The Composition of Air

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Pressure is the normal component of the force per unit area exerted by fluid molecules. In an ideal fluid the pressure at a point is the same in all directions. The unit of pressure commonly used in the atmospheric sciences is 1 mb = 100 h Pa. Temperature is a measure of the kinetic energy of the random molecular motions.

A perfect gas is one that exactly obeys the laws of Boyle and Charles. The equation of state for such a gas is

\[ p\alpha = \frac{R}{\gamma}T , \]

(1)

where \( R' \) is the “specific” gas constant. The specific gas constant varies with the type of gas. Avogadro found that 1 g molecular weight (mole) of any gas occupies 22400 cm\(^3\) at temperature \( T_0 = 0{\circ}C \), at pressure \( p_0 = 1 \). (Obviously, this reference temperature and reference pressure have been arbitrarily chosen, and the particular volume measured, i.e. 22400 cm\(^3\), depends on these choices.) This means that, for the particular case of \( V_0 = 22400 \text{ cm}^3 \), the equation of state becomes

\[ p_0 V_0 = mR'T_0 , \]

(2)

where \( m \) is the molecular weight. Avogadro’s discovery implies that there exists a universal gas constant:

\[ R' = mR' = \frac{p_0V_0}{T_0} . \]

(3)

The equation of state can now be written as

\[ p\alpha = \frac{R'}{m}T . \]

(4)
As shown in Table 1, “dry air” is a mixture of nitrogen, oxygen, argon, carbon dioxide, etc., which are all, practically speaking, perfect gases, and so obey (4). The composition of dry air is nearly homogeneous below 20 km. Except for water vapor and ozone, whose concentrations vary greatly, the concentrations of the other principal constituents of the atmosphere, i.e., N\textsubscript{2}, O\textsubscript{2}, Ar, CO\textsubscript{2}, Ne, He, Kr, H\textsubscript{2}, CH\textsubscript{4}, and N\textsubscript{2}O, are nearly homogeneous up to about 80 km.

<table>
<thead>
<tr>
<th>Gas</th>
<th>Molecular Weight.</th>
<th>(R'), J kg K(^{-1})</th>
<th>Mass fraction of the “dry” portion of the atmosphere.</th>
</tr>
</thead>
<tbody>
<tr>
<td>Nitrogen</td>
<td>28.016</td>
<td>296.7</td>
<td>75.52</td>
</tr>
<tr>
<td>Oxygen</td>
<td>32.000</td>
<td>259.8</td>
<td>23.15</td>
</tr>
<tr>
<td>Argon</td>
<td>39.444</td>
<td>208.1</td>
<td>1.28</td>
</tr>
<tr>
<td>Carbon Dioxide</td>
<td>44.010</td>
<td>188.9</td>
<td>0.0035</td>
</tr>
</tbody>
</table>

Table 1: The composition of “dry air.”

For a mixture of perfect gases occupying volume \(V\) at temperature \(T\), Dalton's law states that:

- each gas completely occupies the volume;
- each gas obeys its own equation of state;
- the total pressure due to the mixture of gases is the sum of the partial pressures exerted by the individual gases.

Therefore,

\[
p_iV = M_iR'\frac{R^*}{m_i}T, \quad i = 1, 2, \ldots, n.
\]

(5)

Here subscript \(i\) denotes a particular species, \(p_i\) is the partial pressure, \(M_i\) is the mass, \(R^*\) is the universal gas constant, and \(m_i\) is the molecular weight. It follows that

\[
V \sum_{i=1}^{n} p_i = T \sum_{i=1}^{n} M_i R'\frac{R^*}{m_i} = R^* T \sum_{i=1}^{n} M_i \frac{1}{m_i}.
\]

(6)

Using
\[ p = \sum_{i=1}^{n} p_i, \quad \rho = \sum_{i=1}^{n} \frac{M_i}{V_i}, \quad M = \sum_{i=1}^{n} M_i, \quad (7) \]

we find that

\[ p = \rho RT, \quad (8) \]

where the effective gas constant of the mixture is

\[ R = \sum_{i=1}^{n} \frac{M_i R_i'}{M} = \sum_{i=1}^{n} \frac{M_i}{m_i}. \quad (9) \]

For dry air, \( R = R_d = 287 \text{ J kg}^{-1} \text{ K}^{-1} \). The apparent molecular weight of dry air is

\[ m_d = \frac{R^*}{R} = 28.966 \text{ g mole}^{-1}. \quad (10) \]

When the effects of moisture are included, (8) is often modified to use the gas constant for dry air, with a “virtual temperature.”