Water Vapor and Other Gases. 3

Terrestrial organisms live in a gaseous medium composed mostly of nitrogen and oxygen. Water vapor is present in varying amounts, and carbon dioxide and other gases, in trace amounts. Organisms exchange oxygen, carbon dioxide, and water vapor with their surroundings. Carbon dioxide is a substrate for photosynthesis and oxygen is a product; while oxygen is a substrate for respiration and carbon dioxide is a product. Exchange of these gases with the environment is therefore a requirement for life. Water vapor almost always moves from the organism to the environment. The humidity of the organism is near 100 percent, while the surroundings are nearly always much drier. The organism must remain in a highly hydrated state in order for biochemical reactions to occur, so the constant loss of water is a threat to survival, and frequent access to liquid water is a necessity for most terrestrial organisms. The intake of liquid water and the loss of water vapor to the environment are usually the most important components of the water budget of an organism.

Water loss is generally viewed as detrimental to the organism, though it may have some benefit in the circulation system of plants. It does, however, have a definite benefit when we consider the energy balance of the organism. Environmental temperatures are often higher than can be tolerated by biological systems. If there were not some mechanism for cooling the organism, it would perish. As water evaporates, roughly 44 kilojoules of energy are required to convert each mole to the vapor state. This is called the latent heat of vaporization. It is 580 times the energy required to change the temperature of one mole of water by one Celsius degree, and therefore represents an enormous sink for energy in the organism environment. Evaporative cooling is a natural way of controlling temperatures of organisms in hot environments. The amount of cooling available to the organism depends on the concentrations of water vapor at the organism surface and in the environment, and on the conductance to water vapor of the organism surface and boundary layer. This chapter discusses terminology for specifying gas concentrations and provides information about concentrations of gases in terrestrial environments.

3.1 Specifying Gas Concentration

Concentrations of the main atmospheric constituents are often expressed as percentages or volume fractions. For mainly historical reasons, which have to do with methods of measurement, water vapor concentration is expressed in a number of different ways. For reasons that will become clearer as we get farther into the subject, there is a substantial advantage to expressing concentrations of all gases in terms of mole fraction (moles of substance per mole of air), and fluxes as moles per square meter per second. The relationship between density or concentration and amount of substance j in a gas is

$$\rho_j = \frac{n_j M_j}{V} \tag{3.1}$$

where n_j is the number of moles, V is the volume of gas, and M_j is the molecular mass. Since the mole fraction of j is the ratio of moles of gas j to moles of air:

$$C_j = \frac{n_j}{n_a} = \frac{M_a \rho_j}{M_j \rho_a}. \tag{3.2}$$

Here, M_a is the molecular mass of air and M_j is the molecular mass of component j. Table 3.1 gives molecular masses for the main constituents of the atmosphere.

The molar density, or ratio ρ_j/M_j, is the same for all gasses. At standard temperature and pressure (STP; 0° C and 101.3 kPa) the molar density of any gas is 44.6 mol m^{-3} (one mole of any gas occupies 22.4 liters). The molar density of gas will show up a lot in our equations, so we give it the special symbol $\hat{\rho}$. The variation of molar density with pressure and temperature is given by the Boyle–Charles law which states that the volume of a gas is inversely proportional to its pressure (p) and directly proportional to its Kelvin temperature (T). Using the Boyle–Charles law the molar density of air can be computed from:

$$\hat{\rho} = 44.6 \frac{p}{101.3} \frac{273.15}{T}. \tag{3.3}$$

TABLE 3.1.   Properties of the major constituents of air.

Gas	Molecular Mass (g/mol)	Mol fraction in air	Density at STP (kg m^{-3})
Nitrogen	28.01	0.78	1.250
Oxygen	32.00	0.21	1.429
Carbon dioxide	44.01	0.00034	1.977
Water vapor	18.02	0 to 0.07	0.804
Air	28.97	1.00	1.292

A middle range temperature for biophysical calculations is 293 K (20° C), giving $\hat{\rho} = 41.4$ mol m^{-3} at sea level (101.3 kPa).

The relationship between volume, temperature, and pressure for a perfect gas is

$$p_j V = n_j RT \tag{3.4}$$

where p_j is the partial pressure of gas j, and R is the gas constant, 8.3143 J mol^{-1} K^{-1}. Substituting Eq. (3.4) into Eq. (3.2) gives

$$C_j = \frac{p_j}{p_a} \tag{3.5}$$

so the mole fraction of a gas can be calculated as the ratio of its partial pressure and the total atmospheric pressure.

One more relationship between mole fraction and other measures is useful. If two gases with initial volumes V_1 and V_2 are mixed to make a volume V_a and the pressure is the same on all three volumes, then the volume fraction, V_1/V_a is equal to the mole fraction, n_1/n_a. Gas concentrations in air are often expressed as percentages, parts per million (ppm), or parts per billion (ppb) on a volume basis (volume of the pure gas divided by the volume of air). It can be seen here that these measures are directly related to the mole fraction.

One more version of the perfect gas law is also useful. The density of a gas is the molecular mass multiplied by the number of moles, and divided by the volume occupied by the gas (Eq. (3.1)). Substituting this into Eq. (3.4) gives the relationship between the partial pressure of a gas and its concentration:

$$p_j = \frac{\rho_j RT}{M_j}. \tag{3.6}$$

Example 3.1. In 1985 the average concentration of CO_2 in the atmosphere of the earth was estimated to be 344 ppm. What is the mole fraction, partial pressure, and density (concentration) of atmospheric CO_2 in air at 20° C?

Solution. Parts per million (ppm) means volumes of CO_2 in 10^6 volumes of air. Since the volume ratio is equal to the mole fraction, 344 ppm is the same as 3.44×10^{-4} moles/mole or 344μ mol/mol. Using Eq. (3.5), $p_c = C_c p_a$. If $p_a = 101$ kPa, then $p_c = 3.44 \times 10^{-4} \times 1.01 \times 10^5$ Pa $= 35$ Pa. For density Eq. (3.6) is used and rearranged to get

$$\rho_c = \frac{p_c M_c}{RT} = \frac{35 \text{ Pa} \times 44 \frac{g}{mol}}{8.31 \frac{J}{mol\,K} \times 293 \text{ K}} = 0.63 \frac{g}{m^3}.$$

To get the units to divide out, you may need to refer to Table 1.2. Note that a Pascal is a Newton per square meter, and that a joule is a Newton-meter. A Pascal is therefore equivalent to a joule per cubic meter.

3.2 Water Vapor: Saturation Conditions

If a container of pure water is uncovered in an evacuated, closed space, water will evaporate into the space above the liquid water. As water evaporates, the concentration of water molecules in the gas phase increases. Finally, an equilibrium is established when the number of molecules escaping from the liquid water equals the number being recaptured by the liquid. If the temperature of the liquid was increased, the random kinetic energy of the molecules would increase, and more water would escape. The equilibrium vapor pressure, established between liquid water and water vapor in a closed system is known as the saturation vapor pressure for the particular temperature of the system. The saturation vapor pressure is the highest pressure of water vapor that can exist in equilibrium with a plane, free water surface at a given temperature. The saturation vapor pressure is shown as a function of temperature in Fig. 3.1. Since the influence from any other gases present in the space above the water is negligible, the vapor pressure above the water surface is essentially the same whether the closed space is initially evacuated or contains air or other gases. The symbol e is used to represent the vapor pressure of water, and the saturation vapor pressure is denoted by $e_s(T)$, indicating that the saturation vapor pressure is determined by temperature. Tables giving saturation vapor pressure as a function of temperature can be found in List (1971) and in Table A.3 of the Appendix of this book. The mole fraction, which we frequently use in future computations, depends on both temperature and pressure. It is computed using Eq. (3.5), by dividing $e_s(T)$ by 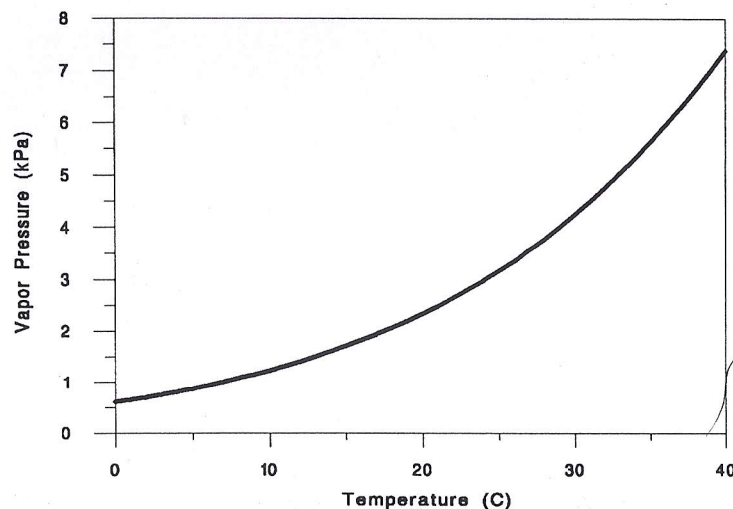the atmospheric pressure. The main variable determining atmospheric pressure is elevation. A relationship, which is accurate enough for biophysical calculations, is

$$p_a = 101.3 \exp\left(\frac{-A}{8200}\right) \tag{3.7}$$

where A is the altitude in meters above sea level and p_a is in kPa.

A convenient, empirical equation (with a close connection to the Clausius Clapeyron equation from thermodynamics) for computing the saturation vapor pressure from temperature is the Tetens formula (Buck, 1981):

$$e_s(T) = a \exp\left(\frac{bT}{T+c}\right) \tag{3.8}$$

where T is the Celsius temperature. The constants a, b, and c can be chosen to optimize the fit of the equation for various ranges of data. For environmental biophysics applications the constants are $a = 0.611$ kPa, $b = 17.502$, and $c = 240.97°$ C. Equation (3.8) can be used in place of tables for finding the saturation vapor pressure. While there are slight differences between the values from Eq. (3.8) and the more generally accepted values in List (1971), the differences are not measurable, nor are they significant for biophysical computations. The Tetens formula can also be used to predict the vapor pressure over ice (which is different from the vapor pressure over water). The coefficients for ice are $b = 21.87$ and $c = 265.5°$ C.

The slope of the saturation mole fraction with respect to temperature is also used frequently in computations. It is obtained by dividing the slope of the saturation vapor pressure function by atmospheric pressure. The slope of the saturation vapor pressure function is obtained by differentiating Eq. (3.8) to obtain

$$\Delta = \frac{bc\, e_s(T)}{(c+T)^2}. \tag{3.9}$$

The slope of the saturation mole fraction is represented by s, and is given by

$$s = \Delta/p_a. \tag{3.10}$$

Example 3.2. Find the saturation vapor pressure at 0, 10, 20, and 30° C, and the mole fraction of water vapor in saturated air at sea level for each of these temperatures.

Solution. When $T = 0$, $\exp(0) = 1$, so Eq. (3.8) gives $e_s(0) = a = 0.611$ kPa. The others require a little more computation, but, using Eq. (3.8) they give $e_s(10) = 1.23$ kPa, $e_s(20) = 2.34$ kPa, and $e_s(30) = 4.24$ kPa. Comparing these to the values in Table A.3 shows them to agree to the number of significant digits shown here. The pressure at sea level is 101 kPa, so the mole fraction at 0° C is

FIGURE 3-1. Saturation vapor pressure of air as a function of temperature.

$C_v = 0.611\,\text{kPa}/101\,\text{kPa} = 0.006$. The other mole fractions are 0.012, 0.023, and 0.042 mol/mol or 6, 12, 23, and 42 mmol/mol.

There are a couple of points from the example that are worth noting. Vapor pressures and mole fractions are used in many calculations throughout this book, so it would be a good idea to find some way of remembering approximate values for these quantities at saturation. Note that the vapor pressure approximately doubles for each 10° C temperature increase. Exactly doubling would give 0.6, 1.2, 2.4, and 4.8 kPa for the four temperatures. All but the last one are within a few percent of the actual value, and it is only a little over ten percent high. If you can remember the vapor pressure at zero, you can therefore estimate saturation vapor pressures at higher temperatures. The other point to note is that conversion to mole fraction at sea level involves division by a number close to 100. The mole fraction, expressed as a percent, is therefore nearly the same as the vapor pressure. This is also the fraction of a saturated atmosphere made up of water vapor at the indicated temperature. The mole fraction, expressed in mmol/mol is just the vapor pressure in kPa multiplied by 10 (and is equal to the vapor pressure in millibars).

3.3 Condition of Partial Saturation

In nature, air is seldom saturated, so we need to know more than just the temperature to specify its moisture condition. Partial saturation can be expressed in terms of ambient vapor pressure or mole fraction, relative humidity, vapor deficit, dew point temperature, or wet bulb temperature. Ambient vapor pressure is simply the vapor pressure that exists in the air, as opposed to saturation vapor pressure, which is the maximum possible vapor pressure for the temperature of the air. Relative humidity is the ratio of ambient vapor pressure to saturation vapor pressure at air temperature:

$$h_r = \frac{e_a}{e_s(T_a)}. \tag{3.11}$$

Relative humidity is sometimes multiplied by 100 to express it as a percent rather than a fraction, but it is always expressed as a fraction in this book.

The relationship between saturation vapor pressure and ambient vapor pressure at various humidities is shown in Fig. 3.2. The curved lines, labeled on the right, show vapor pressures at humidity increments of 0.1 for temperatures from 0 to 40° C.

The vapor deficit is the difference in vapor pressure or mole fraction between saturated and ambient air:

$$D = e_s(T_a) - e_a = e_s(T_a)(1 - h_r) \tag{3.12}$$

where the second relation follows from Eq. (3.11). The vapor deficit at any temperature and relative humidity is the difference, in Fig. 3.2, between the saturation line ($h_r = 1$) at T_a and the line for the ambient relative humidity.

The dew point temperature is the temperature at which air, when cooled without changing its water content or pressure, just saturates. In other
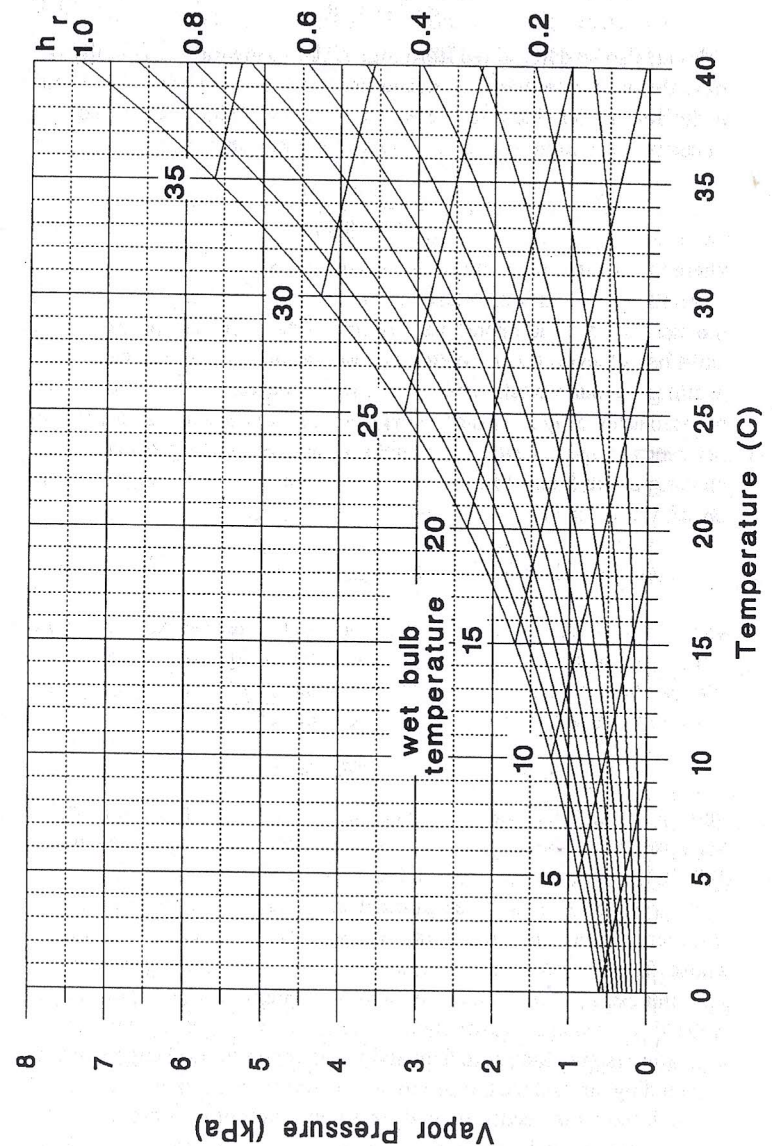

FIGURE 3.2. Vapor pressure-temperature-relative humidity–wet bulb temperature diagram. Wet bulb lines are for sea level pressure.

words, the saturation vapor pressure at dew point temperature is equal to the ambient vapor pressure:

$$e_s(T_d) = e_a. \qquad (3.13)$$

This can also be determined from Fig. 3.2 by following horizontally from the ambient vapor pressure, and reading the temperature at the intersection of the horizontal line and the $h_r = 1$ line. It can be obtained more precisely from Table A.3, and by inverting Eq. (3.8):

$$T_d = \frac{c \ln(e_a/a)}{b - \ln(e_a/a)} \qquad (3.14)$$

where the constants are the same as in Eq. (3.8).

Another important moisture variable is wet bulb temperature, T_w. To find the wet bulb temperature, determine the temperature drop which could be achieved by adiabatic evaporation of water into air (adiabatic meaning "without heat exchange"). Air is cooled by evaporating water into it, but the evaporation of water into the air raises its vapor pressure. Since the change in heat content of the air due to changing its temperature must equal the latent heat of evaporation for the water evaporated into the air, we can write:

$$c_p(T_a - T_w) = \lambda \left[C_s(T_w) - C_{va} \right] = \frac{\lambda \left[e_s(T_w) - e_a \right]}{p_a} \qquad (3.15)$$

where λ is the latent heat of vaporization of water (44 kJ/mol) and c_p is the specific heat of air (29.3 J mol^{-1} K^{-1}). Equation (3.15) is most often written in terms of vapor pressure and used for determining vapor pressure from wet bulb and dry bulb temperatures:

$$e_a = e_s(T_w) - \gamma p_a(T_a - T_w). \qquad (3.16)$$

Here, $\gamma = c_p/\lambda$ is called the thermodynamic psychrometer constant. It has a value of 6.66×10^{-4} C^{-1} with a slight temperature dependence (0.01%/C) due to the temperature dependence of λ.

A psychrometer is an instrument consisting of two thermometers. One thermometer measures the air temperature. The second thermometer, whose "bulb" is covered with a wet cotton wick, measures the wet bulb temperature, T_w. Equation (3.16) is used to determine the vapor pressure of the air from these measurements. Clearly, a real psychrometer is not an adiabatic system since both heat and water vapor are exchanged with the surrounding air, and the thermometers absorb and emit radiation. Later in the book the tools needed to analyze a real psychrometer are developed, but the result of such an analysis yields an equation like Eq. (3.16) with an apparent psychrometer constant γ^*. For an adequately ventilated psychrometer with a good wick and radiation shield the value of the apparent psychrometer constant is close to the thermodynamic constant, but poorly designed or ventilated psychrometers can have much lower constants.

Equation (3.16) defines the family of straight, diagonal lines shown in Fig. 3.2. The wet bulb temperatures are labeled along the $h_r = 1$ line.

Either Eq. (3.16) or Fig. 3.2 can be used to find vapor pressure from wet and dry bulb temperatures, later we give examples of both. Finding wet bulb temperature from vapor pressure and air temperature using Fig. 3.2 is also easily done. Finding T_w from the psychrometer equation, however, is somewhat more challenging, since T_w appears in the linear term and implicitly in $e_s(T_w)$. There is no explicit solution to this equation, but a solution can be found by standard mathematical methods for solving nonlinear equations.

Example 3.3. When air temperature is 23° C and vapor pressure is 1.1 kPa, use Fig. 3.2 to find relative humidity, dew point temperature, and wet bulb temperature.

Solution. Enter Fig. 3.2 at $T = 23°$ C and $e_a = 1.1$ kPa. This point is just below the line for $h_r = 0.4$, so the humidity is estimated to be 0.39. The dew point temperature is the intersection of the $h_r = 1$ line and $e_a = 1.1$ kPa. The dew point temperature is therefore around 8° C. To find the wet bulb temperature, place a straight edge on Fig. 3.2 with the edge parallel to the wet bulb lines and passing through the point $T = 23°$ C, $e_a = 1.1$ kPa. Read the temperature on the scale along the $h_r = 1$ line, it is approximately 14.5° C. Note that the wet bulb temperature always lies between the air temperature and the dew point temperature unless the humidity is 1; then they are all equal.

Example 3.4. Find the vapor pressure and vapor mole fraction at the surface of melting snow at 1300 m elevation.

Solution. The temperature of melting ice and snow is 0° C. Since the surface is pure water, the vapor pressure at the surface is the saturation vapor pressure for that surface temperature. From Table A.3 the saturation vapor pressure at 0° C is 0.611 kPa. Using Eq. (3.7), the pressure at 1300 m elevation is $101.3 \exp(-1300/8200) = 86$ kPa. The vapor mole fraction is $C_v = 0.611$ kPa/86 kPa $= 0.007$ mol/mol or 7 mmol/mol.

Example 3.5. A psychrometer gives an air temperature of 30° C and a wet bulb temperature of 19° C. Use the psychrometer equation and other formulae to find the vapor pressure, relative humidity, dew point temperature, and vapor deficit of the air. Assume the measurement is made at sea level.

Solution. To use the psychrometer equation the saturation vapor pressure at wet bulb temperature needs to be known. Table A.3 gives $e_s(T_w) = 2.20$ kPa. The saturation vapor pressure at air temperature is also needed to find the humidity. Again, from Table A.3 it is $e_s(T_a) = 4.24$ kPa. Using the psychrometer equation (Eq. (3.16)) gives

$$e_a = 2.20 \text{ kPa} - 101 \text{kPa} \times 6.67 \times 10^{-4} \text{C}^{-1}(30 \text{ C} - 19 \text{C}) = 1.46 \text{ kPa}.$$

Now, using Eq. (3.11), $h_r = 1.46\,\text{kPa}/4.24\,\text{kPa} = 0.345$. The vapor deficit (Eq. (3.12)) is $D = 4.24\,\text{kPa} - 1.46\,\text{kPa} = 2.78\,\text{kPa}$. The dew point temperature can be obtained from Table A.3 or from Eq. (3.14). In Table A.3, follow down the vapor pressure column until you find the vapor pressure just smaller than 1.46 kPa. This is at 12° C. The dew point temperature therefore lies between 12° C and 13° C. The difference between the value at 12° C and 1.46 kPa is 58 Pa; the difference between the value at 12 and 13° C is 95 Pa. Our value of 1.46 kPa is therefore 58/95 or 0.6 of the way between 12 and 13° C. The dew point temperature is therefore $T_d = 12.6°$ C. Using Eq. (3.14):

$$T_d = \frac{240.97 \ln \frac{1.46}{0.611}}{17.502 - \ln \frac{1.46}{0.611}} = 12.62°\text{C}.$$

Example 3.6. The vapor pressure of air at a ski resort is 0.4 kPa. A machine to make artificial snow emits water droplets which quickly cool to the wet bulb temperature. At what air temperature will the droplets just reach freezing temperature at an altitude of 2400 m?

Solution. Ice melts at 0° C, but often cools a few degrees below this temperature before freezing (super cools). For this problem we will assume a freezing temperature of −2° C. Rearranging the psychrometer equation (Eq. (3.16)) to find air temperature gives

$$T_a = T_w + \frac{e_s(T_w) - e_a}{\gamma p_a}$$

$$= -2 + \frac{0.53\text{kPa} - 0.4\text{kPa}}{6.67 \times 10^{-4}\text{C}^{-1} \times 0.75 \times 101\text{kPa}} = 0.6°\text{C}.$$

Therefore, expect the droplets to start freezing at an air temperature slightly above zero.

Example 3.7. A humidity sensor at air temperature reads 0.23 when the air temperature is 16° C. What is the vapor pressure, dew point temperature, and wet bulb temperature?

Solution. Entering Fig. 3.2 at the $T = 16°$ C line and $h_r = 0.23$ gives a vapor pressure of about 0.4 kPa. The wet bulb temperature is obtained by moving along a line through the vapor pressure, temperature point, parallel to the psychrometer lines. It is approximately 7.5° C. The dew point temperature cannot be obtained from Fig. 3.2 because it does not go low enough. Table A.3 also fails to go low enough for this problem. Before turning to the formula to obtain this value, try calculating a more precise value for the vapor pressure. From Table A.3, the saturation vapor pressure at 16° C (air temperature) is 1.82 kPa. Rearranging Eq. (3.11) to compute vapor pressure gives:

$$e_a = h_r e_s(T_a) = 0.23 \times 1.82\,\text{kPa} = 0.42\,\text{kPA}.$$

TABLE 3.2. Comparison of various measures of moisture in air for air temperature = 20° C, relative humidity = 0.5, and atmospheric pressure = 100 kPa.

$e_s(20)$ (kPa)	e_a (kPa)	T_d (C)	T_w (C)	C_{va} (mmol/mol)	ρ_v (gm^{-3})	r (g/kg)	q (g/kg)
2.34	1.17	9.3	14.0	11.7	8.65	7.36	7.31

Now, using Eq. (3.14) the dew point temperature is obtained:

$$T_d = \frac{240.97 \ln \frac{0.42}{0.611}}{17.502 - \ln \frac{0.42}{0.611}} = -5.1\,\text{C}.$$

Three other quantities are commonly used by meteorologists to describe vapor concentration in air. These are the absolute humidity, the mixing ratio, and the specific humidity. The absolute humidity, also known as the vapor density, is the mass of water vapor per unit volume of air. It is related to the vapor pressure by Eq. (3.6). For water vapor this becomes

$$\rho_v = \frac{e_a M_w}{RT} \tag{3.17}$$

where e_a has units of Pa. If e_a has units of kPa, and ρ_v has units of g m^{-3}, then for $\mathbf{T} = 293$ K, $\rho_v = 7.4 e_a$.

The mixing ratio r is the mass of water vapor per unit mass of dry air. It can be computed from the mole fraction of water vapor using

$$r = \frac{0.622 C_{va}}{1 - C_{va}}. \tag{3.18}$$

The specific humidity q is the mass of water vapor divided by the mass of moist air, and is related to the mole fraction by:

$$q = \frac{0.622 C_{va}}{1 - 0.378 C_{va}}. \tag{3.19}$$

The units of r and q usually are g/kg. Table 3.2 compares values of r and q with sample values of the other variables representing moisture in air.

3.4 Spatial and Temporal Variation of Atmospheric Water Vapor

The spatial and temporal patterns of vapor pressure in the atmosphere resemble those given in Ch. 2 for temperature. During the day, vapor pressures are highest near a soil or plant surface and decrease with height. At night, vapor pressures tend to be lowest near the surface and increase with height. Vapor pressures tend to be a bit higher in the day than at night and typically reach a minimum at the time temperature is at the minimum. As with temperature, the surface acts as a source of water vapor in the day

and a sink at night (when condensation and dew formation occur), and is therefore responsible for the shape of the vapor pressure profiles. Because the surface acts as a source or sink, and water vapor is transported in the atmosphere, there also exist high frequency random fluctuations in vapor pressure like those for temperature shown in Fig. 2.4.

While these patterns are easily demonstrated, the magnitude of spatial and temporal vapor pressure variation is much smaller than for temperature, and is usually small enough that it can be ignored in comparison with other sources of uncertainty in the measurements. If only the average vapor pressure for a day is known, the best estimate of hourly vapor pressures is that they equal the average for the day. Variation of vapor pressure with height can be described by an equation similar to Eq. (2.1), so a log plot of two or more measured vapor pressures with height would allow extrapolation or interpolation to other heights, as was done with temperature. However, the changes in vapor pressure with height are relatively small, so vapor pressures in an organism microenvironment are similar to the the vapor pressure at measurement height.

Not all measures of atmospheric moisture are as well behaved as vapor pressure or mole fraction, however. Figure 3.3 shows the diurnal variation in relative humidity and vapor deficit for the temperatures in Fig. 2.2, assuming the vapor pressure is constant throughout the day at 1.00 kPa ($T_d = 7°$ C). Note that the humidity is near one and the vapor deficit is near zero early in the morning. In the early afternoon the humidity is around 0.3 and the vapor deficit is 2 kPa. All of this variation is

brought about by the change in temperature, with no change in the vapor concentration in the air.

It is perhaps unfortunate that one of the most common measurements of atmospheric moisture is relative humidity. The measurement itself is essentially useless as an environmental variable except as a means, along with air temperature, of obtaining the vapor pressure, mole fraction, or dew point temperature. Some people compute and report averages of humidity over time periods of a day or longer. It should be clear from Fig. 3.3 that an average humidity is worse than meaningless. In addition to failing to communicate any useful information by itself, averaging individual humidity measurements destroys any possibility of obtaining useful information from the original data because the average humidity depends on the pattern of temperature variation (which is lost in the averaging process). It is best to immediately convert humidity data to vapor pressure or dew point. Then record, average, and process these data.

Averaging the vapor deficit is a slightly different matter. We show later that the vapor deficit gives an estimate of the driving force for evaporation, and is useful in relating transpiration and biomass production in plant communities. The average vapor deficit for the atmosphere is therefore a useful number, but it can be estimated reliably from average vapor pressure and temperature.

3.5 Estimating the Vapor Concentration in Air

Reliable measurements of atmospheric moisture are difficult to obtain, but estimates of the vapor pressure, which are quite reliable, are relatively easy to make. In the absence of airmass changes and advection, the vapor pressure in the air is relatively constant throughout the day and from day to day. It also varies little between indoors and outdoors. Figure 3.3 shows the humidity going to nearly 1.0 at the time of minimum temperature. This behavior is typical of all except arid, summer conditions. Therefore, the minimum daily temperature can often be taken as the dew point temperature.

Example 3.8. Summer minimum temperatures at locations in the Midwestern U.S. can be around 20° C, while in the arid Southwest they are 10° C or below. Compare the vapor pressure of the two locations.

Solution. Assuming the minimum temperature equals the dew point temperature, the vapor pressure can be looked up either in Fig. 3.2 or Table A.3. At 20° C it is 2.3 kPa and at 10° C it is 1.2 kPa. That difference makes an enormous difference in human comfort when temperatures are in the high 30s (C).
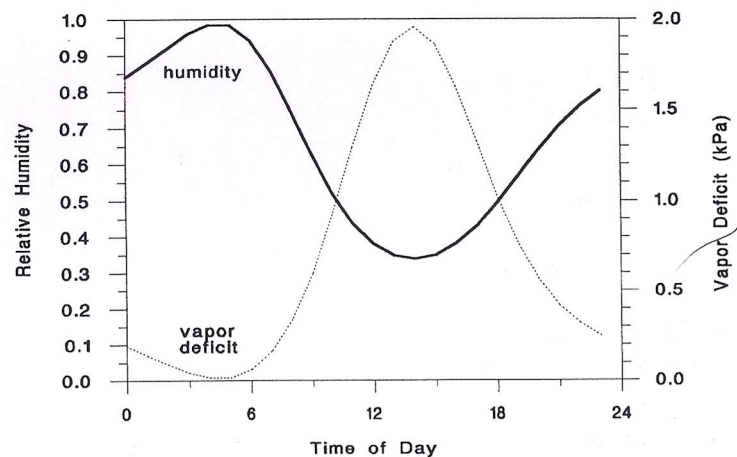

FIGURE 3-3. Diurnal variation in relative humidity and atmospheric vapor deficit for the temperature variation in Fig. 2.2. Vapor pressure is assumed to be constant throughout the day at 1.00 kPa.

Example 3.9. On a particular foggy, cold day the outdoor temperature is $-20°$ C. Compare the humidity outdoors with the humidity in a heated building where air temperature is $22°$ C.

Solution. If there is fog, the humidity outdoors must be 1.0. In order to find the humidity indoors the vapor pressure of the air and the saturation vapor pressure at the indoor air temperature are needed. Assume that the indoor and outdoor vapor pressures are equal. Neither our figure nor our table give values at $-20°$ C, so we use Eq. (3.8):

$$e_a = 0.611 \exp \frac{17.502 \times (-20)}{240.97 - 20} = 0.125 \text{ kPa}.$$

The saturation vapor pressure at $22°$ C is 2.64 kPa (Table A.3), so the indoor humidity is $h_r = 0.125/2.64 = 0.05$. Note that the same amount of vapor in the air gives very different humidities in the two environments.

Vapor pressure or humidity normally is measured in weather stations. Typically such measurements are taken 1.5 to 2 m above the ground in an open area. The humidity of microenvironments in plant canopies or near leaves can be quite different. For example, a tiny mite living on a corn leaf may be experiencing a relative humidity of 70 percent when the humidity measurement at an adjacent weather station indicates 30 percent. This can occur because the mite is small enough to be reside between the leaf surface and the top of the boundary layer surrounding the leaf, so that moisture from the transpiration stream leaving the stomata humidifies the mite environment. This humidification of leaf boundary layers is also important to transpiration because the conductance of the leaf surface (stomatal conductance) is influenced by the humidity in this leaf boundary layer. Humidity can vary greatly among various microenvironments, and we study such effects in following chapters.

References

Buck, A. L. (1981) New equations for computing vapor pressure and enhancement factor. J. Appl. Meteorol. 20:1527-1532.

Geiger, R. (1965) The Climate Near the Ground. Cambridge, Mass.: Harvard University Press.

List, R. J. (1971) Smithsonian Meteorological Tables, 6th. ed. Washington, D. C.: Smithsonian Institution Press.

Problems

3.1. A psychrometer gives air temperature of $34°$ C and wet bulb temperature of $22°$ C. Find the vapor pressure, the vapor mole fraction, the dew point temperature, the relative humidity, and the vapor deficit. The altitude of the site is 1200 m.

3.2. On a hot, humid day your skin surface can be at $36°$ C and covered with perspiration. If the perspiration has the same vapor pressure as pure water, what is the vapor pressure at your skin surface?

3.3. You decide to calibrate a relative humidity sensor at a value of 0.5 in a 0.1 m^3 container. Assuming no water is adsorbed on the walls of the container, how much liquid water must be evaporated into dry air to achieve this humidity at container temperatures of 10 and $40°$ C?

3.4. The method in problem 3.3 is not very useful for calibrating humidity sensors. A better method is to bubble air through water at controlled temperatures. Suppose air is bubbled through water at $10°$ C, and then passed into the container with the humidity sensor, which is at $20°$ C. What is the relative humidity of the air in the chamber? If the air and humidity sensor were really at $21°$ C when you thought they were at $20°$ C, how much error would this cause in your calibration?

3.5. If the outdoor minimum temperature were $-15°$ C on a particular day, estimate the relative humidity in a $22°$ C laboratory on that day.

3.6. The vapor pressure of the air you breathe out is the saturation vapor pressure at body temperature ($37°$ C). If the air you breathe in is at $20°$ C and 0.2 relative humidity, you take 15 breaths per minute, and each breath has a volume of 1 liter, how many 250 ml. glasses of water are required per day to replenish this water loss?

3.7. If air is at $15°$ C and 0.6 relative humidity, find the absolute humidity (or vapor density), the specific humidity, and the mixing ratio.