

Potential temperature and the concept of thermodynamic instability in atmosphere

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This lecture is voted to introduce students to the basic thermodynamic of atmosphere. The meaning of pressure and temperature is defined by means of the classical kinetic theory of matter. After the definition of pressure and temperature the concept of ideal gas, internal energy and work are introduced as well. Internal energy, work and heat are then linked together in the fundamental first principle of thermodynamics. The first principle of thermodynamics is written in the most useful atmospheric form and the specific heats (constant pressure and volume) with their relationship with the gas constant. The variable enthalpy is introduced as well. The attention is then passed to the adiabatic processes and the poisson equation for adiabatic processes is derived. As a consequence the adiabatic lapse rate is introduced and the heuristic concept of stability/instability in atmosphere with some examples taken from real cases. Potential temperature is then introduced as a tool to easily identify the status of static thermodynamic instability.

1. Introduction

Before to start the discussion on potential temperature, it is worth to spend a few words examining its related (and basic) concept of “temperature”. In fact what is temperature and why is it important in physics? This question is not as simple as it might seem and, maybe for this reason, many deservedly renowned books of meteorology give the answer simply saying that “temperature is what is measured by a thermometer”. This answer is undoubtedly right, but it has the disadvantage that it push all the troubles on the naturally following question “what is a thermometer?”. To deal with the definition of temperature. facing both the above mentioned questions, here is adopted the historical

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approach, looking for the reasons why this concept had been introduced and how it happened. As a starting point we choose the semantic analysis. “Temperature” is a noun that comes from the ancient Latin and means “mixing of heat and cool” while “thermometer” comes from the ancient Greek and means “measurer of heat”. Keeping in mind the origin of those names, as well as their current scientific meaning, both “temperature” and “thermometer” appear as two undoubtedly wrong choices. In fact “temperature” is not a mixing of “heat and cool” and a “thermometer” does not measure heat at all. In any case the origin of the nouns “temperature” and “thermometer” gives quite a good idea of the reason why they had been adopted in the past. In particular the natural philosophers noticed that some bodies and the environment in which they lived appeared sometimes warm and sometimes cold, moreover this characteristic was dependent from the person who experienced it. Galileo, trying to quantify this sensation, developed a “thermometer” †. Without going into details (see the historical notes) what Galileo developed was an instrument that make visible the changes of a sensitive media (air) due to the warming and cooling. In particular Galileo noticed that the volume of air increased (decreased) with the increasing (decreasing) of temperature anticipating the experiments of Gay-Lussac and Boyle. Now, after several centuries of experiments and theories, we know that this happens because the motion of the molecules that compose air increase (decrease) their average kinetic energy when thermal energy is given (withdrawn) to the system that they represent. This makes possible to answer to the above question “what is temperature?” in fact we can say that “temperature is a variable related to the average kinetic energy of the constituents of matter”. Before to proceed is important to explain the meaning and importance of the word “average”. In fact if the constituents of matter are too few to make reasonable the concept of “average” then even the related “temperature” can be meaningless. This qualification is not simply an academic one, in fact in the very upper atmosphere this happens. For an example of how temperature and density change with height see Figure 1.

The above definition of temperature has deep roots on the kinetic theory of matter (in particular of aeriforms) according which the average translation kinetic energy of a molecule is given by

† The correct name of the instrument is “thermoscope” because it make visible the changes of temperature.

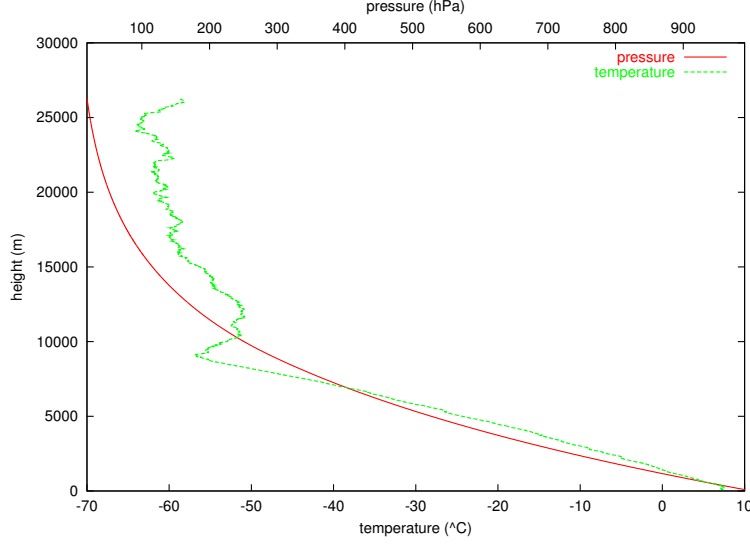


FIGURE 1. Temperature and pressure as a function of height. The data are obtained through the Campofornido atmospheric sounding carried out at 00 UTC of 06th December 2005.

$$\left\langle \frac{1}{2}mv^2 \right\rangle = \frac{3}{2}kT \quad (1.1)$$

while pressure is given by

$$p = \frac{1}{3} \frac{N}{V} \langle mv^2 \rangle \quad (1.2)$$

where m is the molecule mass, v the modulus of its velocity, k is the Boltzmann's constant, p is the pressure, N is the number of molecules/particles and T is the Kelvin's (or absolute) temperature[†]. One of the advantages of this definition is that, for an ideal gas, it gives immediately evidence of the reason why pressure and temperature should be related and why naturally you would expect that putting in contact a warm and cold gas you should obtain a mild mixture, i.e. a flow of heat from the warm gas toward the colder one. Moreover these two equations can be used to obtain the famous *ideal gas law*, in fact equating the $\langle mv^2 \rangle$ we have

$$pV = NkT \quad (1.3)$$

[†] The symbol adopted for the absolute temperature is K without any "°". On the contrary the symbol of Celsius temperature is $^{\circ}C$, where the "°" is strictly connected to the "C" and not to the numerical value

that, using the Avogadro's number \aleph and the number of moles n , can be written in the most common form

$$pV = n\aleph kT = nR^*T \quad (1.4)$$

or, introducing again the molecule/particle mass, in the meteorological form

$$p = \frac{n\aleph m}{V} \frac{k}{m} T = \frac{nm}{V} \frac{\aleph k}{m} T = \rho RT \quad (1.5)$$

Pay attention in not confusing R^* and R in Eq. 1.4 and 1.5, the first is the “ideal gas constant”, which refers to all the gases and vapors that behave like an ideal gas, the second is the “specific gas constant” that refers only to a specific gas or vapor (e.g. the R of the nitrogen is different from that of oxygen). Their difference is not a mere matter of numbers but essentially of dimensions, then they are conceptually different.

From what said above, a possible answer concerning the reason of the importance of temperature in atmospheric sciences is that it triggers the exchange of heat (thermal energy). At the end of this lecture most of the fundamental implications of this answer should (hopefully) result clearly.

2. The first law of thermodynamics

In section 1 it has been said that temperature is related to the average kinetic energy of the constituents of matter (then even of aeriforms). This average kinetic energy is not immediately recognizable as a form of motion, in fact volumes of air in a room are (seem) at rest relatively to the walls, for this reason it is usually called “internal kinetic energy” or simply “internal energy” of the system. This fact can be easily shown using the simple momentum equation for a single particle

$$m \frac{d\mathbf{v}}{dt} = \mathbf{F} \quad (2.1)$$

applying the scalar product with \mathbf{v} to the above equation, we have

$$m\mathbf{v} \cdot \frac{d\mathbf{v}}{dt} = \mathbf{v} \cdot \mathbf{F} \quad (2.2)$$

that, assuming that $\chi = \frac{1}{2}mv^2$ is the kinetic energy of the particle and that its mass is constant, becomes

$$\frac{d\chi}{dt} = \mathbf{v} \cdot \mathbf{F} \quad (2.3)$$

If we have a system composed by several particles, the equation 2.3 becomes

$$\sum_i^N \frac{d\chi_i}{dt} = \sum_i^N \sum_{j \neq i}^{N-1} \mathbf{v}_i \cdot \mathbf{F}_{ij} + \sum_i^N \mathbf{v}_i \cdot \mathbf{F}_i \quad (2.4)$$

where \mathbf{F}_{ij} represents the forces exerted by the particle j on the particle i and \mathbf{F}_i the external force acting on the particle i . In general we can assume that the sum of all the forces \mathbf{F}_{ij} is null, in fact for an ideal gas, the articles interact only through elastic bumps, then the force exerted by one particle is exactly the opposite of the reaction force exerted by the other (in formulas $\mathbf{F}_{ij} = -\mathbf{F}_{ji}$). Nevertheless, even if the sum of these “internal” forces is null, the term $\sum_i^N \sum_{j \neq i}^{N-1} \mathbf{v}_i \cdot \mathbf{F}_{ij}$ is not in general null because of the scalar product with the particle’s velocity. If the number of particles is enough high to give meaning to the average, then for an ideal gas we can define $\sum_i^N \chi_i = U$ as the *internal energy*, $\sum_i^N \mathbf{v}_i \cdot \mathbf{F}_i = W$ as the *work* and $\sum_i^N \sum_{j \neq i}^{N-1} \mathbf{v}_i \cdot \mathbf{F}_{ij} = Q$ as the *heat* of the system and the equation 2.4 becomes

$$\frac{dU}{dt} = Q + W \quad (2.5)$$

The quantity U could, in principle, depend even from the position in space of the molecules/particles, that is from the x direction

$$\frac{d}{dx} \left(\sum_i^N \chi_i \right) = \sum_i^N \frac{1}{2} m_i \frac{d}{dx} (v_i(x, y, z)^2) \neq 0 \quad (2.6)$$

and the same for y and z but if we assume that the kinetic energy and then velocity does not depend from the distance between particles or, in other words, that particles do not interact apart from the mutual bumps, internal energy can be considered as independent from the distance between particles and then independent from the volume. The above assumption is called *ideal gas assumption* and it is extremely important and useful in the atmospheric sciences.

The Eg. 2.5 can be written in a different and more useful form keeping with some physical considerations. In particular the term $\sum_i^N \mathbf{v}_i \cdot \mathbf{F}_i = W$ can be written, remembering that $\mathbf{v}_i = \frac{dl_i}{dt}$ with l_i i^{th} displacement, $\sum_i^N \frac{dl_i}{dt} \cdot \mathbf{F}_i = W$. If we multiply the time

derivative by the following expression $\sigma_i \cdot \frac{\sigma_i}{\sigma_i^2}$, where σ_i is the oriented section of the i^{th} particle, we will have

$$W = \sum_i^N \sigma_i \cdot \frac{\sigma_i}{\sigma_i^2} \frac{d \cdot l_i}{dt} \cdot \mathbf{F}_i = \sum_i^N \frac{d}{dt} (\sigma_i \cdot l_i) \frac{\sigma_i \cdot \mathbf{F}_i}{\sigma_i^2} \quad (2.7)$$

If we assume $p = p_i \equiv -\frac{\sigma_i \cdot \mathbf{F}_i}{\sigma_i^2}$, pressure exerted on the system, and $V_i \equiv \sigma_i \cdot l_i$ volume spanned by the i^{th} particle in the time dt , thanks to the linearity of derivatives, equation 2.5 becomes

$$\frac{dU}{dt} = Q - p \frac{dV}{dt} \quad (2.8)$$

A few words have to be spent concerning the minus sign adopted into the definition of p_i . In fact both σ and \mathbf{F} are vectors, then they are defined even by a versus. Conventionally the positive versus for surfaces is that toward the outside, in any case it is hard to find the outside of a not closed surface. The idea is that to consider surfaces σ and forces \mathbf{F} as opposite when the forces (pressures) are acting “on” the system has the following conceptual advantage: internal energy increases when a system is compressed (i.e. $\frac{dV}{dt} < 0$) and decreases when the system expands against the outside (i.e. $\frac{dV}{dt} > 0$).

3. Specific heats and adiabatic processes

The ideal gas law shown in Eq. 1.4 and in Eq. 1.5 tell us that the three variables p , V and T (or p , ρ and T) are linked together, then fixing two of them the third is automatically determined. In other words the degrees of freedom of the ideal gas are only two. This fact is important to obtain simpler and more useful forms of the first law of thermodynamics (Eq. 2.8). In particular, keeping into account Eq. 2.8 in the case of processes that took place at constant volume (isochoric transformation, $\frac{dV}{dt} = 0$), because we can write the internal energy U as a function of T only, that is $U = U(T)$, we have

$$\frac{dU}{dt} = \frac{\partial U}{\partial T} \frac{dT}{dt} = Q - p \frac{dV}{dt} \quad (3.1)$$

then

$$\frac{dU}{dt} = \frac{\partial U}{\partial T} \frac{dT}{dt} = Q \quad (3.2)$$

The quantity $\frac{\partial U}{\partial T}$ is the internal energy rate of change due to changes in temperature

and, when multiplied by the temperature rate of change with time gives the heat (supplied to or withdrawn from the system). For this reason it is usually called *specific heat at constant volume* and is represented by the symbol C_V (or c_V when referring to the mass unit). This quantity is not generally constant, it depends from the characteristics of the internal energy of the system, but it is usually assumed as constant. The origin of its name comes from the fact that the amount of heat absorbed by a system in the time from t and $t + \delta T$ is equal to

$$\int_t^{t+\Delta t} Q dt = \int_t^{t+\Delta t} C_V \frac{dT}{dt} dt = C_V \Delta T \quad (3.3)$$

Moreover, for an ideal gas, because $\frac{\partial U}{\partial eV} = 0$, in general it is possible to write the time variation of internal energy as

$$\frac{dU}{dt} = \frac{\partial U}{\partial T} \frac{dT}{dt} + \frac{\partial U}{\partial V} \frac{dV}{dt} = C_V \frac{dT}{dt} \quad (3.4)$$

then the first law of thermodynamics for an ideal gas becomes

$$C_V \frac{dT}{dt} = Q - p \frac{dV}{dt} \quad (3.5)$$

If we keep into account Eq. 2.8 for a process that take place at constant pressure (isobaric, e.g. $\frac{dp}{dt} = 0$), we will have

$$\frac{dU}{dt} - \frac{d}{dt}(pV) = \frac{d}{dt}(U + pV) = Q \quad (3.6)$$

This equation says that in a process that occurs at constant pressure, the amount of heat supplied (withdrawn) does not increase (decrease) in a proportional way only the internal energy, but the quantity

$$H = U + PV \quad (3.7)$$

which is usually called *enthalpy* and is referred using the symbol H . This quantity, in an ideal gas, depends only from temperature (remember the relationship between p , V and T) then we can use the chain rule for the differentiation and write

$$\frac{\partial}{\partial T}(U + pV) \frac{dT}{dt} = \frac{\partial H}{\partial T} \frac{dT}{dt} = Q \quad (3.8)$$

The quantity $\frac{\partial H}{\partial T} = C_p$ is called *specific heat at constant pressure* (c_p is the *specific*

heat at constant pressure for mass unit), it is usually assumed as constant with temperature and for an isobaric process gives the following rule to evaluate the amount of heat absorbed (released) in the interval Δt knowing only the initial and final temperatures.

$$\int_t^{t+\Delta t} Q dt = \int_t^{t+\Delta t} C_p \frac{dT}{dt} dt = C_p \Delta T \quad (3.9)$$

C_V and C_p are both related, and their relationship can be obtained with simple algebra by Eq. 3.7. In fact when we explicit the derivatives we derive enthalpy according to temperature, using Eq. 1.4 we have

$$C_p = \frac{\partial H}{\partial T} = \frac{\partial}{\partial T}(U + PV) = \frac{\partial U}{\partial T} + Nk \quad (3.10)$$

then

$$C_p = C_V + Nk = C_V + nR^* \quad (3.11)$$

or, using the specific heats for the mass unit, we have the even simpler form

$$c_p = c_v + R \quad (3.12)$$

where N is the number of molecules of the system and n the number of moles. This equation, because both n and R^* are positive, is telling us that $C_p > C_v$, then for a given increase (decrease) of temperature, the amount of heat absorbed (withdrawn) from a system is higher if the process took place at constant pressure than at constant volume. In other words, if you want to increase the temperature of a volume of air you have to keep constant its volume.

Eq. 3.5 gives us the opportunity to obtain an interesting relationship between two thermodynamic variables in the specific, but not rare, case of a system that does not exchange heat with the outside (these processes are called *adiabatic processes* and are characterized by $Q = 0$). For these processes Eq. 3.5 assumes the form

$$C_V \frac{dT}{dt} = 0 - p \frac{dV}{dt} \quad (3.13)$$

Then, using the ideal gas law (Eq. 1.4), the above equation becomes

$$C_V \frac{d}{dt} \left(\frac{PV}{nR^*} \right) = -p \frac{dV}{dt} \quad (3.14)$$

remembering that from Eq. 3.11 that $nR^* = C_p - C_v$, the above equation becomes

$$\frac{C_V}{C_p - C_V} \left(V \frac{dp}{dt} + p \frac{dV}{dt} \right) = -p \frac{dV}{dt} \quad (3.15)$$

defining $\gamma = \frac{C_p}{C_V} > 1$ we will have

$$\frac{1}{\gamma - 1} \left(V \frac{dp}{dt} \right) = -\frac{1}{\gamma - 1} \left(p \frac{dV}{dt} \right) - p \frac{dV}{dt} \quad (3.16)$$

then, after a few algebra

$$\frac{1}{p} \frac{dp}{dt} = -\gamma \frac{1}{V} \frac{dV}{dt} \quad (3.17)$$

This differential equation can be easily solved obtaining the expression

$$pV^\gamma = const. \quad (3.18)$$

which is called *Poisson expression for the adiabatic processes*, this equation is not in contrast with the ideal gas law, in fact it has been obtained thanks to it, it is just a special case of it obtained when the system does not exchange heat with the environment.

Using the ideal gas law we can derive other two expressions that define relationships between p , V and T , that is

$$p^{\frac{1-\gamma}{\gamma}} T = const. \quad (3.19)$$

and

$$TV^{\gamma-1} = const. \quad (3.20)$$

4. The adiabatic lapse rate

In section 1 the reasons why temperature is important have been shown, moreover in Figure 1 it is clearly shown that temperature generally (even if this cannot be stated using only a single case) decreases with height as well as pressure. From that figure one might expect that, since density is proportional to temperature (for constant pressure, as said by Eq. 1.5), the lower levels (near to the ground) are less dens than the upper ones, then

that atmosphere is unstable[†]. But you have to remember that pressure is not constant, in fact it decreases with height as shown in Figure 1. How can we discover if atmosphere is or is not stable? To face this question, with the adopted operational definition of *stability*, we should answer to the question if a parcel of air slightly displaced from its initial position returns back to it or moves away from it. This question can be faced adopting the simple conceptual model of adiabatic processes (1st assumption: $Q = 0$). In fact air is not a good thermal conductor, it is used as an insulator in houses and in other artificial and natural devices (e.g. feathers are good insulators because they keep air at rest near to the bird's body), for this reason we can consider that during the displacement the volume of air (parcel) does not exchange heat with the environment. Moreover, in our conceptual model of the ascent (descent), we can consider that the pressure of the parcel is nearly the same of the surrounding environment (2nd assumption: $p_p = p_s$). This means that the ascent is enough slow to assure the compensation and that the process takes place in a reversible way.

The 1st assumption makes us valuable the Eq. 3.19, if we differentiate it in respect to the height z we will have

$$\frac{d}{dz}(p^{\frac{1-\gamma}{\gamma}}) = 0 \quad (4.1)$$

expliciting this derivative and with some algebra we will have

$$\left(\frac{1-\gamma}{\gamma}\right)\frac{1}{p}\frac{dp}{dz} + \frac{1}{T}\frac{dT}{dz} = 0 \quad (4.2)$$

Using the 2nd assumption in conjunction with the meteorological form of the ideal gas law (Eq. 1.5) we have

$$\left(\frac{1-\gamma}{\gamma}\right)\frac{1}{\rho RT_s}\frac{dp}{dz} + \frac{1}{T}\frac{dT}{dz} = 0 \quad (4.3)$$

where T_s is the surrounding environmental temperature. At this point we have to explicit the term $\frac{dp}{dz}$. To do this we will use the vertical component of the Navier-Stokes equation for non rotating systems ($2\Omega \times \mathbf{v} = 0$) and for inviscid fluids ($\mathbf{F}_r = 0$), that is

[†] Stability is a complex concept, without going too much into the mathematical details, you can say that a system is stable if it returns to its previous state after small perturbations. Of course one should define "how small is small". This item is beyond the goals of these notes and will not be pursued.

$$\frac{dw}{dt} = -\frac{1}{\rho} \frac{\partial p}{\partial z} - g \quad (4.4)$$

If we make the third assumption that the fluid is in hydrostatic balance (3^{rd} assumption: $\frac{dw}{dt} = 0$) the Eq. 4.7 becomes

$$\left(\frac{1-\gamma}{\gamma}\right) \frac{-g}{RT_s} + \frac{1}{T} \frac{dT}{dz} = 0 \quad (4.5)$$

or, in a more ordered form,

$$\frac{dT}{dz} = -\left(\frac{\gamma-1}{\gamma}\right) \frac{g}{R} \frac{T}{T_s} \quad (4.6)$$

where T_s is the surrounding environment temperature and R is the gas constant. Remembering Eq. 3.12, which gives the relationship between the specific heats for the mass unit, we will have

$$\frac{dT}{dz} = \frac{g}{c_p} \frac{T}{T_s} = -\Gamma_a \frac{T}{T_s} = \quad (4.7)$$

The quantity $\Gamma_a = \frac{g}{c_p} \simeq 9.8 \frac{K}{km}$ is usually called *adiabatic lapse rate* and gives information on how much temperature decreases (increases) when lifted (shrunk) in an environment characterized by an environmental surrounding $T_s = T_s(z)$. Usually the quantity $\frac{T}{T_s}$ does not change so much from unity (remember that we are dealing with absolute temperatures), so it is assumed as $\frac{T}{T_s} \simeq 1$ then, for a reversible adiabatic ascent (descent) of an ideal gas in a hydrostatic environment we will have.

$$\frac{dT}{dz} \simeq -\Gamma_a \quad (4.8)$$

The adiabatic lapse rate can help us in deciding if an atmospheric layer is in a stable or unstable situation. The parcel we lift (shrink) will be in a stable situation if, when lifted (shrunk) will find itself in an environment with a lower (higher) density, otherwise it will buoy moving away from its position. In formulae, the parcel will be in a stable position if lifted when

$$\rho_p(z_0 + \Delta z) > \rho_s(z_0 + \Delta z) \quad (4.9)$$

that, using the ideal gas law (Eq. 1.5) and remembering that $p_s = p_p$, becomes

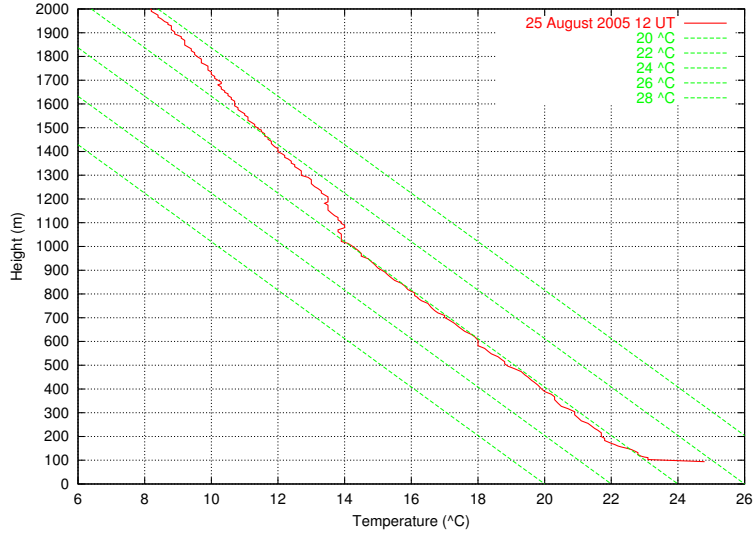


FIGURE 2. Example of the atmospheric vertical profile as observed at 12 UT of 25 August 2005 in Campofornido (UD). Solid line reports observations, dashed line represent the adiabatic lapse rates at 20, 22, 24, 26 and 28 °C.

$$\frac{1}{T_p(z_0 + \Delta z)} > \frac{1}{T_s(z_0 + \Delta z)} \quad (4.10)$$

Using the Taylor expansion for both $T_s(z)$ and $T_p(z)$, remembering that $T_s(z_0) = T_p(z_0)$

$$T_s(z_0) + \frac{dT_s}{dz} \Delta z + o(\Delta z)^2 > T_p(z_0) + \frac{dT_p}{dz} \Delta z + o(\Delta z)^2 \quad (4.11)$$

then the parcel is stable in its position z_0 when

$$\Gamma = \frac{dT_s}{dz} > \Gamma_a = \frac{dT_p}{dz} \quad (4.12)$$

The above equation gives a simple and powerful criterion to determine the dry stability of the atmospheric layers. Atmosphere is statically stable if its vertical thermal gradient Γ is higher than the adiabatic lapse rate Γ_a , it is unstable when $\Gamma_a > \Gamma$. When an atmospheric layer is unstable, then slight changes to its status will grow developing vertical motions. An example of this is shown in Fig. 2

In Fig. 2 it is clearly observable that the grate majority of the layers are in a statically stable situation apart from the lowest layer (near to the ground) and for some other thin layers (e.g. between 100 and 200 m). This is clear because statically unstable layers tend to evolve in stable states (vertical motions). It is hard to find unstable layers because

they are almost always in a transient situation, apart from the very lowest layer (why? Try to explain.). This is even the reason why the layer between ground (Campofornido is 94 m a.m.s.l.) and nearly 1000 m is characterized by an almost adiabatic lapse rate, then the reason why usually temperature decreases with height. This is not because the heating is from the ground (as it is written in several books) but because the energy transport generally took place by means of vertical motions (convection). In fact, as it is possible to notice in Fig. 1 and 2 temperature is not always decreasing with height because in those layers the transport of energy is not driven by vertical motions.

5. The potential temperature

The time is now right to define the *potential temperature*. This can be done easily using the equation 3.19. According to this equation pressure and temperature at a given level z_0 are related to the corresponding values at the level z by the relationship

$$p^{\frac{1-\gamma}{\gamma}} T = p_0^{\frac{1-\gamma}{\gamma}} T_0 \quad (5.1)$$

Remembering that $c_p - c_V = R$ we have

$$T \left(\frac{p}{p_0} \right)^{-\frac{R}{c_p}} = T \left(\frac{p_0}{p} \right)^{\frac{R}{c_p}} = T_0 \quad (5.2)$$

When referring to a fixed p_0 the quantity $\theta = T_0$ is called *potential temperature* and its physical meaning is that of the temperature a parcel in equilibrium with its environment at a pressure p and that will have if lifted or shrunk to the pressure p_0 with an adiabatic process. In this sense potential temperature is not a real temperature that you can measure with a thermometer, but it is useful, as we will see. Usually the referring value for pressure is $p_0 = 1000hPa$, then Eq. 5.3 becomes

$$T \left(\frac{1000}{p} \right)^{\frac{R}{c_p}} = \theta \quad (5.3)$$

For a standard atmosphere we can assume that $\frac{R}{c_p} = 0.286$. The great advantage of potential temperature is that, by construction, it is a conservative quantity for an adiabatic ascent or descent of the particle. In other words if you compute the quantity θ for a particle, it does not change during the ascent or descent of the particle. This advantage is clear if, for the same sounding of figure 2, we compute the potential temperature. In

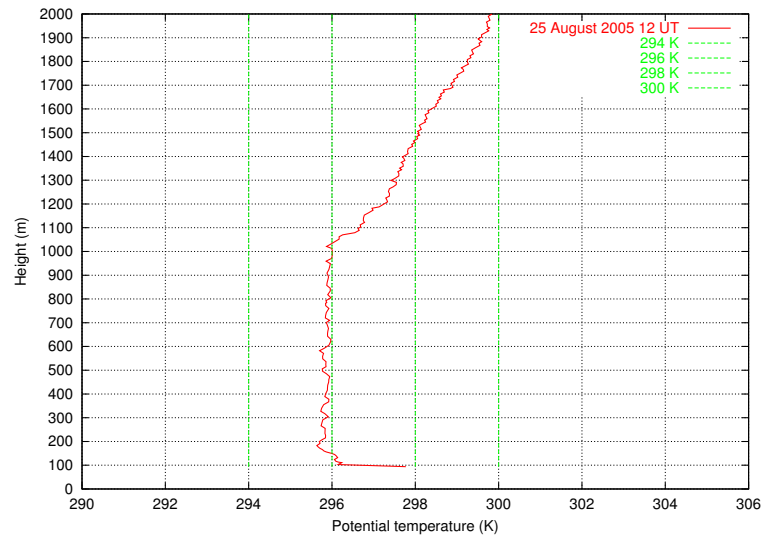


FIGURE 3. Example of the potential temperature obtained through the atmospheric vertical profile observed at 12 UT of 25 August 2005 in Campoformido (UD). Solid line reports observations of potential temperature, dashed line represent the constant potential temperatures at 294, 296, 298, and 300 K.

figure 3 it is much more easy to recognize the statically instable layers from the stable ones.