2. The Equation of State

Related material in Tsonis: Chaps. 2 and 3

At the onset, take careful note of the notation used. In some cases, a variable, such as M, may have a dual definition. Please note any duplication in this chapter, including the Appendix.

2.1 The equation of state and elementary kinetic theory

The equation of state has been derived from both empirical and theoretical approaches. All gases are observed to obey – not exactly, but to a very close approximation – the equation of state (or ideal gas law) over a wide range of conditions. We consider first a derivation of the equation of state from elementary kinetic theory, which is based on classical mechanics. This theory begins with the assumption that gases are "ideal." An ideal gas exhibits the following microscopic characteristics:

- 1. It consists of molecules.
- 2. The molecules are in random motion and obey Newton's laws of motion.
- 3. The total number of molecules is large.
- 4. The volume of molecules is negligible relative to the volume occupied by the gas.
- 5. No appreciable (molecular) forces act on the molecule during a collision.
- 6. Collisions among molecules are elastic and of negligible duration.

In addition, the following assumptions are made regarding the interaction between molecules and a surface membrane or wall which contains the molecules.

- 7. The collision with the wall is elastic.
- 8. There is no loss in momentum, in the direction parallel to the wall, during the collision with the wall (i.e., no friction).





(The material below is extracted from Fleagle and Businger 1980 and Bohren and Albrecht 1998.)

Now we consider the collision of a molecule with the wall according to the geometry shown in Fig. 2.1. The change in momentum (mv) occurs from only that component of motion perpendicular to the wall, and can be expressed as

$$mv\cos\theta - (-mv\cos\theta) = 2mv\cos\theta. \tag{2.1}$$

If we consider collisions from all molecules impacting the wall from all angles, one can show from elementary kinetic theory (e.g., Fleagle and Businger 1980, p. 31) that the incremental change in momentum is

$$mv^2 dn_v sin\theta cos^2 \theta d\theta$$
, (2.2)

where dn_v represents the number of molecules with speeds between v and v+dv per unit volume. Integration between 0 and $\pi/2$ accounts for the collisions from all directions. The integration result is:

$$(mv^2/3)dn_v.$$
 (2.3)

The differential force exerted on the incremental surface dA is found by integrating over all the range of velocities of all molecules:

$$dF = (m/3)(\int v^2 dn_v) dA.$$
 (integrating from 0 to n) (2.4)

Since the definition of pressure is p = F/A, the pressure on the surface is then

$$\mathbf{p} \equiv \mathbf{d}F/\mathbf{d}A = (\mathbf{m}/3)\int \mathbf{v}^2 d\mathbf{n}_{\mathbf{v}},\tag{2.5a}$$

which may be expressed as (assuming a total of n molecules)

$$p = mn\langle v \rangle^2 / 3 \tag{2.5b}$$

where $\langle v \rangle = (1/n) \int v^2 dn_v$ is the average speed of a molecule.

Eq. (2.5b) can be rewritten, using the identity $dN \equiv ndV$ (which assumes that the molecules are distributed uniformly over a differential volume dV), as

$$pV = (1/3)Nm\langle v \rangle^2 \tag{2.6}$$

If we consider a particular volume, such as the molar specific volume α_m (the volume of one kilomole of gas; units m³ kmol⁻¹), which contains N₀ molecules (Avogadro's number), Eq. (2.6) can be expressed as

$$p\alpha_{\rm m} = (1/3)N_0 \langle v \rangle^2. \tag{2.7}$$

At this point we define temperature as being proportional to the average kinetic energy of the molecules according to

$$(3/2)kT = (1/2) \langle mv \rangle^2$$
 (2.8)

where 3/2 k is a constant of proportionality and k is the Boltzmann constant (k = 1.38×10^{-23} J K⁻¹). (This relation is the definition of temperature on a kinetic theory or microscopic basis, and is necessary to link kinetic theory to the equation of state for an ideal gas.) Substitution of Eq. (2.8) into Eq. (2.7) produces the equation of state for an ideal gas:

$$p\alpha_m = kN_0T = R^*T, \tag{2.9}$$

where $R^* = kN_0$ is the universal gas constant. The value of R^* is 8.31441 ± 0.00026 J mol⁻¹ K⁻¹, and hence the value of k = 1.380662 ± 0.000044 x 10⁻²³ J K⁻¹. We note that k can be interpreted as the gas constant for a single molecule. We also note that, for a single gas (or mixture of gases), k, N₀ and hence R* are universal.

The Appendix contains a slightly different derivation of the gas law, using kinetic theory. (Also see Tsonis, pp. 7-10.)

The equation of state can also expressed in terms of the specific volume $\alpha = \rho^{-1}$ according to the definition

$$\alpha = \alpha_{\rm m} / {\rm M},$$

where M is the molecular weight of the gas. Then the equation of state then assumes the more familiar form

$$p\alpha = (R^*/M)T = RT, \qquad (2.10)$$

where R is the specific gas constant for a gas with molecular mass M.

An aside [extracted from Fleagle and Businger, p. 34]

The numerical value of R can be determined experimentally. We find that a specific mass of gas brought into temperature equilibrium (after some extended time) with an ice surface (via conduction) always attains the same volume. A similar behavior is observed for gases reaching equilibrium with boiling water (again, the boiling point has a specific temperature). In both cases, the average kinetic energy of the gas molecules approaches that of the contacting molecules, in which case the systems in contact have uniform temperature and are in equilibrium. It is also observed that when two different masses of gas (or systems) are brought into equilibrium with melting ice, they are also in equilibrium with each other. This is a general thermodynamic principle called the *Zeroth Law of Thermodynamics*.

2.2 The classical equation of state (ideal gas law) – the empirical approach

Laboratory experiments have shown that pressure, volume and temperature of a gas are related by an equation of state. In the laboratory setting, experiments conducted by early physicists/chemists examined the relationship among the 3 variables in the gas law by holding one constant. For a constant temperature (isothermal) process, Boyle (1600's) discovered that the volume V of a gas is inversely proportional to pressure p.

1. Boyles Law: $V \propto p^{-1}$ for an <u>isothermal process</u>.

Later experiments by Gay-Lussac (and Charles?) (1700's) led to two additional "Laws". Bohren and Albrecht (1998) state that the empirical work on the ideal gas law should have been attributed to Gay-Lussac, rather than Charles, who never published work in this area. (In fact, Tsonis makes reference to the First (T,V) and Second (T,p) Laws of Gay-Lussac.)

http://www.grc.nasa.gov/WWW/K-12/airplane/aboyle.html (from google on "Boyle's Law")

2. For a fixed mass of gas at constant pressure (isobaric) the volume of a gas is directly proportion to absolute temperature.

First Law of Gay-Lussac: <mark>V ∝ T for an <u>isobaric</u> process</mark>:

[Charles First Law]

 $dV = \alpha V_0 dT$

(V₀ is the volume at T = 0 °C, and α is the volume coefficient of expansion at constant pressure, with a value of 1/273 deg⁻¹)

The above equation can be integrated to give the relation between V and T:

 $V = V_0(1 + \alpha T)$

A graph is shown in Fig. 3.1 of Tsonis.

http://en.wikipedia.org/wiki/Gay-Lussac's_law

or

http://www.chem.csus.edu/gaslaws/gay-lus.html

3. For a fixed mass of gas at fixed volume (isochoric), the pressure is proportional to the absolute temperature.

Second Law of Gas-Lussac<mark>: p ∝ T for an <u>isochoric</u> process</mark> [Charles <u>Second</u> Law]

 $dp = \beta p_0 dT$

(p_0 is the pressure corresponding to a temperature T of 0 °C. β is the pressure coefficient of thermal expansion at constant volume and has the value 1/273 deg⁻¹.)

Again, we can integrate the above differential equation to get

 $p = p_0(1{+}\beta T)$

A graph of this function (Fig. 3.2, Tsonis) is similar to that of Fig. 3.1. Note that an extrapolation of both graphs to V=0 or p=0 defines absolute zero, -273 °C (but for an *ideal* gas).

Discuss the Application paragraph in Tsonis, page 12, bottom:

Houses are not air tight, and in fact are ventilated with exhaust fans in bathrooms, and with clothes dryer vents.

The above simple laws can be combined to infer the equation of state. Tsonis discusses this in a simple fashion on p. 15 (Section 3.7) of his book. From his equation (3.7), we have

$$pV/T = p'V'/T' = constant = A$$

 $pV = AT$,

where A is a constant, which can be equated to nR^* as follows.

$$pV = nR^*T = m(R^*/M)T = mRT$$
 (2.11)

where n = number of moles (or molar abundance), M = molecular (molar) weight, R* is the universal gas constant and R = R*/M is the specific gas constant (for a given gas, or mixture of gases). This equation can also be expressed in terms of density ($\rho = m/V$), yielding the intensive form

$$p = \rho RT. \tag{2.12a}$$

Using the definition of specific volume ($\alpha = \rho^{-1}$) one gets an expression identical to (2.10),

$$\mathbf{p\alpha} = \mathbf{RT}.$$
 (2.12b)

This latter expression is one that we'll use most. This equation provides a good approximation for the dry atmosphere.

For a mixture of gases, we can utilize the same equation of state by invoking *Dalton's Law of partial pressures*, which states

$$p = \Sigma p_i(T, V)$$
 (p_i is the partial pressure of gas i), (2.13)

or by Amagat's Law of partial volumes

$$V = \Sigma V_{i} (T,p) \tag{2.14}$$

Both relations, 2.13 and 2.14, are exact for ideal gases, but are only approximate for real gases. It can be shown that the equation of state holds for mixtures of gases as well as for the individual component gases, i.e., $p_i\alpha_i = R_iT$. Thus, the value of R for the <u>dry</u> atmosphere in Eqs. (2.12a,b) is $R_d = 287.05 \text{ J kg}^{-1} \text{ K}^{-1}$ (Table 1.1). [We also note that $M_d = 28.96 \text{ g mol}^{-1}$.]

Wikipedia discussion of the ideal gas law: <u>http://en.wikipedia.org/wiki/Ideal_gas_law</u> Hyperphysics, Georgia State Univ.: <u>http://hyperphysics.phy-astr.gsu.edu/hbase/kinetic/idegas.html</u>

Next, we will consider an equation of state for moist air (i.e., air having water vapor also). Before we do so, let us define some moisture variables that are commonly used in atmospheric thermodynamics:

- *water vapor pressure (e)*: the partial pressure due to water vapor molecules (How could this be measured? We will see that it can be determined theoretically/analytically.)
- *mixing ratio*: $r_V = m_V/m_d$
- specific humidity: $q_v = (m_v/(m_v + m_d)) = r_v/(1 + r_v)$ (2.15)

The equivalent form of Eq. (2.12) is thus

$$p = \rho_m T[(m_d R_d + m_v R_v)/(m_d + m_v)] \qquad (mass weighted) \qquad (2.16)$$

where m_d and m_v are the masses of dry air and water vapor, $R_v = R^*/M_{H_20} = 461.5 \text{ J kg}^{-1} \text{ K}^{-1}$, moist air density (ρ_m) is

$$\rho_{\rm m} = ({\rm m}_{\rm d} + {\rm m}_{\rm v}) / {\rm V} = \rho_{\rm d} + \rho_{\rm v} \tag{2.17}$$

and V is the volume of the gas. Eq. (2.16) can be rewritten as

$$p = \rho_m R_d T \left[\frac{m_d + m_v \frac{R_v}{R_d}}{m_d + m_v} \right]$$

Dividing each term in the brackets of the above relation by m_d yields

$$p = \rho_m R_d T \left[\frac{1 + \frac{m_v}{m_d} \frac{R_v}{R_d}}{1 + \frac{m_v}{m_d}} \right]$$
(2.18)

where the quantity $R_v/R_d = \epsilon = 1.61$.

Using the definition of mixing ratio ($r_v = m_v/m_d$), Eq. (2.18) can be rewritten as

$$p = \rho_m R_d T \frac{1 + 1.61 r_v}{1 + r_v}$$
(2.19)

or, by using the definition of specific humidity as

$$p = \rho_m R_d T (1+0.61 q_v)$$
(2.20)

Since r_v only very rarely exceeds 22 g kg⁻¹ in the atmosphere (this can be shown with relationships that will be developed and examined in subsequent chapters) and $q_v \ll 1$ (0.025 at most), then $q_v \sim r_v$ and Eq. (2.20) can be expressed (approximately - this is not exact) as

$$p = \rho_m R_d T (1 + 0.61 r_v)$$
(2.21)

If we define virtual temperature as

$$T_V = (1+0.61q_V)T \cong (1+0.61r_V)T$$
 (2.22)

then (2.20) becomes

$$p = \rho_m R_d T_v. \tag{2.23}$$

Example: Find the density of the air outside at this time.

2.3 An empirical equation of state with corrections for the non ideal nature of gases

Van der Waals' Equation (1873) is an empirical equation of state accounting for the non-ideal nature of gases:

$$(p + aV^{-2})(V - b) = R^*T,$$
 (1 mol) (2.24)

where a and b are specific constants for each gas. Such empirical equations were formulated to account for the non-ideal behavior of gases. [More can be added here on how this equation is derived.] Also, refer to http://chemed.chem.purdue.edu/genchem/topicreview/bp/ch4/deviation5.html

Another empirical equation of state that exhibits a better fit to observed data was formulated by Kammerlingh-Onnes and is given as

$$pV = A + Bp + Cp2 + Dp3 + ...$$

= A(1 + B'p + C'p² +), (1 mol) (2.25)

where A, B', C', ... are called the virial coefficients and are functions of temperature. The first virial coefficient is A=R*T for all gases, as it is the value of pV for p->0. The table below gives the second virial coefficient for dry air. Higher order virial coefficients may be neglected for normal atmospheric conditions (<0.2% error).

T (°C)	$B'(10^{-8} m^2 N^{-1})$	pV/R*T	
	- ()	p = 500 hPa	p = 1000 hPa
-100	-4.0	0.9980	0.9996
-50	-1.56	0.9992	0.9984
0	-0.59	0.9997	0.9994
50	-0.13	0.9999	0.9999

Table 2.1. Second virial coefficient B' for dry air. Taken from Iribarne and Godson (1973)

2.4 A linearized equation of state

Note that (2.21) is a nonlinear equation, i.e., it involves products of at least two atmospheric variables. By linearizing equations about a dry reference state, it is possible to simplify cloud/mesoscale models and provide an easier tool to (sometimes) diagnose atmospheric thermodynamic properties. The dry reference state is defined as a dry horizontally-uniform atmosphere that varies in height (z) only. Reference state variables are denoted with the subscript 0. Departures from these variables are then defined, such as $p'(z) = p(z) - p_0$. By definition, the reference state obeys the gas law $p_0\alpha_0 = R_dT_0$. For our equation of state in Eq. (2.21) we implement the following decompositions of the thermodynamic variables:

 $\alpha = \alpha_0 + \alpha'$, $p = p' + p_0$, $T = T' + T_0$, and $r_v = r_v'$ (since the reference state is dry),

Eq. (2.23) becomes

$$p_{0}(1+p'/p_{0})\alpha_{0}(1+\alpha'/\alpha_{0}) = R_{d}(1+0.61r_{v}')T_{0}(1+T'/T_{0})$$

Now we take the natural logarithm of both sides:

$$\ln(1+p'/p_{O}) + \ln(1+\alpha'/\alpha_{O}) = \ln(1+0.61r_{V}') + \ln(1+T'/T_{O})$$

Expanding in a Taylor's series $[\ln(1+x) = x - x^2/2 + x^3/3 + ... \text{ for } |x|<1]$ we can obtain

$$p'/p_{0} + \alpha'/\alpha_{0} = T'/T_{0} + 0.61 r_{v}'$$
(2.26)

or

$$\alpha'/\alpha_0 = T'/T_0 + 0.61 r_V' - p'/p_0.$$
(2.27)

Eq. (2.27) represents a linearized equation of state for a moist system. Thus, fluctuations in specific volume are produced by the (linear) sum of fluctuations in temperature, moisture and pressure with respect to their basic states.

Question: Explain why Eq. (2.27) is an approximate form of the equation of state. (Note that we certainly cannot refer to it as an ideal gas law.)

Example:

Typical perturbations within a cloud are:

 $\begin{array}{l} T' \ \sim 1 \ K \ (up \ to \ 15 \ K) \\ r_V' \sim 2 \ g \ kg^{-1} \ (up \ to \ 8 \ g \ kg^{-1}) \\ p' \ \sim 0.2 \ mb \ (up \ to \ 1-2 \ mb) \end{array}$

Thus, $T'/T_0 = 1/273 = 0.0037$, $r_v' = 0.002$, and $p'/p_0 = 0.2/800 = 0.00025$.

Temperature and moisture perturbations are comparable and thus provide the most important contributions to density fluctuations in the cloud (or cloud-free) environment. Only in limited regions of cloud systems does p' exceed 0.2-0.4 mb. [It is the density fluctuations that control cloud dynamical processess.]

2.5 *Measurements of temperature, pressure, and water vapor*

Temperature: thermometer, thermister, thermocouples, IR emission, microwave emission (O_2 band) Density: lidar

Pressure: barometer (mercury, aneroid), transducer

Water vapor: wet bulb temperature, RH directly, lidar differential absorption, microwave emission Virtual temperature: radio acoustic sounding system (RASS – speed of sound $\propto T_v$)



Review (from http://www.grc.nasa.gov/WWW/K-12/airplane/eqstat.htm)

NAM	Equation of State (Ideal Gas)	Glenn Research Center
	Properties	
Density = r F	Pressure = p Temperature = T Volume = V Observations	Mass = M
Boyle:	For a given mass, at constant temperature, pressure times the volume is a constant.	^{the} p V= C ₁
Charles and Gay–Lussac:	For a given mass, at constant pressure, the is directly proportional to the temperature	volume e. V = C₂T
Combine	$\mathbf{P} \cdot \mathbf{V} / \mathbf{T} = \mathbf{n} \cdot \mathbf{R} - \mathbf{R} = 8.31 \text{J} / \text{mole} / \text{K}^{2}$	(Universal)
P	p V = n R T $n = number of modelsDivide by mass: Specific Volume = v = \frac{v}{M}p v = \underline{nRT} or p v = R T or p = \frac{1}{M}$	oles <u>olume</u> = <u>1</u> mass = r R r T
v	R = Constant value for e .286 kJ/kg/K ⁰ (fe	ach gas or air)

Refer to the following web site which has the 3-D equation of state diagram shown below. This provides a very good review of the equation of state.

http://hyperphysics.phy-astr.gsu.edu/hbase/kinetic/idegas.html

Example (Wallace and Hobbs, 2006)

If at 0 °C the density of dry air alone is 1.275 kg m⁻³ and the density of water vapor alone is 4.770 x 10^3 kg m⁻³, what is the total pressure exerted by a mixture of the dry air and water vapor at 0 °C?

Solution: From Dalton's law of partial pressures, the total pressure exerted by the mixture of dry air and water vapor is equal to the sum of their partial pressures. The partial pressure exerted by the dry air is

 $p_d = \rho_d R_d T$

where ρ_d is the density of the dry air (1.275 kg m⁻³ at273 K), R_d is the gas constant for 1 kg of dry air (287.0 J K⁻¹ kg⁻¹), and *T* is 273.2 K. Therefore,

 $p_d = 9.997 \text{ x } 10^4 \text{ Pa} = 999.7 \text{ hPa}$

Similarly, the partial pressure exerted by the water vapor is

 $e = \rho_v R_v T$

where ρ_v is the density of the water vapor (4.770 x 10³ kg m³ at 273 K), R_v is the gas constant for 1 kg of water vapor (461.5 J K⁻¹ kg⁻¹), and *T* is 273.2 K.

Therefore,

e = 601.4 Pa = 6.014 hPa

Hence, the total pressure exerted by the mixture of dry air and water vapor is

 $p = p_d + e = 999.7 + 6.014 hPa = 1005.7 hPa.$

Homework

Review the examples on pp. 19-22 of Tsonis

Problems (due in one week)

- 1. (a) Determine the number of molecules in a 1 cm³ volume of air having a pressure of 1 atm. Make any other reasonable assumption if required. [Ans: about $3x10^{19}$ cm⁻³ your answer will be more precise]. (Note, this is similar to problem 3.5 in Tsonis.) (b) What is the mean free path for the average molecule in this volume? Mean free path is determined from $\Delta x_{mfp} = (n\sigma)^{-1}$, where n is the number of molecules per unit volume, $\sigma = \pi d_o^2$ is the collision cross section (σ is about 3 x 10⁻¹⁵ cm² for an air molecule), and d_o is the diameter of an average molecule. You can check your answer with Fig. 1.1b.
- 2. At what pressure is the ideal gas law in error by 1%, for air with T = 0 °C? [Ans: 17 atm; Hint: Use Table 2.1]
- 3. (a) Calculate some extremes in air density at the surface for different scenarios. For example, consider
 (a) International Falls in the winter under high pressure (anticyclone) conditions: T = -40 °F, p=1050 mb, r_v=0.1 g kg⁻¹; (b) Denver in the summer with T = 95 °F, p=850 mb (actual station pressure) and r_v=10 g kg⁻¹. (c) What are some practical implications (e.g., aircraft lift, wind drag on a vehicle)?
- 4. [Fleagle and Businger Prob. 1, ch. 2.] If 10^6 molecules are required in order to ensure a statistically uniform distribution of velocities in all directions, what is the minimum volume in which the state can be defined at standard atmospheric conditions (p=1013 mb, T=0 °C)? [Ans. 37.21x10⁻²¹ m³, which corresponds to a linear distance of $3.34x10^{-7}$ m for a cube. Hint: use the definition dN = ndV].
- 5. Show analytically (using appropriate equations, not just simple calculations) that the density of moist air is less than that for dry air at the same temperature and pressure. Interpret your result. Does this difference have any relevant atmospheric applications? (Hint: Refer to Petty)

Complete the following problems from Tsonis: 3.1 3.2 (Do this via IDL, PV-WAVE, or Excel)

3.7

ATS/ES 441 students: You may eliminate two problems of your choice, or if you turn in all problems, I will ignore the lowest scores on two problems.

Appendix: Derivation of the gas law from kinetic theory using a slightly different approach

(See also Tsonis, pp. 7-10)

We consider a volume of N identical molecules within a cubic container having a volume V. We begin with Newton's Second Law: $\mathbf{F} = d(m\mathbf{v})/dt$. (mv is momentum, and v is the velocity vector.) The momentum change that a molecule of mass m experiences after collision with a flat surface is $2mv_x = \int F_x dt$

During a finite time interval Δt , the number of molecules that strike an area A of the surface is $nv_xA\Delta t/2$, where n = N/V is the number density of identical molecules. The time-integrated force on the area A is the following (where F_x is the instantaneous force on the area A):

$$\frac{1}{2}nv_{x} 2mv_{x} A\Delta t = \int_{0}^{t} F_{x} dt$$

The average force over a time interval $\Delta t = \int dt$ is

$$\left\langle F_{x}\right\rangle = \frac{\int_{0}^{t} F_{x} dt}{\Delta t}$$

Since pressure is defined as force per unit area (p = F/A), we can use the previous two equations to write $p = nm/y^{2}$

$$p = nm \langle v_x^2 \rangle$$

We have assumed that all molecules have the same speed. This is not true, but we can use an average speed $\langle v_x \rangle$ in place of v_x , which is shown in the preceding equation. We also need to be more realistic by including all three motion components. On average, the kinetic energy of all components is identical, i.e., $\langle v_x^2 \rangle = \langle v_y^2 \rangle = \langle v_z^2 \rangle$

Then we make the assumption that the total speed v is related to its components v_x , v_y , and v_z according to

$$\left\langle \mathbf{v}^{2}\right\rangle = \left\langle \mathbf{v}_{x}^{2} + \mathbf{v}_{y}^{2} + \mathbf{v}_{z}^{2}\right\rangle = \left\langle \mathbf{v}_{x}^{2}\right\rangle + \left\langle \mathbf{v}_{y}^{2}\right\rangle + \left\langle \mathbf{v}_{z}^{2}\right\rangle$$

Question for ATS 541 students: Under what conditions is this true?

Then the momentum of an ensemble of N molecules within a volume V can be expressed

$$p = \frac{1}{3} \frac{N}{V} \left\langle mv^2 \right\rangle$$

Now we make a statement, based on our wisdom, that relates the momentum to absolute temperature:

$$\frac{1}{3}\langle mv^2 \rangle = kT$$
 or $\left\langle \frac{1}{2}mv^2 \right\rangle = \frac{3}{2}kT$

where k is Boltzmann's constant (k = $1.38 \times 10^{-23} \text{ J K}^{-1}$). (Note: T \propto mean molecular kinetic energy)

The most basic form gas law can then be written as

$$pV = NkT$$

We can rewrite the above form as

$$p = m \frac{N}{V} \frac{k}{m} T = \rho R T$$

Since $\rho = Nm/V$, and the gas constant R is defined as R = k/m.

True or False?

- 1. Atmospheric pressure is just the weight of the atmosphere above us.
- 2. Absolute zero is the temperature at which all motion ceases.
- 3. As temperature increases, so does pressure, and vice versa.
- 4. Cold air is denser that hot air.

Relation between T and p?

