Classroom

In this section of *Resonance*, we invite readers to pose questions likely to be raised in a classroom situation. We may suggest strategies for dealing with them, or invite responses, or both. "Classroom" is equally a forum for raising broader issues and sharing personal experiences and viewpoints on matters related to teaching and learning science.

Condensation of Water Vapour in Ambient Atmosphere Under Applied Pressure at Room Temperature
A Demonstration Experiment

Amit R Morarka[1,2]
[1] Abasaheb Garware College
Karve Road, Pune 411 004, India.
[2] Department of Electronics
Science, Savitribai Phule Pune
University, Pune 411 007, India.
Email: amitmorarka@gmail.com

The article demonstrates the condensation of water vapour under pressure and at room temperature, using a disposable syringe and an ambient light source. The small amount of trapped water vapour inside a syringe along with the air is compressed manually to one-sixth of its original volume. The compressed water vapour inside the syringe can be easily seen in the form of a white condensate on the inner side of the transparent, circular wall of the syringe. The condensate cloud can be observed in the atmosphere by releasing it in the open air while holding the syringe in front of a light source over a dark background. This model can be used as an analogous experiment while explaining the concept of condensation of gases under pressure – from school to postgraduate students.

1. Introduction

Liquefied petroleum gas (LPG), dry ice (solid carbon dioxide), air filled balloon, tyer tubes, etc., are few examples of compressed gases that we come across in day to day life. Other compressed gases like liquid oxygen (O_2), nitrogen (N_2), helium (He), etc.,

Keywords
Condensation, pressure, gases, water vapour, temperature, vapour pressure, kinetic gas theory, state parameters, ideal gas.

are seen only in research laboratories and hospitals. These gases are compressed under the application of high pressure inside a chamber which liquefies them. To understand any concept about the compression/expansion of gases, we revert back to the basic equation of state [1] for an ideal gas which is given as,

$$PV = n\mathrm{R}T\,, \tag{1}$$

where,
P = Pressure of the gas
V = Volume occupied by the gas
n = Number of moles
R = Gas constant
T = Temperature of the gas

The parameters in (1) are called the 'state parameters'. This implies that one can get information about the state (equilibrium state and various properties) of the gas using all of these three parameters. In (1), we assume that for a given ideal gas, 'R' is the universal gas constant. The number of molecules for a given mass of a gas is also a constant. This implies that the number of moles 'n' for that gas is a constant. If the compression/expansion of the gas is taking place at a constant temperature (isothermal process), then (1) can be modified in the following way:

Case 1: For an uncompressed gas, we have,

$$P_1V_1 = n\mathrm{R}T\,, \tag{2}$$

where,
P_1 = Pressure of the gas before compression
V_1 = Volume of the gas before compression

Case 2: For a compressed gas, we have,

$$P_2V_2 = n\mathrm{R}T\,, \tag{3}$$

where,

P_2 = Pressure of the gas after compression
V_2 = Volume of the gas after compression

Combining (2) and (3), we get,

$$P_1V_1 = P_2V_2 \,. \tag{4}$$

Equation (4) provides us with three known (P_1, V_1, and P_2) and one unknown parameter (V_2). For a given gas, the value of V_2 can be estimated to a good degree of accuracy using (4).

Of the many properties of a gas, one which attracts much attention is its compressibility. Roughly speaking, compressibility [1] is a measure of the decrease in a unit of volume (ratio of change in volume to the original volume = $\Delta V/V_O$) for a unit increase in pressure (ΔP). In principle, all the three phases of matter viz; solid, liquid and gas can be compressed. Practically, it is observed that solids and liquids undergo compression under extreme pressures. On the other hand, gases can be compressed at relatively lower pressures. The property of compressibility of gases has been utilized in many industrial and household applications.

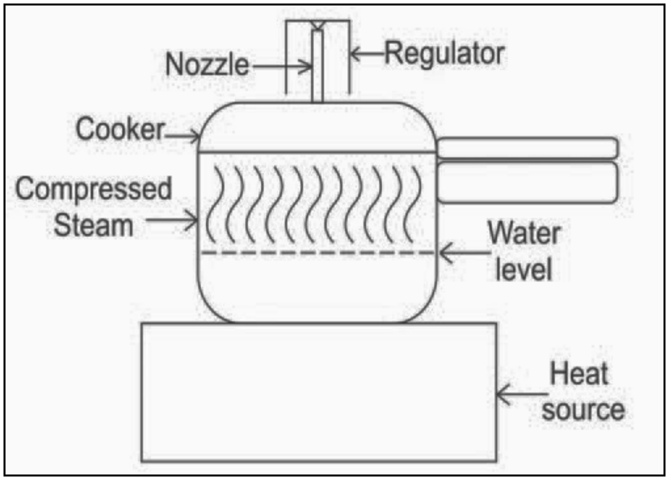
Of the many properties of a gas, one which attracts much attention is its compressibility.

Examples of such uses are but not limited to include the compressors in air conditioners and refrigerators, pressure cookers, tyre tube of bicycles and inflated balloons which have been used by most of us in our day to day life.

To further explain the compressibility of gases, we will concentrate on two examples mentioned above – pressure cooker and inflated balloon.

A familiar sound to our ears is the ‘whistling’ of the pressure cooker. As the water inside the cooker boils at 100$^{\text{o}}$C it becomes steam. This steam is inhibited from leaving the cooker through the vent pipe by a small, lightweight, freely suspended metallic plug called the ‘regulator’ as shown in *Figure* 1. As the heating continues, steam is also continuously generated inside the cooker. Due to the increase in the quantity of steam inside the constant volume of the cooker, the pressure also increases. The pressure of the steam inside the cooker keeps on increasing to a point where

Figure 1. Schematic of a household pressure cooker on a heating source.



the steam pushes the regulator against the gravity and the vent pipe opens. At the same instant, the accumulated steam from inside the cooker escapes through the vent pipe into the atmosphere while making the whistling sound.

Steam occupies approximately 1600 times the volume as that occupied by the same mass of water when it was in the liquid state.

The loss of steam reduces the pressure inside the cooker. The reduced pressure has insufficient force to lift the regulator against the gravitational force due to which the outflow of steam ceases. Since the heating is continued, the cycle repeats itself till water is present inside the cooker. But wait a minute, amid all of this functioning of cooker explained, where is the compression of gas happening? The answer to this question lies in the statement 'pressure of the steam' which was made in the above explanation. We know that matter in the gaseous state occupies the entire volume available to it. As the water (having some mass) is converted into steam, the steam expands in volume approximately 1600 times as compared to the volume occupied by the same mass of water when it was in the liquid state. This is given as,

$$V_{\mathrm{s}} = 1600 \times (V_{\mathrm{w}}), \tag{5}$$

where, V_{s} is the volume of steam and V_{w} is the volume of water.

Therefore, if 1 ml of water (initial volume V_{w}) is converted en-

Box 1. Hooke's Law

Hooke's law of elasticity states that as long as a material regains its original dimensions when the applied forces acting on the material are removed, the material is said to be elastic in nature. Materials which get permanently deformed after the removal of the deforming force are said to be plastic in nature. In short, in an elastic material applied stress is directly proportional to the strain produced in the material.

tirely into steam at 100^{o}C, it will occupy 1600 ml of volume. If this steam is collected in a container having volume less than 1600 ml, the steam will be said to be under pressure [1]. Steam under pressure is utilized for cooking as well as in locomotive engines.

Similarly, when a balloon is inflated, air from the mouth is pushed into the balloon. As we push more and more air inside it, the balloon keeps on expanding against two pressures. One exerted by the atmosphere and the other by the elasticity of the material of the balloon. As the balloon is inflated continuously against the atmospheric pressure, a point is reached where the elastic limit of the balloon material is crossed and the material permanently deforms (Hooke's law of elasticity, *Box* 1), that is, it bursts! In this example, the air inside the balloon is said to be under pressure till the point where the balloon ruptures. If the balloon is inflated below its rupture limit and released without sealing its inlet, we observe the balloon propelling in random directions due to the expulsion of compressed air from the balloon as shown in *Figure* 2. The balloon will keep propelling in random directions till the air pressure inside the balloon equalizes with the atmospheric pressure. The application of this concept is used in propelling rockets and satellites into the Earth's orbit and in deep space for cosmic research.

Both of the above examples are based on the compression of gases (steam and air). The compressed air coming out of the balloon cannot be seen by the naked eye. On the other hand, steam coming out of the cooker, which is under pressure, is no doubt a di-

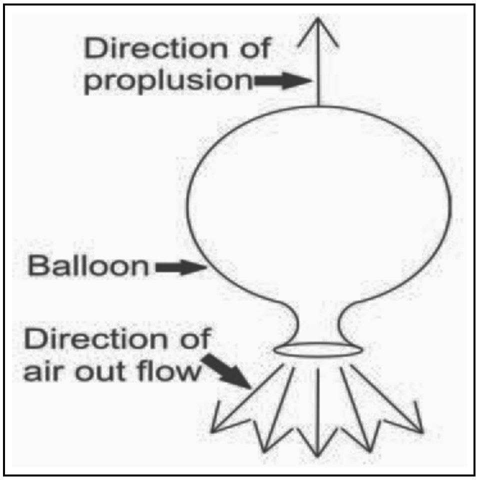


Figure 2. Schematic of a propelling balloon.

rect example of compression of gases. But to observe this practically, we need a pressure cooker type system and a heating source to boil the water. In many schools and undergraduate colleges, such equipment may not be readily available. Above all, demonstrating the experiment of compression of gases in the classroom while the concept is discussed, using bulky apparatus involving high temperatures or high pressures is difficult. Moreover, during such demonstrations, students do not get an experiential learning of the physical concept individually which is required to trigger their interest in physical sciences. Addressing these challenges, a simple technique was developed using readily available materials like a disposable syringe to demonstrate compression of gases (water vapour) at room temperature.

2. Experiment

Humidity (water vapour) present in the atmosphere was used as the working gas to be compressed. Ambient temperature at the time of experiment was measured by a mercury thermometer to be 27^{o}C. Air is a mixer of various gases including nitrogen ($\approx$78%), oxygen ($\approx$20%) and other gases (remaining 1%) like carbon (mono and dioxide), rare gases and water vapour. The amount of water vapour present in the atmosphere is highly dependent on the tem-

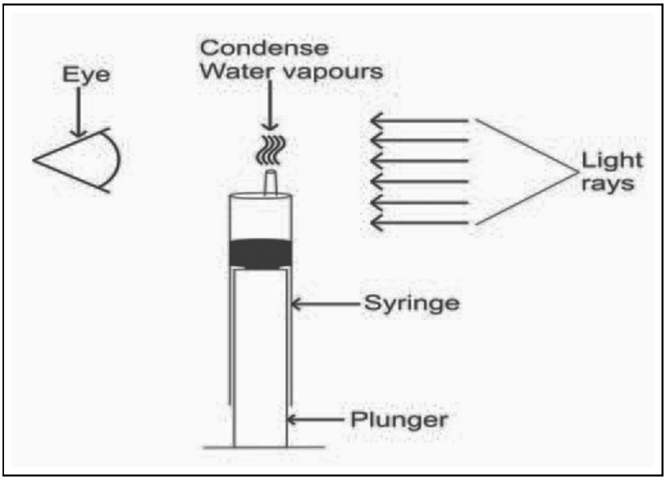


Figure 3. Setup for observing condensed water vapour.

perature of the atmosphere and elevation from the sea level. As shown in *Figure* 3, the setup consists of a light source and 5–20 ml disposable syringe. To observe the condensed vapour coming out of the syringe, one has to align their eyes to the path of the light as depicted in *Figure* 3. Any type of light source can be used for the experiment.

The demonstration of condensation of water vapour under the application of pressure using a syringe consists of four simple steps as shown in *Figures* 4(a–d).

Step [a] – Holding the syringe
As shown in *Figure* 4a, a 6 ml syringe is held using both the hands. The outlet of the syringe is blocked using the index finger.
Step [b] – Compressing the air
As shown in *Figure* 4b, the plunger of the syringe is pushed inside, compressing the air.
Step [c] – Expansion of compressed air
As shown in *Figure* 4c, the index finger is removed once the volume of 1 ml is obtained by compressing the air by the plunger. As the finger is removed, the compressed air rushes out with a small 'hissing' sound.
Step [d] – Expulsion of condensed vapour

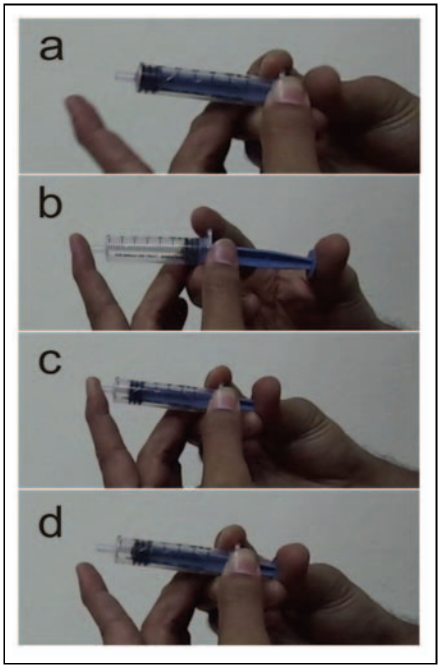


Figure 4. Images depicting the steps for compressing air using a syringe. **(a)** Holding the syringe; **(b)** Compressing the air; **(c)** Expansion of the compressed air; **(d)** Expulsion of the condensed vapour.

As shown in *Figure* 4d, the instant the index finger is removed, the condensed vapour inside the 1 ml volume is completely expelled outside by the pushing of the plunger until its volume inside the syringe is zero.
The last two steps are to be executed in progression with approximately less than a second of delay between them to observe the effect properly.

The process was recorded using a Sony cyber shot camera (8X optical zoom and 16.1 MP resolution). Snapshots from the movie were taken and labeled to create the figures. Practically, the phenomenon can be easily observed with the naked eye in the presence of ambient light sources.

3. Results and Discussions

According to the kinetic theory of gases (KTG) [*Box* 2], an increase in the pressure of gas is associated with an increase in its temperature [1, 2]. Going by this, the gas (water vapour) under

Box 2. Kinetic Theory of Gases

Kinetic theory of gases (KTG) as its name suggests, is a theoretical framework which deals with the dynamics of the gas molecules based on its translational kinetic energy. According to the theory, above its absolute temperature, every gas molecule can move in random directions. Its kinetic energy directly depends upon the absolute temperature of the gas. These molecules which are in constant motion collide with each other and with the surrounding surfaces. This collision with any surface is called as the 'pressure of the gas'. According to the KTG, the pressure of the gas is directly proportional to its absolute temperature. Similarly, the kinetic energy of the gas molecules has a direct dependence on the absolute temperature of the gas. These are given as:

$$P = \frac{1/3Nmv^2}{V}. \tag{6}$$

Here, P = Pressure of the gas, N = Number of molecules, V = Volume of the gas, m = Mass per molecule, and v = Speed of the molecules.

$$E = \frac{3}{2}\mathrm{K}T. \tag{7}$$

Here, E = Energy, K = Boltzmann constant, T = Absolute temperature.

Equation (6) gives us the pressure exerted by the gas and (7) gives us the translational kinetic energy of the gas molecules. For detailed derivations of these equations, please refer to [1, 2].

study, instead of condensing should heat up under the increased pressure inside the syringe. But to our surprise, this does not happen! The vapour does get condensed as seen in *Figures* 5a and 5b.

According to the kinetic theory of gases, an increase in the pressure of gas is associated with an increase in its temperature.

To account for this discrepancy, temperature itself comes to our rescue. In KTG, the increase in temperature is the result of compression of the gas taking place adiabatically. Meaning, there is no heat exchange (perfect thermal insulation) between the gas and the walls of the container in which the gas is being compressed. In our case, the syringe (container) is not a perfect thermal insulator. Adding to this, air compression inside the syringe is happening at a slow rate (since it is being manually compressed). Due to these conditions, even if the air inside heats up under the compression, the syringe wall conducts heat by the time the entire mass of air is compressed to its minimum achievable volume. This is called

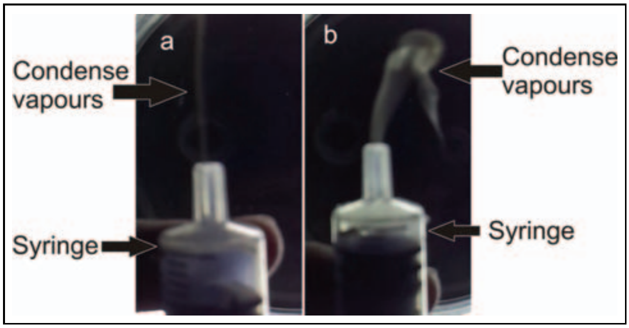


Figure 5. Condensed water vapour are coming out of the syringe nozzle. **(a)** Compression off air without water droplet inside the syringe **(b)** Compression of air with water droplet inside the syringe. Presence of droplet increases the humidity inside the syringe before compression. Difference in the amount of fog coming out of the two nozzles can be easily seen.

as an 'isothermal compression'. Hence, no cognizable heating effects are observed in this process. At the minimum volume when the air is fully compressed, we can estimate the final temperature and pressure using the equation for adiabatic compression as given in *Box* 3.

At such a high pressure, the water vapour present inside the syringe along with the air undergoes compression. Under compression, the density of the vapour increases leading to increased interactions between the water molecules. We know that water molecules possess strong hydrogen bonding. Hence, they are attracted towards each other, causing coalescence of water molecules. The effect that we observe is the formation of a white condensate on the inside wall of the syringe, as shown in *Figures* 5a and 5b respectively. As the nozzle is opened, the compressed water vapour is forced out along with the compressed air due to the pressure difference. As the water condensate comes out, it undergoes a rapid expansion against the already present atmospheric pressure. The energy required for this expansion of the condensed water vapour comes from its internal energy. This is because as expansion is a fast process, there is not enough time for the heat energy to be transferred from the surrounding to the condensate. The condensed vapour is using its internal energy, and as per the KTG, reduction in internal energy is associated with the reduction of temperature. Hence we observe the formation of a thick fog as we complete the expulsion of condensed water vapour through the syringe nozzle. This is called 'adiabatic cooling' due to 'adi-

Box 3. Adiabatic Compression of Gases

Adiabatic compression is a process in which an ideal gas is compressed in a very short instant of time. In such a short instant there is not enough time for the heat energy to dissipate from the system to its surrounding. Also ideally, there should not be any exchange of heat energy taking place between the gas molecules and the container in which the gas is being compressed. In the present case, the gas under study is air (a mixture of gases and water vapour), which is being compressed slowly and in a not so perfect insulating plastic syringe tube. The equation for estimating the final temperature and pressure is given as:

$$T_2 = T_1(V_2/V_1)^{-\gamma}, \tag{8}$$

where, (8) is the equation for estimating final temperature of the real gas, in this case air. The constant γ for air = 0.4.

$$P_2 = P_1(V_1/V_2)(T_2/T_1), \tag{9}$$

where, (9) is the equation for estimating final pressure of the real gas, in this case air. For a detailed mathematical derivation please refer [2].

We have initial values given as: P_1 = 1 atmosphere, V_1 = 6e-3 liter, V_2 = 1e-3 liter, T_1 = 300 K, T_2 = ? and P_2 = ?
Using (8), we get T_2 = 600 K or 327°C.
Using (9)), we get P_2 = 12 atmosphere.

At such a high temperature, practically, the plastic syringe should meltdown within a few seconds. But it does not happen! This is due to the fact that even though we started with the assumption of adiabatic compression of the air, in practice, the syringe wall is not a perfect insulator. As the process of compression is slow, the rise in temperature is countered by dissipation of heat energy through the walls of the syringe. Hence, the actual temperature rise is very less, and the corresponding actual rise in pressure is also less but still greater than one atmosphere. We may feel a slight increase in the temperature of the syringe wall as we compress the air. An example of adiabatic compression is the working of a diesel engine.

abatic expansion' of the water vapour.

The water vapour present in the air at room temperature condenses under the applied pressure inside the syringe. Since the demonstration of the phenomenon largely depends upon the amount of water vapour present in the air at room temperature, the experiment was performed under two different conditions. In first, the air was compressed inside the syringe as it was (*Figure* 5a). In the second condition (*Figure* 5b), an arbitrarily small volume of liquid water (droplet) was placed inside the syringe, and then the air was compressed. The difference in the formation of the fog can be easily observed in the figures. As the compressed air is released, formation of fog of higher density is seen in *Figure* 5b. However, isn't it peculiar that under applied pressure, only water vapour undergoes condensation? This can be explained on the basis of vapour pressure of the gases [*Box* 4]. In our experiment, the ambient water vapour is isothermally compressed at a slow rate. At the start of the compression, the vapour is unsaturated since it is trapped from the ambient atmosphere. As the inside volume of the syringe is gradually decreased, the pressure of the air-vapour mixture increases. The vapour density also increases towards a threshold density known as 'saturated vapour density'. As this state is reached, further decrease in the volume causes the vapour to condense and appear as a white formation on the inner side of the circular wall of the syringe. This is condensed water. The sudden release of the compressed mixture of air causes an adiabatic expansion of the mixture and further cooling of the already condensed water vapour. This is seen as the foggy expulsion from the syringe nozzle at the completion of steps 'c' and 'd' in our experiment. If we compare the vapour pressures of all the other gases present in the air, we see that most of them are in a gaseous state at room temperature with high vapour pressures. If we assume that real gases obey the ideal gas laws with some approximations, the magnitude of pressure required to compress these gases at room temperature is quite high. Hence, during the manual compression of trapped air inside the syringe, only water vapour present in the syringe condenses at room temperature.

If we assume that real gases obey the ideal gas laws with some approximations, the magnitude of pressure required to compress these gases at room temperature is quite high.

Note that depending upon the seasonal and geographical variations, the amount of water vapour in the atmosphere changes. This implies that the compression of air in the syringe without water droplet carried out in the summer will yield less amount of condensed vapour compared to the same experiment carried out in the rainy season. Similarly, the experiment carried out in a desert will yield less condensed vapour compared to the experiment carried out near a sea beach or large waterbody.

Box 4. Vapor Pressure

Let us assume that a closed container is half filled with water at a temperature of 25°C. From the KTG, we know that every molecule has kinetic energy corresponding to the absolute temperature of the gas. Hence, the water molecules will also have some kinetic energy. The molecules with the highest kinetic energy will have enough energy to leave the surface of the liquid water as water vapour by breaking their attractive surface bonds into open air – the process is called evaporation.

In a closed container, when the rate of molecules leaving the surface of the liquid water is equal to the rate of the molecules returning back to it, the system is said to be in 'thermodynamic equilibrium'. In such a system, the presence of water vapour, since it is in gaseous state, will have certain pressure associated with it (*Box* 2). This pressure is called the 'saturated vapour pressure'. At this pressure, vapour has the maximum density. Any further increase in the density of the vapour will cause the vapour to condense to form water droplets. For different gases/vapour, vapour pressure has different values at ambient temperature. *Table* 1 summarizes the approximate values of the vapour pressure of the gases [3]. (Note that 760 Torr = 760 mmHg = 1 atmosphere.)

Name of the gases/vapour	Vapour pressure
Water	17.5424 Torr at 20°C
Carbon dioxide	42753 Torr at 20°C
Oxygen	$\gg$ 750 Torr at −183°C
Helium	$\gg$ 750 Torr at −268°C
Nitrogen	$\gg$ 750 Torr at −196°C

Table 1. Values of vapour pressure for different gases/vapour.

As seen from *Table* 1, water vapour has the lowest value of vapour pressure at ambient temperature compared to other gases which are part of the air used in our compression experiment. Condensing other gases by applying manual pressure alone at ambient temperature (20°C) is impossible. For example, the value of vapour pressure of oxygen is much greater than 750 Torr at −183°C. This signifies that oxygen, under one atmospheric pressure at temperature of −183°C completely evaporates. As the temperature of gaseous oxygen is increased upto the ambient temperature (20°C), the vapour pressure of the oxygen becomes >750 Torr. As evident, application of low manual pressure (such as on the trapped air inside the syringe) will not cause the condensation of oxygen. For such gases to be condensed, simultaneously their temperature needs to be decreased while applying high pressure. The same explanation goes for the other gases in the table. Hence, only water vapour present in the atmospheric air can be condensed at room temperature using manually applied pressure.

4. Conclusion

There are not many experiments/demonstrations in the school, undergraduate and postgraduate coursework which improves our understanding of thermodynamics. This simple experiment could be used for demonstrating the concept right from higher-secondary to postgraduate students. Formation of fog under varied atmospheric conditions, and the condensation of gases could be discussed using this demonstration model as an analogy. The demonstration covers concepts like isothermal compression, condensation of water vapour under pressure, adiabatic expansion, and compressibility of gases. It is widely seen that parameters like high cost and unavailability of materials hinder practical learning. Hence, it is crucial that simple experiments be designed from materials procured easily from our surroundings and demonstrated along with theoretical studies for clearer understanding of concepts and inspiring scientific interest.

Suggested Reading

[1] **Sears and Zemansky, *University Physics with Modern Physics*, 13th Edition, ISBN 13: 978-0-321-69686-1; ISBN 10: 0-321-69686-7, Addison-Wesley publisher, p.1598, 2011.**

[2] **Douglas C Giancoli, *Physics for Scientists and Engineers with Modern Physics*, 4th Edition, ISBN-13: 978-0-13-149508-1, ISBN-10: 0-13-149508-9, Pearson Education, Inc., p.1221, 2008.**

[3] `https://en.wikipedia.org/wiki/Vapor_pressure`