Chemistry 106, Exam 3. January 23, 2003.

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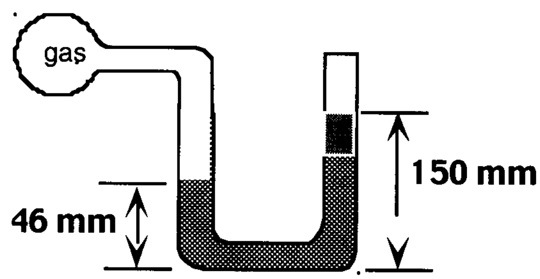
Useful information: PV=nRT, R = 0.0821 L•atm/mol•K = 8.314 J/mol•K, J = kgm^2/s^2,

$u = \sqrt{\frac{3RT}{MW}}$, $\frac{u_1}{u_2} = \sqrt{\frac{MW_2}{MW_1}}$, 1 atm = 760 mm Hg = 101 kPa   van der Waals Eqn

(1) The height of the mercury in the right arm open to atmospheric pressure (760 mmHg) is ~~100~~ 150 mm and the height in the left arm is ~~120~~ 46 mm.



What is the pressure of the gas in the bulb?

(A) 760 mmHg   (C) 656 mmHg

(B) 864 mmHg   (D) 103 mmHg

(2) At constant volume, the pressure of gas **Y** increases with increasing temperature because as the temperature increases,

(A) molecules of **Y** move faster.

(B) the molecular volume of **Y** increases.

(C) the mass of **Y** molecules increases.

(D) molecular collisions are more elastic.

(E) molecules are closer together.

(3) A student collected 40 mL of H_2 gas when the temperature was 20 °C and the pressure was 720 mmHg. The next day the temperature was 20 °C, but he had only 38.4 mL of gas. The new pressure is

(A)	691 mmHg	(D)	750 mmHg
(B)	700 mmHg	(E)	760 mmHg
(C)	721 mmHg		

$(0.947\,atm)(0.040\,L) = n(0.0821\,\frac{L\,atm}{mol\,K})(293\,K)$

$1.57\times10^{-3}\,mol = n$

$720\,mmHg = 0.947\,atm = P_1$

$V_1 = 0.040\,L$

$T = 293\,K$

$P(0.0384\,L) = (1.57\times10^{-3}\,mol)(0.0821\,\frac{L\,atm}{mol\,K})(293\,K)$

$P = 0.986\,atm = 750\,mmHg$

(4) When equal masses of methane gas and oxygen are added to an empty container at 25 °C, the fraction of the pressure exerted by the oxygen is

	Molar Masses
CH_4	16. g·mol^{-1}
O_2	32. g·mol^{-1}

32 g of CH_4 = 2 moles

32 g of O_2 = 1 mole

$X_{O_2} = \frac{1\,mol}{3\,moles} = 0.33$

(A) 0.16 total pressure.

(B) 0.33 total pressure.

(C) 0.50 total pressure.

(D) 0.67 total pressure

(5) Given a mixture of gases: 15.0 g CH_4; 7.00 g N_2; and 0.500 g H_2. What is the total pressure at 20 °C if confined to a 1.000 L container?

	Molar Masses
CH_4	16.0 $g{\cdot}mol^{-1}$
N_2	28.0 $g{\cdot}mol^{-1}$
H_2	2.01 $g{\cdot}mol^{-1}$

$15.0\,g \div 16.0\,g/mol = 0.938\,moles\ CH_4$

$7.00\,g \div 28.0\,g/mol = 0.250\,moles\ N_2$

$0.500\,g \div 2.01\,g/mol = 0.249\,moles\ H_2$

total $1.44\,moles\ gas$

(A) 21.8 atm

(B) 34.5 atm

(C) 10.9 atm

(D) 4.85 atm

$T = 293\,K$

$V = 1.00\,L$

$P(1.00\,L) = (1.44\,moles)(0.0821\,\frac{L\,atm}{mol\,K})(293\,K)$

$P = 34.5\,atm$

(6) Nitrogen at room temperature is a more nearly ideal gas than sulfur dioxide. Which is the best ***theoretical*** explanation for this?

(A) Nitrogen is a element; sulfur dioxide is a compound.

(B) Nitrogen contains two atoms per molecule, while sulfur dioxide contains three atoms per molecule.

(C) Nitrogen is quite inert, while sulfur dioxide is fairly reactive.

(D) Sulfur dioxide may be more readily liquefied than nitrogen.

(E) The molecules of sulfur dioxide have a greater attraction for one another than the molecules of nitrogen have for one another.

(7) List Four Postulates of the Kinetic Molecular Theory of Gases

(8) The US Navy built an experimental undersea habitat that had an atmosphere composed of 79.0 % He, 17% Ne, and 4.0 % O_2 (by mole).

(a) What is the partial pressure of each of these gases 59 m below the surface of the water (P = 6.91 atm)?

$P_{O_2} = X_{O_2} P_{tot}$

~~0.0404~~

$P_{O_2} = (0.040)(6.91 \text{ atm}) = 0.276 \text{ atm}$

$P_{He} = (0.790)(6.91 \text{ atm}) = 5.46 \text{ atm}$

$P_{Ne} = (0.17)(6.91 \text{ atm}) = 1.17 \text{ atm}$

(b) How does the P_{O2} above compare to the partial pressure of oxygen (20 mole%) in the atmosphere (1 atm)?

$P_{O_2} = (0.20)(1 \text{ atm}) = 0.20 \text{ atm}$

In the undersea habitat, it is slightly higher.

(9) What is the pressure of water at 150 °C if 1.00 mole is in a 35.0 L container? Use both the Van der Waal's equation and the ideal gas law. Which pressure is higher? (a = 5.537 $L^2 \bullet atm/mol^2$, b = 0.03049 L/mol for water).

$\left(P + \frac{n^2 a}{V^2}\right)(V - nb) = nRT \quad T = 150\,^\circ C = 423\,K$

$\left(P + \frac{(1\,mole)^2 (5.537\,\frac{L^2 \cdot atm}{mole^2})}{(35.0\,L)^2}\right)\left(35.0\,L - (1.00\,mole)(0.03049\,\frac{L}{mole}\right) = (1.00\,mole)(0.0821\,\frac{L\,atm}{mole\,K})(423\,K)$

$(P + 4.52\times10^{-3}\,atm)(35.0\,L - 0.03049\,L) = 34.7\,L\,atm$

$(P + 4.52\times10^{-3}\,atm)(35.0\,L) = 34.7\,L\,atm$

$P + 4.52\times10^{-3}\,atm = 0.991\,atm$

$\boxed{P = 0.987\,atm}$ Van der Waals

ideal $P(35.0\,L) = (1.00\,mole)(0.0821\,\frac{L\,atm}{mole\,K})(423\,K)$

$P = 0.992\,atm$

The ideal gas pressure is higher

(10) The Union Carbide company's fortune was founded on the reaction of calcium carbide and water to form acetylene gas (C_2H_2).

$CaC_{2(s)} + H_2O_{(l)} \text{-----}> C_2H_{2(g)} + Ca(OH)_{2(s)}$

If 25.0 g of ~~acetylene~~ is added to 200 mL of water at 20 °C and 1.00 atm of pressure, what volume of acetylene is formed?

Calcium Carbide

FW = 64.10 g/mole

$25.0\,g \div 64.10\,g/mole = 0.390\,moles$

$(1.00\,atm)\,V = (0.390\,moles)(0.0821\,\frac{L\,atm}{mole\,K})(293\,K)$

$\boxed{V = 9.38\,L}$

(11) What is the density of $CHCl_3$ at 529 mm Hg pressure and a temperature of 273 K (give density in g/L)?

$CHCl_3$ FW = 119.4 g/mole

assm 1.00 L at 273 K

$\frac{529 \text{ mm Hg}}{760 \text{ mmHg/atm}} = 0.696 \text{ atm}$

$(0.696 \text{ atm})(1.00 \text{ L}) = n (0.0821 \frac{\text{L atm}}{\text{mole K}})(273 \text{ K})$

$0.0311 \text{ moles} = n$

$0.0311 \text{ moles} \times 119.4 \text{ g/mole} = 3.71 \text{ g}$

$\frac{0.0311 \text{ moles}}{\text{L}} = \boxed{\frac{3.71 \text{ g}}{\text{L}}}$

(12) The speed of light (c) is 3.00 x 10^8 m/s. What temperature would would be required for the average H_2 molecule to have a velocity of 0.5 c?

$0.5 c = 0.5(3.00 \times 10^8 \text{ m/s}) = 1.50 \times 10^8 \text{ m/s}$

$MW = 2.02 \text{ g/mole} = 0.00202 \text{ kg/mole}$

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

$(1.50 \times 10^8 \text{ m/s})^2 = \left(\sqrt{\frac{3(8.314 \text{ J/mole K})(T)}{0.00202 \text{ kg/mole}}}\right)^2$

$2.25 \times 10^{16} \frac{m^2}{s^2} = \frac{3(8.314 \text{ J/mole K})T}{0.00202 \text{ kg/mole}}$

$2.25 \times 10^{16} \frac{m^2}{s^2} = (1.234 \times 10^4 \frac{m^2}{s^2 K})T$

$\boxed{1.82 \times 10^{12} \text{ K} = T}$

(Extra Credit) Methane, CH_4, diffuses in a given apparatus at the rate of 30 mL·min^{-1}. At what rate would a gas with a molar mass of 100 diffuse under the same conditions?

CH_4 MW = 16.05 g/mole = 0.01605 kg/mole

	Molar Mass
CH_4	16. g·mol^{-1}

(A) 0.77 mL·min^{-1}

(B) 6.7 mL·min^{-1}

(C) 12 mL·min^{-1}

(D) 30 mL·min^{-1}

(E) 75 mL·min^{-1}

$\frac{u_{100}}{u_{CH_4}} = \sqrt{\frac{0.0165 \text{ kg/mole}}{0.100 \text{ kg/mole}}}$

$\frac{u_{100}}{u_{CH_4}} = 0.406$ (slower)

$0.406 \times \frac{30 \text{ mL}}{\text{min}} = 12 \text{ mL/min}$