Worksheet 8 - Partial Pressures and the Kinetic Molecular Theory of Gases

Dalton’s Law of Partial Pressures states that the sums of the pressures of each gas in the mixture add to give the total pressure \( P_{\text{tot}} = P_1 + P_2 + P_3 + \ldots \)

1. The stopcock between a 5.00 L bulb containing hydrogen at 295 torr and an 8.00 L bulb containing nitrogen at 530 torr is opened. What is the total pressure of the mixture (assume constant \( T = 25.0^\circ\text{C} \))?

\[
P_{\text{tot}} = P_{\text{oxygen}} + P_{\text{nitrogen}}
\]

You can treat each gas separately. After mixing, there will be the same number of moles of each gas, at the same \( T \), just in a different volume.

a) For hydrogen, \( P_1 V_1 = P_2 V_2 \)

\[
V_1 = 5.00 \quad \text{L} \quad V_2 = 5.00 + 8.00 = 13.00 \quad \text{L}
\]

\[
P_1 = 0.388 \quad \text{atm} \quad P_2 = 0.149 \quad \text{atm}
\]

\[
P_1 = 295 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 0.388 \text{ atm}
\]

\[
P_2 = \frac{P_1 V_1}{V_2} = \frac{(0.388 \text{ atm})(5.00 \text{ L})}{13.00 \text{ L}} = 0.149 \text{ atm}
\]

b) For nitrogen, \( P_1 V_1 = P_2 V_2 \)

\[
V_1 = 8.00 \quad \text{L} \quad V_2 = 5.00 + 8.00 = 13.00 \quad \text{L}
\]

\[
P_1 = 0.697 \quad \text{atm} \quad P_2 = 0.429 \quad \text{atm}
\]

c) \( P_{\text{tot}} = P_{O_2} + P_{N_2} = 0.149 \text{ atm} + 0.429 \text{ atm} = 0.578 \text{ atm} \)

The total pressure is equal to the sum of the partial pressure of each gas.
2. A sample of nitrogen gas is collected over water at 20.°C and a pressure of 1.00 atm. The volume of the collected gas is 250.0 L. What is the mass of N₂ collected? (At 20.°C, the vapor pressure of water is 17.5 torr).

\[
P_{\text{water}} = 17.5 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 0.023 \text{ atm}
\]

\[
P_{\text{tot}} = P_{\text{nitrogen}} + P_{\text{water}} \quad P_{\text{nitrogen}} = 1.00 \text{ atm} - 0.023 \text{ atm} = 0.977 \text{ atm}
\]

\[
n = \frac{PV}{RT} = \frac{(0.977 \text{ atm})(250.0 \text{ L})}{\left(0.08206 \frac{L \cdot \text{ atm}}{\text{mol} \cdot K}\right)(20 + 273 \text{ K})} = 10.16 \text{ mol N}_2
\]

\[
10.16 \text{ mol N}_2 \times \frac{28.02 \text{ g N}_2}{1 \text{ mol N}_2} = 285 \text{ g N}_2
\]

\[
\text{mass of N}_2 = 285 \text{ g}
\]

Skip question #3 – mol fractions will not be covered on exam #1

The Kinetic Molecular Theory of Gases explains the behavior of ideal gases. Shown below are the assumptions that this model makes:

*The volume of the gaseous molecules is negligible compared to the total volume in which the gas is contained.* Ideal gases are viewed as having mass but no volume.

*The gaseous molecules are in constant, random motion. Their collisions with the walls of the container are the cause of the pressure exerted by the gas.*

*Attractive and repulsive forces between molecules are negligible.* Ideal gases do not interact, chemically, with any other matter.

*The average kinetic energy of the molecules is proportional to the absolute temperature (K) of the gas.* At any given temperature the molecules of all gases have the same average kinetic energy.
4. Consider two gases, A and B, in containers of equal volume. Both are at the same V, T and P.

<table>
<thead>
<tr>
<th></th>
<th>A</th>
<th>B</th>
</tr>
</thead>
<tbody>
<tr>
<td>mass</td>
<td>0.34 g</td>
<td>0.48 g</td>
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</table>

Are the following statements true or false? Why?

a) The number of molecules of A is equal to the number of molecules of B
   True – both samples have equal P, V and T, so they must also have the same n

b) The molar mass of A is greater than the molar mass of B
   False – since both samples have equal moles, the value of g/mol is smaller for A

c) Both samples have the same average kinetic energy.
   True – average kinetic energy depends only on temperature, which is identical for both samples

d) The molecules of A have the same average velocity as the molecules of B
   False – kinetic energy is equal to \( \frac{1}{2} \text{mass} \times \text{velocity}^2 \). The masses of the samples are different, so the velocities must also be different to maintain constant kinetic energy

e) The molecules of A collide with the container walls more frequently than the molecules of B
   True – the molecules of A are moving at a higher average velocity, so they will collide more frequently with the walls of the container. The total pressures of the two gases can remain equal, however, because the molecules of B collide with the container walls with more force per collision (because B has a higher mass).
**Graham’s Law of Effusion (Diffusion)** Graham found that the average velocity of a molecule is inversely proportional to the square root of the molar mass, M.

\[ u = \text{average velocity} = \text{constant} \times \frac{1}{\sqrt{M}} \]

For two gases: \[ \frac{u_1}{u_2} = \sqrt{\frac{M_2}{M_1}} \]

5. In the lecture demo, HCl and NH₃ gases were inserted into opposite ends of a tube.

a) Which gas is expected to travel faster?
   \[ \text{NH}_3 \text{ (it has the lighter molar mass)} \]

b) Which gas is expected to travel farther?
   \[ \text{NH}_3 \text{ (it has the higher velocity)} \]

c) If the distance each gas travels is directly proportional to the velocity of the gas, calculate the ratio of the distances traveled.

\[ \frac{\text{distance}_{\text{NH}_3}}{\text{distance}_{\text{HCl}}} = \sqrt{\frac{M_{\text{HCl}}}{M_{\text{NH}_3}}} = \sqrt{\frac{36.45 \text{ g/mol}}{17 \text{ g/mol}}} = 1.5 \]

d) Does this calculation agree with what you saw in lecture?
The experiment in lecture wasn’t complete by the end of the class period. We could see the ammonia vapor in the tube, and it had traveled more than halfway through the tube without reacting with the HCl. We don’t know exactly where the two gases would meet, but it was past the halfway point (and closer to the HCl end).

e) Real gases do not always behave ideally. List two ways to increase the ideality of this experiment

Gases behave most ideally at high temperature or low pressure. The experiment could be altered by increasing the temperature of the room (or applying a heat source near the tube). The pressure could be decreased by either increasing the volume of the tube (more space for the gas particles to travel before colliding with the wall of the tube) or decreasing the amount of sample (less likely for gas particles to collide with other gas particles).