

## Lecture 3. The basic properties of the natural atmosphere

Objectives:

1. Composition of air.
2. Pressure.
3. Temperature.
4. Density.
5. Concentration. Mole. Mixing ratio.
6. Gas laws.
7. Dry air and moist air.

Readings: Turco: p.11-27, 38-43, 366-367, 490-492; Brimblecombe: p. 1-5

### 1. Composition of air.

The word *atmosphere* derives from the Greek *atmo* (vapor) and *spherios* (sphere).

The Earth's atmosphere is a mixture of gases that we call **air**. Air usually contains a number of small particles (atmospheric aerosols), clouds of condensed water, and ice cloud.

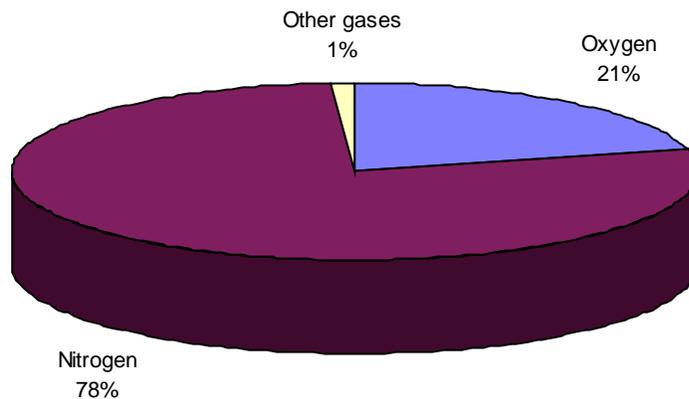
NOTE : The atmosphere is a thin veil of gases; if our planet were the size of an apple, its atmosphere would be thick as the apple peel. Some 80% of the mass of the atmosphere is within 10 km of the surface of the Earth, which has a diameter of about 12,742 km.

The Earth's atmosphere as a mixture of gases is characterized by pressure, temperature, and density which vary with altitude (will be discussed in Lecture 4).

The atmosphere below about 100 km is called Homosphere. This part of the atmosphere consists of uniform mixtures of gases as illustrated in Table 3.1.

Table 3.1. The composition of air.

Gases	Fraction of air
<b>Constant gases</b>	
Nitrogen, N <sub>2</sub>	<b>78.08%</b>
Oxygen, O <sub>2</sub>	<b>20.95%</b>
Argon, Ar	<b>0.93%</b>
Neon, Ne	<b>0.0018%</b>
Helium, He	<b>0.0005%</b>
Krypton, Kr	<b>0.00011%</b>
Xenon, Xe	<b>0.000009%</b>
<b>Variable gases</b>	
Water vapor, H <sub>2</sub> O	<b>4.0%</b> (maximum, in the tropics) <b>0.00001%</b> (minimum, at the South Pole)
Carbon dioxide, CO <sub>2</sub>	<b>0.0365%</b> (increasing ~0.4% per year)
Methane, CH <sub>4</sub>	~ <b>0.00018%</b> (increases due to agriculture)
Hydrogen, H <sub>2</sub>	~ <b>0.00006%</b>
Nitrous oxide, N <sub>2</sub> O	~ <b>0.00003%</b>
Carbon monoxide, CO	~ <b>0.000009%</b>
Ozone, O <sub>3</sub>	~ <b>0.000001% - 0.0004%</b>
Fluorocarbon 12, CF <sub>2</sub> Cl <sub>2</sub>	~ <b>0.00000005%</b>



- Some gases in Table 3.1 are called **constant gases** because the ratio of the number of molecules for each gas and the total number of molecules of air do not change substantially from time to time or place to place.
- The fraction of many other gases do change significantly over time. Such gases typically exist only in small quantities, thus, they are referred to as trace gases. Yet trace compounds are very important for atmospheric physical and chemical processes.
- Water vapor is a ‘special’ gas because its abundance depends both on temperature and availability of water. For example, when temperature is low, water vapor easily condenses to form liquid or deposits as ice.
- Ozone is a trace gas that is not emitted into the atmosphere. Instead, it is formed only by photochemical reactions. Ozone fraction is typically peak at about 25-32 km in altitude.

NOTE: the actual composition of the atmosphere is much richer than is presented by Table 3.1. The actual number of **trace constituents** reaches into the thousands.

NOTE: atmospheric gases (except some molecules) can not escape to space because of Earth gravitational forces which keep them near the Earth.

The composition of the atmosphere is a key factor determining its properties. The Earth’s atmosphere serves us in several capacities:

- atmospheric gases participate in essential life processes (e.g., oxygen is used in respiration by animals and humans; carbon dioxide is used for plant respiration);
- atmosphere is a shield for protection from dangerous ultraviolet rays in sunlight, and is a blanket that traps heat and keeps Earth’s surface comfortably warm.

## 2. Pressure.

Pressure is force, divided by the area over which it is applied:

$$P = \text{pressure} = \text{force/area}$$

Due to the Earth's gravity, the weight of air creates a pressure called the atmospheric pressure.

NOTE: distinguish between mass and weight.

Mass is an absolute measure of the amount of a material, weight is the measure of the force of gravity acting on a given mass

1 atmosphere is the pressure exerted by the atmosphere at sea level at temperature 0°C, abbreviated **1 atm**. At higher altitudes, the pressure is smaller because less of the atmosphere is pressing down from above. In fact, in the atmosphere, pressure decreases exponentially with altitude (see Lecture 4).

Barometer is an instrument that measures the atmospheric pressure.

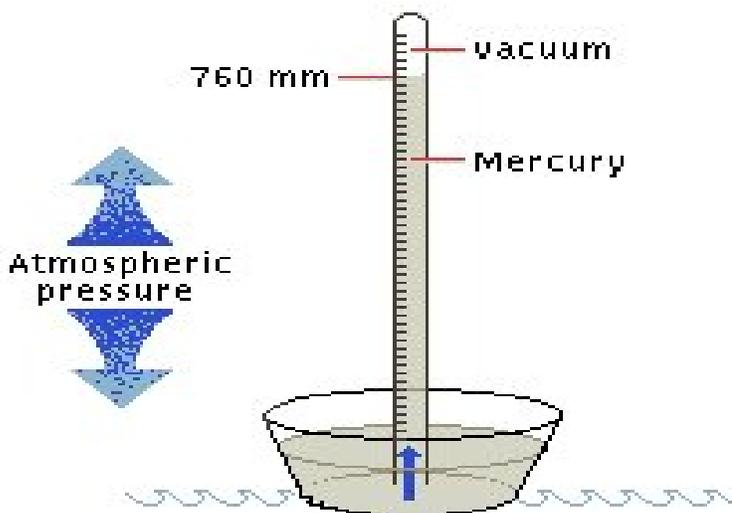


Table 3.2 Units of pressure.

Unit	Symbol	Value of atmospheric P in that unit (at sea level and 0°C)
atmosphere	<b>atm</b>	<b>1 atm</b>
millimeters of mercury	<b>mm Hg (torr)</b>	<b>760 mm Hg (760 torr)</b>
pascal	<b>Pa</b>	<b>1.01325x10<sup>5</sup> Pa</b>
millibar	<b>mbar (mb)</b>	<b>1013.25 mbar</b>

- Pressure Conversion:

$$1 \text{ bar} = 10^3 \text{ mb} = 10^5 \text{ Pa} = 0.986923 \text{ atm} = 10^6 \text{ g cm}^{-1} \text{ s}^{-2}$$

- SI units for pressure: Pascal

$$\text{Pa} = \text{N m}^{-2} = \text{m}^{-1} \text{ kg s}^{-2}$$

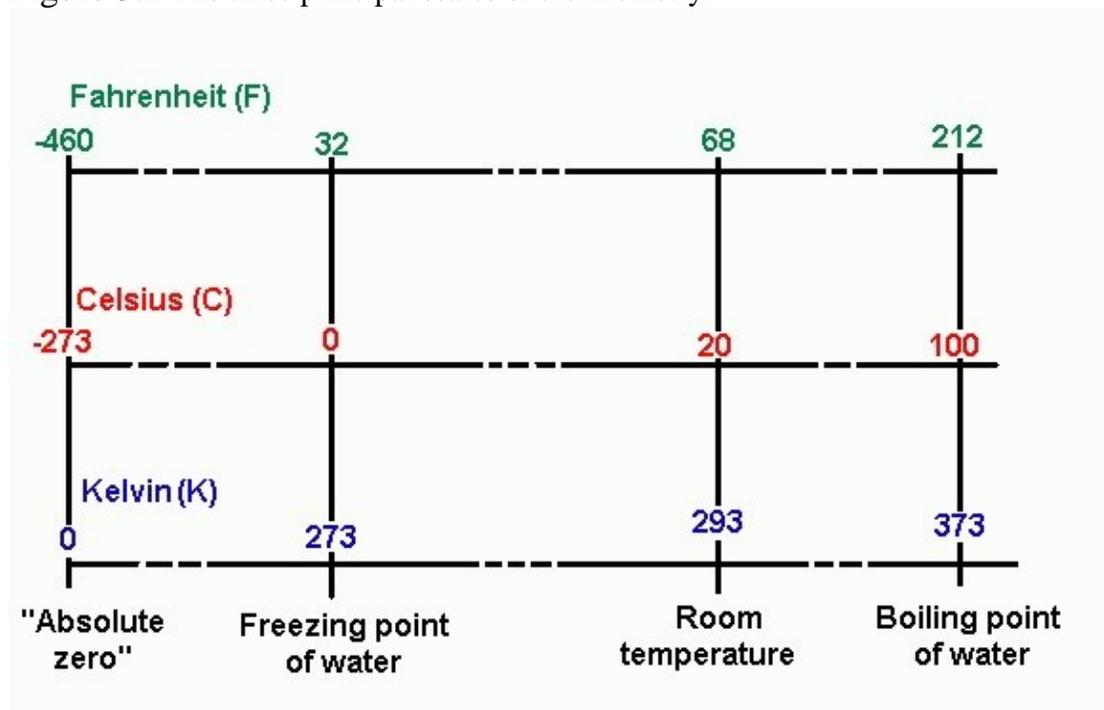
### 3. Temperature.

**Temperature** measures the amount of internal energy that a substance holds. This energy is due to movement of atoms and molecules comprising the substance (for example, the speed of the molecule in a gas). The higher the temperature is, the faster the atoms and molecules are moving.

**Air Temperature** is a measure of the kinetic energy or average speed of an air molecule.

**Thermometers** are instruments to measure temperature. The three most common temperature scales are Fahrenheit, Centigrade (or Celsius), and Kelvin scales.

Figure 3.2 The three principal scales of thermometry.



Conversion rules:

- to convert from °F to °C:

$$^{\circ}\text{C} = (^{\circ}\text{F} - 32^{\circ}) \cdot 5/9$$

- to convert from °C to °F:

$$^{\circ}\text{F} = ^{\circ}\text{C} \cdot 5/9 + 32^{\circ}$$

- to convert from °C to °K:

$$^{\circ}\text{C} = ^{\circ}\text{K} - 273.16$$

**4. Density** of a gas (or any liquid and solid material) is a measure of the mass of that material present in a specified volume of space:

$$\rho = \text{density} = \text{mass} / \text{volume} = \mathbf{M / V}$$

Units:  $\text{kg/m}^3$  if M is in kg and V is in  $\text{m}^3$

$\text{g/cm}^3$  if M is in g and V is in  $\text{cm}^3$

**Air density** is the mass of air per unit volume of air.

## 5. Concentration. Mole. Mixing ratio.

Number concentration is the number of gas molecules in the sample of the volume:

$$n = \text{number concentration} = \text{total number of molecules} / \text{volume} = N/V$$

where N is the total number of gas molecules in the volume V.

$$\text{Units: } \begin{array}{l} \text{number/m}^3 = \text{molecules/m}^3 \text{ if } V \text{ is in } \text{m}^3 \\ \text{number/cm}^3 \text{ if } V \text{ is in } \text{cm}^3 \end{array}$$

Mass concentration is the mass of gas molecules in the sample of the volume:

$$q = \text{mass concentration} = \text{total mass of molecules} / \text{volume} = M/V$$

where M is the total mass of gas molecules in the volume V.

$$\text{Units: } \text{kg/m}^3 ; \text{g/cm}^3 ; \mu\text{g/m}^3 ;$$

NOTE: mass concentration of a gas = gas density

- Relationship between **q** and **n**:

if each molecule has mass m, then total mass is  $M = N \cdot m$  and

$$q = M/V = N m / V = n m$$

thus **q = n m**

NOTE: the mass of a single molecule is very small, and the number of molecules is usually enormous! For instance, the number concentration of air is

$$n = 2.55 \times 10^{19} \text{ molec./cm}^3 \text{ at } p=1013 \text{ mbar and } T = 288 \text{ K.}$$

Mole is introduced to operate with large collections of atoms or molecules.

**One mole** of any substance consists of  $6.022 \times 10^{23}$  units of that substance.

This number  $N_A = 6.022 \times 10^{23}$  is called Avogadro's number.

Example:

1 mole of eggs is  $6.022 \times 10^{23}$  eggs (just as a dozen eggs is 12 eggs);

1 mole of oxygen atoms is  $6.022 \times 10^{23}$  oxygen atoms;

The International System of Units (SI) defines the **mole** (abbreviation mol) as the number equal to the number of atoms in exactly 12 grams of pure  $^{12}\text{C}$  (carbon twelve). Also,  $^{12}\text{C}$  is as the standard in the system of atomic masses. In this system  $^{12}\text{C}$  is assigned a mass of exactly 12 atomic mass units (amu), and the masses of all other atoms are given relative to the standard (more details are given in Lecture 5).

Relationships between atomic mass units and grams:

Since  $6.022 \times 10^{23}$  atoms of  $^{12}\text{C}$  (each with a mass of 12 amu) have a mass of 12 grams, then

$$(6.022 \times 10^{23} \text{ atoms}) (12 \text{ amu/atom}) = 12 \text{ g}$$

$$6.022 \times 10^{23} \text{ amu} = 1 \text{ g}$$

or

$$1 \text{ amu} = 1.66 \times 10^{-24} \text{ g}$$

**The mass of one mole,  $m_g$**  of the element is equal to its atomic mass in grams.

Example:

the mass of 1 mole of H (hydrogen) atoms is 1.008 g because hydrogen atomic mass is 1.008 amu;

the mass of 1 mole of O (oxygen) atoms is 16 g because oxygen atomic mass is 16 amu.

**The mass of one mole,  $m_g$**  (or molar mass, or molecular weight) of a chemical compound, which is a collection of atoms, is obtained by summing the masses of atoms.

NOTE: the terms molar mass and molecular weight mean exactly the same thing: the mass in grams of one mole of a compound.

**Problem:** What is molecular weight of methane ( $\text{CH}_4$ )?

**Solution.**

Mass of 1 mole of C = 12.011 g

Mass of 4 mole of H = 4 x 1.008 g

Thus, mass of 1 mol of  $\text{CH}_4$  = 16.043 g

NOTE: that  $m_g$  relates to  $m$  as :  $m_g = m N_A$

**The number of moles** can be expressed as :

$\mu = \text{number of moles} = \text{mass of gas} / \text{molecular weight} = M / m_g$

or as

$\mu = \text{number of gas molecules} / \text{Avagadro's number} = N / N_A$

**Molar concentration** is the number of moles a gas in the sample of the volume:

$c = \text{molar concentration} = \text{number of moles} / \text{volume} = \mu / V$

where  $\mu$  is the number of moles of a gas in the volume  $V$ .

Units:  $\text{mol/m}^3$  if  $V$  is in  $\text{m}^3$   
 $\text{mol/cm}^3$  if  $V$  is in  $\text{cm}^3$

- Relationship between **c** and **n**:

since  $c = \mu/V$  and  $\mu = N/N_A$ , we obtain

$$c = \mu/V = (N/N_A)/V = (N/V)/N_A = n/N_A$$

thus  **$c = n / N_A$**

- Relationship between **c** and **q**:

since  $n = q / m$ , we obtain

$$c = n/N_A = (q/m)/N_A = q/(m N_A) = q/m_g$$

thus  **$c = q / m_g$**

**Problem:** Although air is a mixture of gases, it is often that the term “air molecule” is used. What is an average mass of the air molecule (take molecular weight of air  $m_a = 28.97$  g/mol)?

**Solution.**

One mole of air contains the Avogadro’s number of air molecules, therefore average mass of air molecule =  $m_a / N_A = 28.97$  (g mol<sup>-1</sup>) /  $6.022 \times 10^{23}$  mol<sup>-1</sup>)  
thus average mass of air molecule =  $4.8096 \times 10^{-23}$  g

**Mixing ratio** is defined as the ratio of the amount (or mass) of the substance in a given volume to the total amount (or mass) of all constituents in that volume.

Since air is a mixture of gases, the volume mixing ratio of i-th gas in the air is

**volume mixing ratio =**

**number of molecules of i-th gas/number of air molecules**

or

$$r_i = N_i / N_a$$

where  $N_i$  is the number of molecules of i-th gas, and  $N_a$  is the number of molecules of air (total number of molecules of all individual gases in the atmosphere:  $N_a = \sum N_i$ ).

NOTE:  $r_i = N_i / N_a = n_i / n_a = \mu_i / \mu_a = c_i / c_a$

NOTE: Table 3.1 presents volume mixing ratios of individual gases in the atmosphere expressed in percent.

For instance, the volume mixing ratio of carbon dioxide (CO<sub>2</sub>) is (see Table 3.1):

$r_{CO_2} = 0.000365$ , which is very small number. Therefore, it is multiplied by  $10^6$  and expressed as

$r_{CO_2} = 365$  **ppmv** (parts per million by volume).

In turn, mass mixing ratio of i-th gas is

**mass mixing ratio = mass of molecules of i-th gas/mass of air molecules**

or

$$r_{i,m} = M_i / M_a$$

where  $M_i$  is the mass of molecules of i-th gas, and  $M_a$  is the mass of molecules of air (total mass of molecules of all individual gases in the atmosphere:  $M_a = \sum M_i$ ).

NOTE:  $r_{i,m} = M_i / M_a = \rho_i / \rho_a$

Mass mixing ratio and volume mixing ratio are related as:

$$r_{i,m} = \rho_i / \rho_a = m_i n_i / m_a n_a = m_i / m_a \cdot n_i / n_a = m_i / m_a r_i$$

thus, mass mixing ratio of a certain component is its volume mixing ratio multiplied by its molecular weight and divided by the molecular weight of air ( $m_a=28.97$ )

Mass and volume mixing ratios are distinguished by an added v(for volume)and m(for mass), such as

ppmv is parts per million by volume ; ppmm is parts per million by mass

NOTE: often in literature v or m are omitted. In such case ppm units typically refer to volume mixing ratio unless noted in text otherwise.

Commonly used mixing fraction:

one part per million	1 ppm	( $1 \times 10^{-6}$ )
one part per billion	1 ppb	( $1 \times 10^{-9}$ )
one part per trillion	1 ppt	( $1 \times 10^{-12}$ )

**Problem:** A typical volume mixing ratio for ozone in urban air is  $r_{O_3} = 0.1$  ppmv. What is ozone mass mixing ratio? (take molecular weight of air  $m_a = 28.97$  g/mol)?

**Solution.**

Molecular weight of ozone is  $m_{O_3} = 48.0$  g mol<sup>-1</sup>. Therefore we have

$$r_{O_3,m} = m_{O_3} \cdot r_{O_3} / m_a = 48 \text{ (g mol}^{-1}\text{)} \cdot 0.1 \text{ ppmv} / 28.97 \text{ (g mol}^{-1}\text{)} = 0.17 \text{ ppmm}$$

thus

mass mixing ratio for ozone is  $r_{O_3,m} = 0.17$  ppmm

## 6. Gas laws.

Atmosphere is a gas mixture, therefore it obeys the gas laws. The behavior and properties of ideal gas mixtures is explained by kinetic molecular theory (part of physics).

- Atmospheric gases differ only slightly from the ideal gases, therefore ideal gas laws work in the atmosphere.

Postulates of the kinetic molecular theory:

An **ideal gas** consists of particles (molecules or atoms) that have the following properties:

1. The particles are so small compared with the distances between them that the volume of the individual particles can be assumed to be negligible (zero).
  2. The particles are constantly moving in a random way. They collide with each other and with the walls of the container producing the pressure exerted by the gas.
  3. The particles are assumed to exert no forces on each other (they are assumed neither to attract nor to repel each other).
  4. Collisions of the particles are perfectly elastic; no kinetic energy is lost during a collision. The average kinetic energy of a collection of gas particles is assumed to be directly proportional to the temperature of the gas.
- 
- A particular state of a gas is described by P (pressure); T (temperature); V (volume); and n (gas concentration) or mixing ratio. These variables are called state variables and are related by gas laws.

**Ideal gas law (also known as the equation of state):** says that the pressure exerted by a gas is proportional to its temperature and inversely proportional to its volume:

$$P V = \mu R T$$

where **R** is the universal gas constant. If pressure **P** is in atmospheres (atm), volume **V** in liters (L) and temperature **T** in degrees Kelvin (K), thus **R** has value

$$R = 0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1}$$

NOTE: T must be in degree Kelvin when used in the gas laws.

**Ideal gas law** is a combined form of three gas laws: Boyle's law, Charles's law, and Avogadro's law. These gas laws are based on experimental measurements of the properties of gases, and they can be explained with the kinetic molecular theory of gases.

**Boyle's law:**  $V \sim 1/P$  (at constant T and  $\mu$ )

NOTE: Robert Boyle (1627-1691), Irish chemist

**Charles's law:**  $V \sim T$  (at constant P and  $\mu$ )

NOTE: Jacques Charles (1746-1823), French physicist

**Avogadro's law:**  $V \sim \mu$  (at constant P and T)

NOTE: Amadeo Avogadro (1778-1850), Italian chemist

Avogadro postulated in 1811: that the equal volumes of gases at the same T and P contain the same number of molecules.

**Standard temperature and pressure (abbreviated STP) conditions:**

$$T = 0^{\circ}\text{C} = 273.2 \text{ K and } P = 1 \text{ atm}$$

**Problem:** What is the volume of 1 mole of ideal gas at STP?

**Solution.**

From the ideal gas law, the volume is given by

$$V = \mu R T / P = 1 \text{ (mol)} \cdot 0.08206 \text{ (L atm K}^{-1} \text{ mol}^{-1}) \cdot 273.2 \text{ (K)} / 1 \text{ (atm)} = \mathbf{22.42 \text{ (L)}}$$

This volume of **22.42** liters is called the molar volume of an ideal gas at STP.

**Dalton's law of Partial Pressures** states that for a fixed temperature and volume (container), the total pressure of a mixture of gases is the sum of the partial pressures that each of the individual component gases would have if placed alone in the same volume (container).

$$P = P_1 + P_2 + P_3 + \dots = \sum P_i$$

where  $P_i$  is the partial pressure of  $i$ -th gas. If each gas behaves ideally, the  $P_i$  can be calculated from the ideal gas law:

$$P_i = \mu_i R T / V$$

Thus, the total pressure of the mixture is

$$P = \sum P_i = \sum (\mu_i R T / V) = (RT/V) \sum \mu_i = \mu_{\text{tot}} (RT/V)$$

where  $\mu_{\text{tot}} = \sum \mu_i$  is the total number of moles of gases in the mixture.

**Mole fraction** is the ratio of the number of moles of a given component in a mixture to the total number of moles in the mixture:

$$\xi_i = \mu_i / \mu_{\text{tot}}$$

From the ideal gas law

$$\mu_i = P_i (V / RT)$$

Therefore,

$$\xi_i = \mu_i / \mu_{\text{tot}} = P_i (V / RT) / \{P (V / RT)\} = P_i / P$$

rearranging this expression as

$$P_i = \xi_i P$$

that is, the partial pressure of a particular component of a gaseous mixture is equal to the mole fraction of that component times the total pressure.

NOTE: mole fraction is equivalent to the volume fraction.

**Problem:** The mole fraction of nitrogen in the air is 0.7808. Calculate the partial pressure of N<sub>2</sub> in air when the atmospheric pressure is 760 torr.

**Solution.**

The partial pressure of N<sub>2</sub> can be calculated as follows:

$$P_{\text{N}_2} = \xi_{\text{N}_2} P = 0.7808 \cdot 760 \text{ (torr)} = 593 \text{ (torr)}$$

## The kinetic molecular theory of gases

derives relationships between state variables similar to that observed experimentally; and explains why gases behave in the observed fashion.

A main assumption in the **kinetic molecular theory of gases** is that the distribution of velocities of ideal gas molecules is described by the Maxwell-

Boltzmann velocity distribution law:

$$f(v) = 4\pi \left\{ \frac{m}{2\pi k_B T} \right\}^{3/2} v^2 \exp(-mv^2/2k_B T)$$

where

$v$  is velocity (m/s);  $m$  is mass of a gas molecule (kg),  $T$  is temperature (K), and  $k_B$  is Boltzmann's constant =  $1.38066 \times 10^{-23}$  (J/K).

NOTE: J denotes Joule which is unit of energy in SI system:  $J = \text{kg m}^2/\text{s}^2$

Analysis of the expression for  $f(v)$  yields the following equation for the most probable velocity  $v$  (the velocity possessed by the greatest number of gas particles):

$$v = (2k_B T / m)^{1/2}$$

Thus, the Maxwell-Boltzmann velocity distribution function implies that the most probable velocity of a molecule in a gas is directly proportional to the square root of the temperature of the gas and inversely proportional to the square root of the mass of the molecule.

It is easy to derive an expression for P from **kinetic molecular theory of gases** :

$$\mathbf{P = n k_B T}$$

where n is concentration of gas molecules.

Comparing this expression to one followed from the ideal gas law, we have

$$\mathbf{n k_B T = P = \mu R T / V}$$

then, substituting  $n = \mu N_A / V$ , we obtain

$$\mathbf{k_B \mu N_A / V = \mu R T / V}$$

therefore,

$$\mathbf{k_B = R / N_A}$$

NOTE: that gas constant and Boltzmann's constant are related by Avogadro's number:

$R = N_A k_B$ . In fact, it is useful to think of  $k_B$  as the gas law constant per molecule.

NOTE: R and  $k_B$  can be expressed in alternative units. For instance,

$$k_B = 1.38066 \times 10^{-23} \text{ (J/K)} = 1.38066 \times 10^{-19} \text{ (cm}^3 \text{ mb K}^{-1}\text{)} = \dots$$

$$R = 8.3145 \text{ (J mol}^{-1} \text{ K}^{-1}\text{)} = 8.314 \times 10^4 \text{ (cm}^3 \text{ mb mol}^{-1} \text{ K}^{-1}\text{)} = \dots$$

NOTE: the Maxwell-Boltzmann velocity distribution function gives the following equation for the mean velocity of a gas molecule:

$$\mathbf{v = (8k_B T / \pi m)^{1/2} = (8RT / \pi m_g)^{1/2}}$$

where m is the mass of a gas molecule,  $m_g$  is the molecular weight (or molar mass); and  $\pi$  is a mathematical constant  $\pi=3.141592\dots$ The mean velocity of a gas molecule is often used in various atmospheric chemistry applications.

**NOTE:** important to remember: the equation of state can be written in several forms:

using molar concentration of a gas,  $c = \mu/v$ :  $\mathbf{P = c T R}$

using number concentration of a gas,  $n = c N_A$ :  $\mathbf{P = n T R/N_A}$  or  $\mathbf{P = n T k_B}$

using mass concentration of a gas,  $q = c m_g$ :  $\mathbf{P = q T R /m_g}$

**NOTE:** be careful to write the quantities with the proper units and to select the appropriate value for R when you use the ideal gas law.

## **7. Dry air and moist air.**

Total atmospheric pressure can be divided into the partial pressure exerted by air ( $P_d$ ) and the partial pressure by water vapor ( $P_v$ ): thus

$$\mathbf{P_a = P_d + P_v}$$

Similar the number concentration of air molecules can be divided into a dry and moist component: thus

$$\mathbf{n_a = n_d + n_v}$$

where  $\mathbf{n_d}$  is the number concentration of dry air,  $\mathbf{n_v}$  is the number concentration of water vapor, and  $\mathbf{n_a}$  is the number concentration of air.

Equation of state for dry air:

$$P_d = \mu_d R T / V = n_d k_B T = \rho_d R_d T$$

where  $\rho_d$  is the mass density of dry air, and  $R_d$  is the gas constant for dry air defined as

$$R_d = R / m_d$$

Equation of state for water vapor:

$$P_v = \mu_v R T / V = n_v k_B T = \rho_v R_v T$$

where  $\rho_v$  is the mass density of water vapor, and  $R_v$  is the gas constant for water vapor

defined as  $R_v = R / m_{H_2O}$

NOTE: total air = moist air = dry air + water vapor

**Specific humidity** is the ratio of mass of water vapor to mass of moist air.

$$q_s = \rho_v / (\rho_v + \rho_d) = \rho_v / \rho_a$$

**Relative humidity** is defined as the ratio of the partial pressure of water vapor to its

saturation vapor pressure at a given temperature:

$$RH = 100 P_v / P_v^*$$

where the factor 100 is used because RH is usually expressed in percent, and  $P_v^*$  is the

saturation water vapor pressure, which is vapor pressure in the equilibrium state of the

water-water vapor system at a given temperature.

NOTE: RH=90% for very humid air (summer in Florida), RH=20 % for very dry air (summer in desert)

- Condensation of water vapor releases heat, which is called the **latent heat of condensation.**
- Evaporation of water vapor adsorbs heat, which is called the **latent heat of evaporation.**
- Latent heat is one of the principal energy sources in the atmosphere.