Chapter 5 Suggested end-of chapter problems with solutions

^{21.} If a barometer were built using water $(d = 1.0 \text{ g/cm}^3)$ instead of mercury $(d = 13.6 \text{ g/cm}^3)$, would the column of water be higher than, lower than, or the same as the column of mercury at **1.00 atm**? If the level is different, by what factor? Explain.

- Answer: The column of water would have to be 13.6 times taller than a column of mercury. When the pressure of the column of liquid standing on the surface of the liquid is equal to the pressure of air on the rest of the surface of the liquid, then the height of the column of liquid is a measure of atmospheric pressure. Because water is 13.6 times less dense than mercury, the column of water must be 13.6 times longer than that of mercury to match the force exerted by the columns of liquid standing on the surface.
 - 23. Boyle's law can be represented graphically in several ways. Which of the following plots does *not* correctly represent Boyle's law (assuming constant *T* and *n*)? Explain.



Answer:

- The P versus 1/V plot is incorrect. The plot should be linear with <u>positive</u> slope and a *y*-intercept of zero. PV = k, so P = k(1/V). This is in the form of the straight-line equation y = mx + b. The *y*-axis is pressure, and the *x*-axis is 1/V.
- 29. Do all the molecules in a **1-mole** sample of $CH_4(g)$ have the same kinetic energy at **273 K**? Do all molecules in a **1-mole** sample of $N_2(g)$ have the same velocity at **546 K**? Explain.

Answer:

No; at any nonzero Kelvin temperature, there is a distribution of kinetic energies. Similarly, there is a distribution of velocities at any nonzero Kelvin temperature. The reason there is a distribution of kinetic energies at any specific temperature is because there is a distribution of velocities for any gas sample at any specific temperature.

- 31. As $\mathbf{NH}_3(g)$ is decomposed into nitrogen gas and hydrogen gas at constant pressure and temperature, the volume of the product gases collected is twice the volume of \mathbf{NH}_3 reacted. Explain. As $\mathbf{NH}_3(g)$ is decomposed into nitrogen gas and hydrogen gas at constant volume and temperature, the total pressure increases by some factor. Why the increase in pressure and by what factor does the total pressure increase when reactants are completely converted into products? How do the partial pressures of the product gases compare to each other and to the initial pressure of \mathbf{NH}_3 ?
- Answer: 2 NH₃(g) \rightarrow N₂(g) + 3 H₂(g); as reactants are converted into products, we go from 2 moles of gaseous reactants to 4 moles of gaseous products (1 mol N₂ + 3 mol H₂). Because the moles of gas doubles as reactants are converted into products, the volume of the gases will double (at constant P and T).

$$PV = nRT$$
, $P = \left(\frac{RT}{V}\right)n = (constant)n$; pressure is directly related to n at constant T and V.

As the reaction occurs, the moles of gas will double, so the pressure will double. Because 1 mole of N₂ is produced for every 2 moles of NH₃ reacted, $P_{N_2} = (1/2)P_{NH_3}^o$. Owing to the 3 : 2 mole ratio in the balanced equation, $P_{H_2} = (3/2)P_{NH_3}^o$.

Note: $P_{total} = P_{H_2} + P_{N_2} = (3/2)P_{NH_3}^o + (1/2)P_{NH_3}^o = 2P_{NH_3}^o$. As we said earlier, the total pressure will double from the initial pressure of NH₃ as reactants are completely converted into products.

33. From the values in Table 5.3 for the van der Waals constant *a* for the gases
H₂, CO₂, N₂, and CH₄, predict which of these gas molecules show the strongest intermolecular attractions. Not needed this semester

Answer: The values of *a* are: H₂,
$$\frac{0.244 \text{ atm } \text{L}^2}{\text{mol}^2}$$
; CO₂, 3.59; N₂, 1.39; CH₄, 2.25

Because a is a measure of intermolecular attractions, the attractions are greatest for CO₂.

34. Without looking at a table of values, which of the following gases would you expect to have the largest value of the van der Waals constant *b*: H₂, N₂, CH₄, C₂H₆, or C₃H₈? Not needed this semester

Answer:

The van der Waals constant b is a measure of the size of the molecule. Thus C₃H₈ should have the largest value of b because it has the largest molar mass (size).

41. A diagram for an open-tube manometer is shown below.



If the flask is open to the atmosphere, the mercury levels are equal. For each of the following situations where a gas is contained in the flask, calculate the pressure in the flask in torr, atmospheres, and pascals.



c. Calculate the pressures in the flask in parts a and b (in torr) if the atmospheric pressure is **635 torr**.

Answer:

If the levels of mercury in each arm of the manometer are equal, then the pressure in the flask is equal to atmospheric pressure. When they are unequal, the difference in height in millimeters will be equal to the difference in pressure in millimeters of mercury between the flask and the atmosphere. Which level is higher will tell us whether the pressure in the flask is less than or greater than atmospheric.

a.
$$P_{\text{flask}} < P_{\text{atm}}$$
; $P_{\text{flask}} = 760. - 118 = 642$ torr

$$642 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 0.845 \text{ atm}$$

$$0.845 \text{ atm} \times \frac{1.013 \times 10^5 \text{ Pa}}{\text{atm}} = 8.56 \times 10^4 \text{ Pa}$$

b. $P_{\text{flask}} > P_{\text{atm}}$; $P_{\text{flask}} = 760$. torr + 215 torr = 975 torr

975 torr
$$\times \frac{1 \text{ atm}}{760 \text{ torr}} = 1.28 \text{ atm}$$

$$1.28 \text{ atm} \times \frac{1.013 \times 10^5 \text{ Pa}}{\text{atm}} = 1.30 \times 10^5 \text{ Pa}$$

c. $P_{\text{flask}} = 635 - 118 = 517 \text{ torr}; P_{\text{flask}} = 635 + 215 = 850. \text{ torr}$

42.

a. If the open-tube manometer in Exercise 41 contains a non-volatile

silicone oil $\left(\text{density} = 1.30 \text{ g/cm}^3 \right)$ instead of mercury

 $(\text{density} = 13.6 \text{ g/cm}^3)$, what are the pressures in the flask as shown in

parts a and b in torr, atmospheres, and pascals?

- b. What advantage would there be in using a less dense fluid than mercury in a manometer used to measure relatively small differences in pressure?
- Answer: _______ a. The pressure is proportional to the mass of the fluid. The mass is proportional to the volume of the column of fluid (or to the height of the column assuming the area of the column of fluid is constant).

 $d = density = \frac{mass}{volume}$; in this case, the volume of silicon oil will be the same as the volume of mercury in Exercise 41.

$$V = \frac{m}{d}$$
; $V_{Hg} = V_{oil}$; $\frac{m_{Hg}}{d_{Hg}} = \frac{m_{oil}}{d_{oil}}$, $m_{oil} = \frac{m_{Hg}d_{oil}}{d_{Hg}}$

Because P is proportional to the mass of liquid:

$$P_{oil} = P_{Hg}\left(\frac{d_{oil}}{d_{Hg}}\right) = P_{Hg}\left(\frac{1.30}{13.6}\right) = (0.0956)P_{Hg}$$

This conversion applies only to the column of silicon oil.

 $P_{\text{flask}} = 760. \text{ torr} - (0.0956 \times 118) \text{ torr} = 760. - 11.3 = 749 \text{ torr}$

749 torr ×
$$\frac{1 \text{ atm}}{760 \text{ torr}}$$
 = 0.986 atm; 0.986 atm × $\frac{1.013 \times 10^{3} \text{ Pa}}{\text{ atm}}$ = 9.99 × 10⁴ Pa

 $P_{flask} = 760. \text{ torr} + (0.0956 \times 215) \text{ torr} = 760. + 20.6 = 781 \text{ torr}$

781 torr
$$\times \frac{1 \text{ atm}}{760 \text{ torr}} = 1.03 \text{ atm}; \ 1.03 \text{ atm} \times \frac{1.013 \times 10^5 \text{ Pa}}{\text{ atm}} = 1.04 \times 10^5 \text{ Pa}$$

b. If we are measuring the same pressure, the height of the silicon oil column would be $13.6 \div 1.30 = 10.5$ times the height of a mercury column. The advantage of using a less dense fluid than mercury is in measuring small pressures. The height difference measured will be larger for the less dense fluid. Thus the measurement will be more precise.

43. A particular balloon is designed by its manufacturer to be inflated to a volume of no more than 2.5 L. If the balloon is filled with 2.0 L helium at sea level, is released, and rises to an altitude at which the atmospheric pressure is only 500. mm Hg, will the balloon burst? (Assume temperature is constant.)

Answer:

At constant n and T, PV = nRT = constant, so $P_1V_1 = P_2V_2$; at sea level, P = 1.00 atm = 760. mm Hg.

$$V_2 = \frac{P_1 V_1}{P_2} = \frac{760. \text{ mm} \times 2.0 \text{ L}}{500. \text{ mm Hg}} = 3.0 \text{ L}$$

The balloon will burst at this pressure because the volume must expand beyond the 2.5 L limit of the balloon.

Note: To solve this problem, we did not have to convert the pressure units into atm; the units of mm Hg canceled each other. In general, only convert units if you have to. Whenever the gas constant R is not used to solve a problem, pressure and volume units must only be consistent and not necessarily in units of atm and L. The exception is temperature, which must <u>always</u> be converted to the Kelvin scale.

- 50. The average lung capacity of a human is **6.0 L**. How many moles of air are in your lungs when you are in the following situations?
- Answer:

a. At sea level (T = 298 K, P = 1.00 atm).

b. 10. m below water (T = 298 K, P = 1.97 atm).

c. At the top of Mount Everest (T = 200. K, P = 0.296 atm).

a.
$$n = \frac{PV}{RT} = \frac{1.00 \text{ atm} \times 6.0 \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 298 \text{ K}} = 0.25 \text{ mol air}$$

b. $n = \frac{1.97 \text{ atm} \times 6.0 \text{ L}}{0.08206 \text{ L atm/K mol} \times 298 \text{ K}} = 0.48 \text{ mol air}$
c. $n = \frac{0.296 \text{ atm} \times 6.0 \text{ L}}{0.08206 \text{ L atm/K mol} \times 200. \text{ K}} = 0.11 \text{ mol air}$

Air is indeed "thinner" at high elevations.

51. The steel reaction vessel of a bomb calorimeter, which has a volume of 75.0 mL, is charged with oxygen gas to a pressure of 14.5 atm at 22°C. Calculate the moles of oxygen in the reaction vessel.

Answer:

PV = nRT,
$$n = \frac{PV}{RT} = \frac{14.5 \text{ atm} \times (75.0 \times 10^{-3} \text{ L})}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 295 \text{ K}} = 0.0449 \text{ mol } \text{O}_2$$

58. What will be the effect on the volume of an ideal gas if the pressure is doubled and the absolute temperature is halved?

Answer:

The pressure is doubled so $P_2 = 2P_1$ and the absolute temperature is halved so $T_2 = \frac{1}{2}T_1$ (or $T_1 = 2T_2$). The moles of gas did not change, so $n_2 = n_1$. The volume effect of these changes is:

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}, \ \frac{V_2}{V_1} = \frac{P_1n_2T_2}{P_2n_1T_1} = \frac{P_1T_2}{P_2T_1} = \frac{P_1 \times T_2}{2P_1 \times 2T_2} = 1/4$$

The volume of the gas decreases by a factor of four when the pressure is doubled and the absolute temperature is halved.

61. An ideal gas is contained in a cylinder with a volume of 5.0 × 10² mL at a temperature of 30.°C and a pressure of 710. torr. The gas is then compressed to a volume of 25 mL, and the temperature is raised to 820.°C. What is the new pressure of the gas?

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er:
$$\frac{PV}{T} = nR = constant, \quad \frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$P_2 = \frac{P_1 V_1 T_2}{V_2 T_1} = 710. \text{ torr} \times \frac{5.0 \times 10^2 \text{ mL}}{25 \text{ mL}} \times \frac{(273 + 820.) \text{ K}}{(273 + 30.) \text{ K}} = 5.1 \times 10^4 \text{ torr}$$

65. Consider the following reaction:

$$4\mathrm{Al}(s) + 3\mathrm{O}_2(g)
ightarrow 2\mathrm{Al}_2\mathrm{O}_3(s)$$

It takes **2.00 L** of pure oxygen gas at STP to react completely with a certain sample of aluminum. What is the mass of aluminum reacted?

Answer: STP: T = 273 K and P = 1.00 atm; at STP, the molar volume of a gas is 22.42 L.

$$2.00 \text{ L } \text{O}_2 \times \frac{1 \text{ mol } \text{O}_2}{22.42 \text{ L}} \times \frac{4 \text{ mol } \text{Al}}{3 \text{ mol } \text{O}_2} \times \frac{26.98 \text{ g } \text{Al}}{\text{ mol } \text{Al}} = 3.21 \text{ g } \text{Al}$$

Note: We could also solve this problem using PV = nRT, where $n_{O_2} = PV/RT$. You don't have to memorize 22.42 L/mol at STP.

66. A student adds **4.00** g of dry ice (solid **CO**₂) to an empty balloon. What will be the volume of the balloon at STP after all the dry ice sublimes (converts to gaseous **CO**₂)?

Answer:

$$CO_2(s) \rightarrow CO_2(g); 4.00 \text{ g } CO_2 \times \frac{1 \text{ mol } CO_2}{44.01 \text{ g } CO_2} = 9.09 \times 10^{-2} \text{ mol } CO_2$$

At STP, the molar volume of a gas is 22.42 L. $9.09 \times 10^{-2} \text{ mol CO}_2 \times \frac{22.42 \text{ L}}{\text{mol CO}_2} = 2.04 \text{ L}$

1 0 0

Ethene is converted to ethane by the reaction **Not needed this semester**

$$\mathrm{C_2H_4}(g) + \mathrm{H_2}(g) \overset{\mathrm{Catalyst}}{\longrightarrow} \mathrm{C_2H_6}(g)$$

 C_2H_4 flows into a catalytic reactor at 25.0 atm and 300.°C with a flow rate of 1000. L/min. Hydrogen at 25.0 atm and 300.°C flows into the reactor at a flow rate of 1500. L/min. If 15.0 kg C_2H_6 is collected per minute, what is the percent yield of the reaction?

Answer:

From the balanced equation, ethene reacts with hydrogen in a 1:1 mole ratio. Because T and P are constant, a greater volume of H₂ and thus more moles of H₂ are flowing into the reaction container than moles of ethene. So ethene is the limiting reagent.

In 1 minute:

$$n_{C_{2}H_{4}} = \frac{PV}{RT} = \frac{25.0 \text{ atm} \times 1000. \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 573 \text{ K}} = 532 \text{ mol } C_{2}H_{4} \text{ reacted}$$

Theoretical yield =
$$\frac{532 \operatorname{mol} C_2 H_4}{\operatorname{min}} \times \frac{1 \operatorname{mol} C_2 H_6}{\operatorname{mol} C_2 H_4} \times \frac{30.07 \operatorname{g} C_2 H_6}{\operatorname{mol} C_2 H_6} \times \frac{1 \operatorname{kg}}{1000 \operatorname{g}}$$

= 16.0 kg C₂H₆/min

Percent yield = $\frac{15.0 \text{ kg/min}}{16.0 \text{ kg/min}} \times 100 = 93.8\%$

77. An unknown diatomic gas has a density of **3.164 g/L** at STP. What is the identity of the gas?

Answer: Molar mass = $\frac{dRT}{P}$, where d = density of gas in units of g/L. Molar mass = $\frac{3.164 \text{ g/L} \times \frac{0.08206 \text{ L atm}}{\text{K mol}} \times 273.2 \text{ K}}{1.000 \text{ atm}} = 70.98 \text{ g/mol}$

The gas is diatomic, so the average atomic mass = 70.93/2 = 35.47 u. From the periodic table, this is chlorine, and the identity of the gas is Cl₂.

87. Consider the three flasks in the diagram below. Assuming the connecting tubes have negligible volume, what is the partial pressure of each gas and the total pressure after all the stopcocks are opened?



- 89. The partial pressure of $CH_4(g)$ is 0.175 atm and that of $O_2(g)$ is 0.250 atm in a mixture of the two gases.
 - a. What is the mole fraction of each gas in the mixture?
 - b. If the mixture occupies a volume of **10.5 L** at **65°C**, calculate the total number of moles of gas in the mixture.
 - c. Calculate the number of grams of each gas in the mixture.

Answer: a. Mole fraction
$$CH_4 = \chi_{CH_4} = \frac{P_{CH_4}}{P_{total}} = \frac{0.175 \text{ atm}}{0.175 \text{ atm} + 0.250 \text{ atm}} = 0.412$$

$$\chi_{O_2} = 1.000 - 0.412 = 0.588$$

b.
$$PV = nRT$$
, $n_{total} = \frac{P_{total} \times V}{RT} = \frac{0.425 \text{ atm} \times 10.5 \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 338 \text{ K}} = 0.161 \text{ mol}$

c. $\chi_{CH_4} = \frac{n_{CH_4}}{n_{total}}, n_{CH_4} = \chi_{CH_4} \times n_{total} = 0.412 \times 0.161 \text{ mol} = 6.63 \times 10^{-2} \text{ mol CH}_4$ $6.63 \times 10^{-2} \text{ mol CH}_4 \times \frac{16.04 \text{ g CH}_4}{\text{mol CH}_4} = 1.06 \text{ g CH}_4$

 $n_{O_2} = 0.588 \times 0.161 \text{ mol} = 9.47 \times 10^{-2} \text{ mol } O_2; 9.47 \times \text{mol } O_2 \times \frac{32.00 \text{ g } O_2}{\text{mol } O_2} = 3.03 \text{ g } O_2$

91. Small quantities of hydrogen gas can be prepared in the laboratory by the addition of aqueous hydrochloric acid to metallic zinc.

$$\mathrm{Zn}(s) + 2\mathrm{HCl}(aq)
ightarrow \mathrm{ZnCl}_2(aq) + \mathrm{H}_2(g)$$

Typically, the hydrogen gas is bubbled through water for collection and becomes saturated with water vapor. Suppose **240. mL** of hydrogen gas is collected at **30.°C** and has a total pressure of **1.032 atm** by this process. What is the partial pressure of hydrogen gas in the sample? How many grams of zinc must have reacted to produce this quantity of hydrogen? (The vapor pressure of water is **32 torr** at **30°C**.)

Answer:
$$P_{\text{total}} = P_{\text{H}_2} + P_{\text{H}_2\text{O}}$$
, 1.032 atm $= P_{\text{H}_2} + 32 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}}$, $P_{\text{H}_2} = 1.032 - 0.042 = 0.990 \text{ atm}$
 $n_{\text{H}_2} = \frac{P_{\text{H}_2} \text{V}}{\text{RT}} = \frac{0.990 \text{ atm} \times 0.240 \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 303 \text{ K}} = 9.56 \times 10^{-3} \text{ mol H}_2$
 $9.56 \times 10^{-3} \text{ mol H}_2 \times \frac{1 \text{ mol Zn}}{\text{mol H}_2} \times \frac{65.38 \text{ g Zn}}{\text{mol Zn}} = 0.625 \text{ g Zn}$

100. The oxides of Group 2A metals (symbolized by **M** here) react with carbon dioxide according to the following reaction:

$$MO(s) + CO_2(g) \rightarrow MCO_3(s)$$

A 2.85-g sample containing only MgO and CuO is placed in a 3.00-L container. The container is filled with CO₂ to a pressure of 740. torr at 20.°C. After the reaction has gone to completion, the pressure inside the flask is 390. torr at 20.°C. What is the mass percent of MgO in the mixture? Assume that only the MgO reacts with CO₂.

Answer:

The partial pressure of CO_2 that reacted is 740. – 390. = 350. torr. Thus the number of moles of CO_2 that react is given by:

$$n = \frac{PV}{RT} = \frac{\frac{350.}{760} \text{ atm} \times 3.00 \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 293 \text{ K}} = 5.75 \times 10^{-2} \text{ mol CO}_2$$

5.75 × 10⁻² mol CO₂ × $\frac{1 \text{ mol MgO}}{1 \text{ mol CO}_2}$ × $\frac{40.31 \text{ g MgO}}{\text{mol MgO}} = 2.32 \text{ g MgO}$
Mass % MgO = $\frac{2.32 \text{ g}}{2.85 \text{ g}} \times 100 = 81.4\% \text{ MgO}$

101. Calculate the average kinetic energies of $CH_4(g)$ and $N_2(g)$ molecules at 273 K and 546 K. Not needed this semester

Answer:

 $KE_{avg} = (3/2)RT$; the average kinetic energy depends only on temperature. At each temperature, CH_4 and N_2 will have the same average KE. For energy units of joules (J), use R = 8.3145 J/K•mol. To determine average KE per molecule, divide the molar KE_{avg} by Avogadro's number, 6.022×10^{23} molecules/mol.

At 273 K:
$$KE_{avg} = \frac{3}{2} \times \frac{8.3145 \text{ J}}{\text{K mol}} \times 273 \text{ K} = 3.40 \times 10^3 \text{ J/mol} = 5.65 \times 10^{-21} \text{ J/molecule}$$

At 546 K:
$$KE_{avg} = \frac{3}{2} \times \frac{8.3145 \text{ J}}{\text{K mol}} \times 546 \text{ K} = 6.81 \times 10^3 \text{ J/mol} = 1.13 \times 10^{-20} \text{ J/molecule}$$

102. Calculate the root mean square velocities of $CH_4(g)$ and $N_2(g)$ molecules at 273 K and 546 K. Not needed this semester

Answer:
$$\mu_{\rm rms} = \left(\frac{3\,{\rm RT}}{{\rm M}}\right)^{1/2}$$
, where ${\rm R} = \frac{8.3145\,{\rm J}}{{\rm K\,mol}}$ and ${\rm M} = {\rm molar\ mass\ in\ kg}$.
For CH₄, ${\rm M} = 1.604 \times 10^{-2}\,{\rm kg}$, and for N₂, ${\rm M} = 2.802 \times 10^{-2}\,{\rm kg}$.
For CH₄ at 273 K: $\mu_{\rm rms} = \left(\frac{3 \times \frac{8.3145\,{\rm J}}{{\rm K\,mol}} \times 273\,{\rm K}}{1.604 \times 10^{-2}\,{\rm kg/mol}}\right)^{1/2} = 652\,{\rm m/s}$

Similarly, μ_{rms} for CH₄ at 546 K is 921 m/s.

For N₂ at 273 K:
$$\mu_{\rm rms} = \left(\frac{3 \times \frac{8.3145 \text{ J}}{\text{K mol}} \times 273 \text{ K}}{2.802 \times 10^{-2} \text{ kg/mol}}\right)^{1/2} = 493 \text{ m/s}$$

Similarly, for N₂ at 546 K, $\mu_{rms} = 697$ m/s.

- 107. Consider a **1.0-L** container of neon gas at STP. Will the average kinetic energy, average velocity, and frequency of collisions of gas molecules with the walls of the container increase, decrease, or remain the same under each of the following conditions?
 - a. The temperature is increased to $100^\circ C$.
 - b. The temperature is decreased to $-50^\circ C$.
 - c. The volume is decreased to 0.5~L .
 - d. The number of moles of neon is doubled.

Answer:		a	b	с	d
	Avg. KE	increase	decrease	same (KE \propto T)	same
	Avg. velocity	increase	decrease	same $(\frac{1}{2} \text{ mv}^2 = \text{KE} \propto \text{T})$	same
	Wall coll. freq	increase	decrease	increase	increase

Average kinetic energy and average velocity depend on T. As T increases, both average kinetic energy and average velocity increase. At constant T, both average kinetic energy and average velocity are constant. The collision frequency is proportional to the average velocity (as velocity increases, it takes less time to move to the next collision) and to the quantity n/V (as molecules per volume increase, collision frequency increases).

111. Freon-12 is used as a refrigerant in central home air conditioners. The rate of effusion of Freon-12 to Freon-11 (molar mass = 137.4 g/mol) is 1.07:1. The formula of Freon-12 is one of the following: CF_4 , CF_3Cl , CF_2Cl_2 , $CFCl_3$, or CCl_4 . Which formula is correct for Freon-12?

Answer:

Graham's law of effusion:
$$\frac{\text{Rate}_1}{\text{Rate}_2} = \left(\frac{M_2}{M_1}\right)^{1/2}$$

Let Freon-12 = gas 1 and Freon-11 = gas 2:

$$\frac{1.07}{1.00} = \left(\frac{137.4}{M_1}\right)^{1/2}, \ 1.14 = \frac{137.4}{M_1}, \ M_1 = 121 \text{ g/mol}$$

The molar mass of CF_2Cl_2 is equal to 121 g/mol, so Freon-12 is CF_2Cl_2 .

112. The rate of effusion of a particular gas was measured and found to be 24.0 mL/min . Under the same conditions, the rate of effusion of pure methane (CH₄) gas is 47.8 mL/min . What is the molar mass of the unknown gas?

Answer:
$$\frac{\text{Rate}_1}{\text{Rate}_2} = \left(\frac{M_2}{M_1}\right)^{1/2}$$
; $\text{rate}_1 = \frac{24.0 \text{ mL}}{\text{min}}$; $\text{rate}_2 = \frac{47.8 \text{ mL}}{\text{min}}$; $M_2 = \frac{16.04 \text{ g}}{\text{mol}}$; $M_1 = ?$
 $\frac{24.0}{47.8} = \left(\frac{16.04}{M_1}\right)^{1/2} = 0.502$, $16.04 = (0.502)^2 \times M_1$, $M_1 = \frac{16.04}{0.252} = \frac{63.7 \text{ g}}{\text{mol}}$

114. It took **4.5 minutes** for **1.0 L** helium to effuse through a porous barrier. How long will it take for **1.0 L Cl₂** gas to effuse under identical conditions?

Answer: $\frac{\text{Rate}_1}{\text{Rate}_2} = \left(\frac{M_2}{M_1}\right)^{1/2}$, where M = molar mass; let gas (1) = He and gas (2) = Cl_2.

Effusion rates in this problem are equal to the volume of gas that effuses per unit time (L/min). Let t = time in the following expression.

$$\frac{\frac{1.0 \text{ L}}{4.5 \text{ min}}}{\frac{1.0 \text{ L}}{\text{ t}}} = \left(\frac{70.90}{4.003}\right)^{1/2}, \quad \frac{\text{t}}{4.5 \text{ min}} = 4.209, \text{ t} = 19 \text{ min}$$

115. Calculate the pressure exerted by $0.5000\ mole$ of N_2 in a 1.0000-L container at 25.0°C

- a. using the ideal gas law.
- b. using the van der Waals equation.
- c. Compare the results.

Answer:
a.
$$PV = nRT$$

 $P = \frac{nRT}{V} = \frac{0.5000 \text{ mol} \times \frac{0.08206 \text{ L atm}}{\text{K mol}} \times (25.0 + 273.2) \text{ K}}{1.0000 \text{ L}} = 12.24 \text{ atm}}$
b. $\left[P + a \left(\frac{n}{V}\right)^2\right] (V - nb) = nRT; \text{ for } N_2: a = 1.39 \text{ atm } L^2/\text{mol}^2 \text{ and } b = 0.0391 \text{ L/mol}}$
 $\left[P + 1.39 \left(\frac{0.5000}{1.0000}\right)^2 \text{ atm}\right] (1.0000 \text{ L} - 0.5000 \times 0.0391 \text{ L}) = 12.24 \text{ L atm}}$
 $(P + 0.348 \text{ atm})(0.9805 \text{ L}) = 12.24 \text{ L atm}$
 $P = \frac{12.24 \text{ L atm}}{0.9805 \text{ L}} - 0.348 \text{ atm} = 12.48 - 0.348 = 12.13 \text{ atm}}$

c. The ideal gas law is high by 0.11 atm, or $\frac{0.11}{12.13} \times 100 = 0.91\%$.

116. Calculate the pressure exerted by $0.5000\ mole$ of N_2 in a 10.000-L container at 25.0°C

- a. using the ideal gas law.
- b. using the van der Waals equation.
- c. Compare the results.
- d. Compare the results with those in Exercise 115.

Answer:
a.
$$PV = nRT$$

 $P = \frac{nRT}{V} = \frac{0.5000 \text{ mol} \times \frac{0.08206 \text{ L atm}}{\text{K mol}} \times 298.2 \text{ K}}{10.000 \text{ L}} = 1.224 \text{ atm}}$
b. $\left[P + a \left(\frac{n}{V}\right)^2\right] (V - nb) = nRT; \text{ for } N_2: a = 1.39 \text{ atm } L^2/\text{mol}^2 \text{ and } b = 0.0391 \text{ L/mol}}$
 $\left[P + 1.39 \left(\frac{0.5000}{10.000}\right)^2 \text{ atm}\right] (10.000 \text{ L} - 0.5000 \times 0.0391 \text{ L}) = 12.24 \text{ L atm}}$
 $(P + 0.00348 \text{ atm})(10.000 \text{ L} - 0.0196 \text{ L}) = 12.24 \text{ L atm}}$
 $P + 0.00348 \text{ atm} = \frac{12.24 \text{ L atm}}{9.980 \text{ L}} = 1.226 \text{ atm}, P = 1.226 - 0.00348 = 1.223 \text{ atm}}$
c. The results agree to $\pm 0.001 \text{ atm } (0.08\%).$

d. In Exercise 115, the pressure is relatively high, and there is significant disagreement. In Exercise 116, the pressure is around 1 atm, and both gas laws show better agreement. The ideal gas law is valid at relatively low pressures.

Not needed this semester

160. Nitrogen gas (N_2) reacts with hydrogen gas (H_2) to form ammonia gas (NH_3) . You have nitrogen and hydrogen gases in a **15.0-L** container fitted with a movable piston (the piston allows the container volume to change so as to keep the pressure constant inside the container). Initially the partial pressure of each reactant gas is **1.00 atm**. Assume the temperature is constant and that the reaction goes to completion.

- a. Calculate the partial pressure of ammonia in the container after the reaction has reached completion.
- b. Calculate the volume of the container after the reaction has reached completion.

Answer:

a. Initially $P_{N_2} = P_{H_2} = 1.00$ atm, and the total pressure is 2.00 atm ($P_{total} = P_{N_2} + P_{H_2}$). The total pressure after reaction will also be 2.00 atm because we have a constant-pressure container. Because V and T are constant before the reaction takes place, there must be equal moles of N₂ and H₂ present initially. Let $x = \text{mol } N_2 = \text{mol } H_2$ that are present initially. From the balanced equation, $N_2(g) + 3 H_2(g) \rightarrow 2 NH_3(g)$, H_2 will be limiting because three times as many moles of H₂ are required to react as compared to moles of N₂. After the reaction occurs, none of the H₂ remains (it is the limiting reagent).

Mol NH₃ produced = $x \mod H_2 \times \frac{2 \mod NH_3}{3 \mod H_2} = 2x/3$

Mol N₂ reacted = $x \mod H_2 \times \frac{1 \mod N_2}{3 \mod H_2} = x/3$

Mol N₂ remaining = $x \mod N_2$ present initially – $x/3 \mod N_2$ reacted = $2x/3 \mod N_2$

After the reaction goes to completion, equal moles of $N_2(g)$ and $NH_3(g)$ are present (2x/3). Because equal moles are present, the partial pressure of each gas must be equal $(P_{N_2} = P_{NH_3})$.

 $P_{total} = 2.00 \text{ atm} = P_{N_2} + P_{NH_3}$; solving: $P_{N_2} = 1.00 \text{ atm} = P_{NH_3}$

b. $V \propto n$ because P and T are constant. The moles of gas present initially are:

 $n_{N_{2}} + n_{H_{2}} = x + x = 2x \text{ mol}$

After reaction, the moles of gas present are:

$$n_{N_2} + n_{NH_3} = \frac{2x}{3} + \frac{2x}{3} = 4x/3 \text{ mol}$$

$$\frac{V_{after}}{V_{initial}} = \frac{n_{after}}{n_{initial}} = \frac{4x/3}{2x} = \frac{2}{3}$$

The volume of the container will be two-thirds the original volume, so:

$$V = 2/3(15.0 L) = 10.0 L$$