Dalton’s Law, or the Law of Partial Pressures, states that the total pressure exerted by a mixture of gases is equal to the sum of the partial pressures of the gases in the mixture.

**Explanation**

Based on the kinetic theory of gases, a gas will diffuse in a container to fill up the space it is in and does not have any forces of attraction between the molecules. In other words, the different molecules in a mixture of gases are so far apart that they act independently; they do not react with each other. The pressure of an ideal gas is determined by its collisions with the container, not collisions with molecules of other substances since there are no other collisions. A gas will expand to fill the container it is in without affecting the pressure of another gas. So it can be concluded that the pressure of a certain gas is based on the number of moles of that gas and the volume and temperature of the system. Since the gases in a mixture of gases are in one container, the Volume (V) and Temperature (T) for the different gases are the same as well. Each gas exerts its own pressure on the system, which can be added up to find the total pressure of the mixture of gases in a container. This is shown by the equation

\[ P_{total} = P_A + P_B + \ldots \tag{1} \]

**Derivation**

We have, from the Ideal Gas Law

\[ PV = nRT \tag{2} \]

If we know the molar composition of the gas, we can write

\[ n_{total} = n_a + n_b + \ldots \tag{3} \]

Again, based on the kinetic theory of gases and the ideal gas law, Dalton’s law can also be applied to the number of moles so that the total number of moles equals the sum of the number of moles of the individual gases. Here, the pressure, temperature and volume are held constant in the system. The total volume of a gas can be found the same way, although this is not used as much. This yields the equation,

\[ P_{total} V = n_{total} RT \tag{4} \]

We can rearrange this equation to find the total number of moles. Sometimes masses of each sample of gas are given and students are asked to find the total pressure. This can be done by converting grams to moles and using Dalton’s law to find the pressure.

**Mole Ratio**

From the partial pressure of a certain gas and the total pressure of a certain mixture, the mole ratio, called \( X_i \), of a gas can be found. The mole ratio describes what fraction of the mixture is a specific gas. For example, if oxygen exerts 4 atm of pressure in a mixture and the total pressure of the system is 10 atm, the mole ratio would be 4/10 or 0.4. The mole ratio
applies to pressure, volume, and moles as seen by the equation below. This also means that 0.4 moles of the mixture is made up of gas i.

\[X_i = \frac{P_i}{P_{\text{tot}}} = \frac{n_i}{n_{\text{tot}}} = \frac{V_i}{V_{\text{tot}}} \tag{5}\]

The mole ratio, \(\{X_i\}\) is often used to determine the composition of gases in a mixture. The sum of the mole ratios of each gas in a mixture should always equal one since they represent the proportion of each gas in the mixture.

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**Collection of a Gas Over Water**

The Law of Partial Pressures is commonly applied in looking at the pressure of a closed container of gas and water. The total pressure of this system is the pressure that the gas exerts on the liquid. The gas is made up of whatever sample of gas there is plus the evaporated water. The pressure of the gas on the liquid consists of the pressure of the evaporated water and the pressure of the gas collected. Based on Dalton's law, the pressure of the gas collected can be calculated by subtracting the pressure of the water vapor from the total pressure.

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**Real Gases**

Real gases are gases that do not behave ideally. That is, they violate one or more of the rules of the kinetic theory of gases. Real gases behave ideally when the gases are at low pressure and high temperature. Therefore at high pressures and low temperatures, Dalton's law is not applicable since the gases are more likely to react and change the pressure of the system. For example, if there are forces of attraction between the molecules, the molecules would get closer together and the pressure would be adjusted because the molecules are interacting with each other.

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**Problems**

1. A sample of gas A evaporates over water in a closed system. What is the pressure of gas A if the total pressure is 780 torr and water vapor pressure is 1 atm?

2. There is a mixture of 4 moles of hydrogen gas, 8 moles of oxygen, 12 moles of helium, and 6 moles of nitrogen in a closed container. What is the total number of moles in this system?

3. If there is a mixture of hydrogen and oxygen gas in a container with 10 moles total and the mole ratio of hydrogen is .67 mol_{Hydrogen} to 1 mol_{Total}, how many moles of each gas are there?

4. 24.0 L of nitrogen gas at 2 atm and 12.0 L of oxygen gas at 2 atm are added to a 10 L container at 273 K. Find the partial pressure of nitrogen and oxygen and then find the total pressure.

5. Flourine gas is in a 5.0 L container that is 25 C and 2 atm. A certain amount of hydrogen with a partial pressure of .5 is added to the container. What is the mole ratio of hydrogen?

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**Solutions**

1.

- Convert pressure to same units so 780 torr=1.03 atm
• Subtract water vapor pressure from total pressure to get partial pressure of gas A: \( P_A = 1.03 \text{ atm} - 1 \text{ atm} = 0.03 \text{ atm} \)

2. The law of partial pressures also applies to the total number of moles if the other values are constant, so

\[
4 \text{ mol Hydrogen} + 8 \text{ mol Oxygen} + 12 \text{ mol Helium} + 6 \text{ mol Nitrogen} = \text{30 moles total}
\]

3.

• \( 10 \text{ mol total} \times 0.67 \text{ mol Hydrogen} = 6.7 \text{ moles H}_2 \text{ gas} \)

• \( 10 \text{ mol total} - 6.7 \text{ mol Hydrogen} = 3.3 \text{ moles O gas} \)

4.

• Find the number of moles of oxygen and nitrogen using \( PV = nRT \) which is \( n = \frac{PV}{RT} \)
  1. oxygen: \( \frac{(1 \text{ atm})(12 \text{ L})}{(0.08206 \text{ atm L mol}^{-1} \text{ K}^{-1})(273 \text{ K})} = 0.536 \text{ moles oxygen} \)
  2. nitrogen: \( \frac{(1 \text{ atm})(24.0 \text{ L})}{(0.08206 \text{ atm L mol}^{-1} \text{ K}^{-1})(273 \text{ K})} = 1.07 \text{ moles of Nitrogen} \)
  3. add to get \( n_{\text{tot}}: 0.536 \text{ mol Oxygen} + 1.07 \text{ mol Nitrogen} = 1.61 \text{ moles total} \)

• Use \( PV = nRT \) or \( P = \frac{(nRT)}{V} \) to find the total pressure
  1. \( P_{\text{tot}} = \frac{(1.61 \text{ mol total})(0.08206 \text{ atm L mol}^{-1} \text{ K}^{-1})(273 \text{ K})}{(10.0 \text{ L})} = 3.61 \text{ atm} \)

• \( P_A/P_{\text{tot}} = n_A/N_{\text{tot}} \) can be rearranged to \( P_A = P_{\text{tot}}(n_A/N_{\text{tot}}) \) to find the partial pressures
  1. \( P_{\text{Oxygen}} = (3.61 \text{ atm total})(0.536 \text{ mol Oxygen}/1.61 \text{ mol total}) = 1.20 \text{ atm Oxygen} \)
  2. \( P_{\text{Nitrogen}} = 3.61 \text{ atm total} - 1.20 \text{ atm Oxygen} = 2.41 \text{ atm Nitrogen} \)

5.

• total pressure is \( 2 \text{ atm Fluorine} + 0.5 \text{ atm Hydrogen} = 2.5 \text{ atm total} \)

• remember that \( p_H/P_{\text{tot}} = n_H/n_{\text{tot}} = \nu_H/\nu_{\text{tot}} \) so the mole ratio of hydrogen to the mixture is is \( 0.5/2.5 = \frac{2 \text{ mol Hydrogen} \text{ to 1 mol total}}{2.5} \)

References