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The Composition of Air

David A. Randall

Department of Atmospheric Science
Colorado State University, Fort Collins, Colorado 80523

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 = 1 h Pa.

Temperature is a measure of the kinetic energy of the random molecular motions.

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

$$p\alpha = R' \, T \ , \tag{1}$$

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. For the particular case of V_0 = 22400 cm^3, the equation of state becomes

$$p_0 \, V_0 = mR' \, T_0 \ , \tag{2}$$

where m is the molecular weight. Avogadro's discovery implies that there exists a *universal* gas constant:

$$R^* = mR' = \frac{p_0 \, V_0}{T_0} \ . \tag{3}$$

The equation of state can now be written as

$$p\alpha = \frac{R^*}{m} T \ . \tag{4}$$

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^2, O^2, Ar, CO^2, Ne, He, Kr, H_2, CH_4, and N_2O, are nearly homogeneous up to an altitude of about 80 km.

Table 1: The composition of "dry air."

Gas	*Molecular Weight*	R', $\mathrm{J\,kg^{-1}\,K^{-1}}$	*Mass fraction of the dry portion of the atmosphere*
Nitrogen	28.016	296.7	75.52
Oxygen	32.000	259.8	23.15
Argon	39.444	208.1	1.28
Carbon Dioxide	44.010	188.9	0.0035

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,

$$p_i\,V = M_i\,R_i'\,T = M_i\,\frac{R^*}{m_i}\,T,\ \ i = 1,\ 2,\ \ldots,\ n\ . \tag{5}$$

Here subscript i denotes a particular species, p_i is the partial pressure and, M_i is the mass. It follows that

$$V\sum_{i=1}^{n} p_i = T\sum_{i=1}^{n} M_i\,R_i' = R^*\,T\sum_{i=1}^{n}\frac{M_i}{m_i}\ . \tag{6}$$

Using

$$p \equiv \sum_{i=1}^{n} p_i,\ \rho \equiv \sum_{i=1}^{n}\frac{M_i}{V},\ \ M \equiv \sum_{i=1}^{n} M_i\ , \tag{7}$$

we find that

$$p = \rho R T \ , \tag{8}$$

where the effective gas constant of the mixture is

$$R \equiv \sum_{i=1}^{n} \frac{M_i \, R_i'}{M} = \sum_{i=1}^{n} \frac{M_i}{m_i} \ . \tag{9}$$

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

$$m_d = \frac{R^*}{R} \equiv 28.966 \, g \ \mathrm{mole}^{-1} \ . \tag{10}$$

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