## Ideal Gas Mixtures

***Dalton's Law of Partial Pressures***

The *partial pressure* of a gas in a mixture is the pressure it would exert if alone in the container. *Dalton's law of partial pressures* says that the total pressure of a mixture of gases is the sum of the partial pressures. Thus, for a mixture of *n* gases,

$$p_{Total} = p_1 + p_2 + p_3 + \;\cdots\; = \sum_{i=1}^{n} p_i \qquad (1)$$

where

$$p_i = n_i\left(\frac{RT}{V}\right)\;,\quad p_{Total} = n_{Total}\left(\frac{RT}{V}\right)\quad \text{and}\quad n_{Total} = n_1 + n_2 + n_3 + \;\ldots\; = \sum_{i=1}^{n} n_i \qquad (2)$$

It follows from Eq(2) that

$$\frac{p_i}{p_{Total}} = \frac{n_i(RT/V)}{n_{Total}(RT/V)} = \frac{n_i}{n_{Total}} = X_i \qquad (3)$$

where $X_i$ is the *mole fraction* of component *i*.

***Vapor Pressure, Relative Humidity, and the Dew Point***

A gas commonly present in gas mixtures is water vapor, $H_2O(g)$, which exerts a partial pressure, $p_{H_2O}$. The maximum value of $p_{H_2O}$ at a given temperature is the *vapor pressure* of water. The *vapor pressure* of a gas is the pressure exerted by a vapor *in equilibrium with its liquid in a closed container.* In experiments in which a gas is collected over water, $p_{H_2O}$ contributes to the total gas pressure in the container (see exercise (1) below).

Weather reports often give the *relative humidity* which is the percent of the equilibrium vapor pressure at the reported temperature which is actually present in the atmosphere:

$$\text{Relative Humidity} \;=\; \frac{p_{H_2O}}{(\text{vapor pressure})} \times 100\% \qquad (4)$$

Since the vapor pressure increases as the temperature increases, a relative humidity of 90% indicates a much higher $p_{H_2O}$ on a hot day in summer than on a cold day in winter! As the temperature drops for a given $p_{H_2O}$, the *relative humidity* increases (since the vapor pressure decreases as T decreases); when the *relative humidity* reaches 100% the *dew point* has been reached, and water vapor begins to condense (as "dew" or "frost").

***Average Molecular Speed (Root-Mean-Square Speed)***

At a given temperature a molecule of any gas has the *same* average kinetic energy, $\overline{KE} = \frac{1}{2}m\overline{u^2}$.

This implies that gas molecules with *low* molar masses must have *higher* average speeds $\overline{u^2}$ than those with high molar masses (why?). As shown by Chang (p. 196), the average (root-mean-square) speed of a gas molecule with molar mass *M* may be calculated at a temperature *T* :