H_6}{78.12\text{ g}} \cdot \frac{15\text{ mol } O_2}{2\text{ mol } C_6H_6} = .0091\text{ mol } O_2$$

$$P_{O_2} = \frac{n_{O_2}RT}{V} = \frac{(.0091\text{ mol } O_2)(.0821)(295\text{ K})}{4.75\text{ L}} = \boxed{.046\text{ atm}}$$

* Method 1

9. Potassium superoxide reacts with CO_2 to give oxygen gas.

$$4\,KO_2\,(s) + 2\,CO_2\,(g) \rightarrow 2\,K_2CO_3\,(s) + 3\,O_2\,(g)$$

16.0 g

P=1.24 atm
T=23°C
V=4.00 L

* These conditions apply only to the GAS not the KO_2 solid

a) If you combine 16.0 g of KO_2 with the CO_2 in a 4.00L tank, in which the gas pressure is 1.24 atm at 23 ° C, which reactant is consumed completely?

aka what is the LR?

$$\frac{16.0\text{ g } KO_2}{} \cdot \frac{1\text{ mol } KO_2}{71.1\text{ g}} \cdot \frac{3\text{ mol } O_2}{4\text{ mol } KO_2} = .169\text{ mol } O_2$$

compare

$$\boxed{\therefore LR = KO_2}$$

$$n_{CO_2} = \frac{(1.24\text{ atm})(4.00\text{ L})}{(.0821)(296\text{ K})} = \frac{.204\text{ mol } CO_2}{} \cdot \frac{3\text{ mol } O_2}{2\text{ mol } CO_2} = .306\text{ mol } O_2$$

* Method 2

9. Potassium superoxide reacts with CO_2 to give oxygen gas.

$$4\,KO_2\,(s) \;+\; 2\,CO_2\,(g) \;\rightarrow\; 2\,K_2CO_3\,(s) \;+\; 3\,O_2\,(g)$$

16.0 g

P=1.24 atm
T=23°C
V=4.00 L

* These conditions apply only to the GAS not the KO_2 solid

a) If you combine 16.0 g of KO_2 with the CO_2 in a 4.00L tank, in which the gas pressure is 1.24 atm at 23° C, which reactant is consumed completely?

aka what is the LR?

The SHORTCUT

*Divide By coefficient

$$\frac{16.0\text{ g }KO_2 \,\Big|\, 1\text{ mol }KO_2}{\Big|\, 71.1\text{ g}} = .225\text{ mol }KO_2 / 4 = .05625$$

compare

$$n_{CO_2} = \frac{(1.24\text{ atm})(4.00\text{L})}{(.0821)(296\text{ K})} = .204\text{ mol}CO_2 / 2 = .102 \quad \boxed{\therefore LR = KO_2}$$

9. Potassium superoxide reacts with CO_2 to give oxygen gas.

$$4\,KO_2\,(s) \;+\; 2\,CO_2\,(g) \;\rightarrow\; 2\,K_2CO_3\,(s) \;+\; 3\,O_2\,(g)$$

P=? atm
2.50L
25°C

b) If the O_2 gas is captured from the reaction, what is its pressure in a 2.50L flask at 25° C?

P_{O_2}=?

From part A: .169 mol O_2 made from the LR

$$P_{O_2} = \frac{(.169\text{ mol})(.0821)(298\text{K})}{2.50\text{L}} = \boxed{1.65\text{ atm}}$$

10. A sample of gas collected over water at 42 ° C occupies a volume of 1.00L. The wet gas has a pressure of 0.986 atm. The gas is dried and the dry gas occupies 1.04 L with a pressure of 1.00 atm at 90. ° C. What is the vapor pressure of the water at 42 ° C?

WET GAS	DRY GAS
Water vapor + gas sample	gas sample, NO water vapor
V=1.00 L T= 42 °C P_{wet}=0.986 atm	V=1.04L P=1.00 atm T=90.°C $n = \frac{(1 \text{ atm})(1.04L)}{(.0821)(363 K)}$ n= .0349 mol gas
$P_{wet\,gas} = P_T = P_{H_2O} + P_{gas}$ $P_{H_2O} = P_T - P_{gas}$ *at 42°C	So at T=42°C when gas was collected OVER H_2O in a 1.0L flask: $P_{gas} = \frac{(.0349 \text{ mol})(.0821)(315 K)}{1.00L}$ P_{gas} = .903 atm
P_{H_2O}=0.986-.903 $\boxed{P_{H_2O} = .083 \text{ atm}}$	

11. Compare and Contrast a real gas and an ideal gas. Explain how to make a real gas behave more like an ideal gas. (details are needed just like when explaining a "good collision")

A real gas has attractions between molecules (intermolecular forces) to some degree and an Ideal gas does not. Real gases exist and Ideal gases do not. To make a real gas behave more like an ideal gas, the IMF's need to be reduced. A high T and large Volume (↓P) would allow molecules to move fast and be farther apart so that attractions aren't recognized as much.

Effusion = gas escapes container through a pin hole

12. In each of the pair of gases below, tell which effuses faster and why:

a) CO_2 and F_2 b) O_2 and N_2 c) C_2H_4 and B_2H_6

a) $MM_{CO_2} = 44.01$

$MM_{F_2} = 38$

$\therefore$ F_2 effuses faster b/c it is lighter

b) $MM_{O_2} = 32$

$MM_{N_2} = 28.02$

$\therefore$ N_2 effuses faster

c) $MM_{C_2H_4} = 28.06$

$MM_{B_2H_6} = 27.68$

$\therefore$ B_2H_6 effuses faster

13. Argon gas is ten times as dense as helium gas at the same temperature and pressure. Which gas effuses faster? How much faster?

Graham's Law

$$\frac{\text{rate He}}{\text{rate Ar}} = \sqrt{\frac{MM_{Ar}}{MM_{He}}} = \sqrt{\frac{39.95}{4.00}} = \sqrt{10} = 3.16$$

He effuses 3.16 times faster than Ar at the same T and P.