Experiment 14
Introduction to Thermodynamics

Advanced Reading:
Serway & Jewett - 8th Edition
Chapters 19, 20 & 21

Equipment:
Balloons
Arbor Scientific Fire Syringe
Dewar(s)
Liquid nitrogen
Masses
table clamp, rod and right angle clamp
vernier caliper

Objective:

The objective of this experiment is to examine a few fundamental concepts of thermodynamics. Particular interest will be given to ideal gas laws internal energy, adiabatic compression and the first law of thermodynamics.

Theory:

While thermodynamics is one of the most important and useful physics concepts, in today's physics curricula it has been somewhat relegated to 'stepchild' status and is usually barely touched upon at the end of the end of the first semester. *This has some historical precedence since Max Planck (one of the founders of quantum physics) was advised in 1874 not to study physics " because after the discovery of the first two laws of thermodynamics, all that was left for a theoretical physicist to do was to tidy up a few loose ends."*

Suppose that the temperatures of different gases are measured at different pressures with a gas thermometer. Experiments show that the thermometer reading is nearly independent of the type of gas used as long as the gas pressure is low and the temperature is well above the point at which the gas liquefies (see Figures 19.3 and 19.5). If a straight line is extended towards negative temperature it is seen that "in every case the pressure is zero when the temperature is -273.15 ºC. This is the basis for the absolute zero temperature scale.

For a gas, the quantities volume V, pressure P and temperature T are related for a sample of gas of mass m. In general the equation that interrelates these quantities is called an **equation of state**. If the gas is maintained at low pressure (or density) the equation is 'quite simple' and can be determined from experimental results. **A low density gas is called an ideal gas.**

Suppose an ideal gas is confined to a cylindrical cylinder whose volume can be varied by means of a movable piston (see Figure 19.12). If we assume that the cylinder does not leak and the mass of the gas remains constant the following information can be obtained:

• For a gas at constant temperature, its pressure is proportional to volume (Boyle's law.)

• For a gas with constant pressure the volume is directly proportional to the temperature **(Charles's law.)**

• For a gas with volume, the temperature is directly proportional to the temperature (Gay-Lussac's law.)

These observations are summarized by the **equation of state for an ideal gas**:

$PV = nRT$ Equation 1

where n is the number of moles of a substance.

In this experiment we will use the equation of state, a balloon and liquid nitrogen to estimate the value of absolute zero.

The first law of thermodynamics is one of the most important laws of physics. First it is important to make a distinction between **internal energy** and **heat**. Internal energy is all of the energy of a system associated with its microscopic components (when viewed from a reference frame as rest with respect to the center of mass.) Heat is defined as the transfer of energy across the boundary of the system due to a temperature difference between the system and its surroundings. Heat and work are ways of changing the energy of a system.

The first law of thermodynamics is a special case of the law of conservation of energy that describes processes in which only the internal energy internal energy ΔE_{int} changes and the only energy transfers are by heat Q and work W:

$$\Delta E_{\text{int}} = Q + W \qquad \text{Equation 2}$$

Internal energy is a state variable like pressure, volume and temperature.

An **adiabatic process** is one in which no energy leaves of enters the system by heat (i.e., Q= 0.) This can be achieved by either insulating the walls of the container or by performing the process rapidly so that there is negligible time for energy to transfer by heat. Thus Equation 2 becomes

$$\Delta E_{\text{int}} = W \text{ (adiabatic process)} \qquad \text{Equation 3}$$

This shows that if a gas is compressed adiabatically, such that W is positive then ΔE_{int} is positive and the temperature of the gas increases. *This is how a diesel engine combusts the fuel air mixture while using no spark plugs.*

A process that occurs at constant temperature is called an **isothermal process**. A plot of P vs. V at constant temperature for an ideal gas yields a hyperbolic curve called an isotherm. *The internal energy of an ideal gas is a function of temperature only*. Thus $\Delta E_{\text{int}} = 0$ and we can conclude that for an isothermal process the energy transfer Q must be equal to the negative of the work done on the gas, i.e., $Q = -W$. Any energy that enters the system is transferred out of the system by work resulting in $\Delta E_{\text{int}} = 0$.

As noted above an adiabatic process is one in which no energy is transferred between a system and its surroundings. The diesel is one example and another is the slow expansion of a gas that is thermally insulated from its surroundings. All three variables in the ideal gas law (P, V and T) change during an adiabatic process.

For an adiabatic process undergoing infinitesimal changes in volume dV and temperature dT , the first law of thermodynamics can be used to obtain the following relationships.

$$PV^{\gamma} = \text{constant} \qquad \text{Equation 4}$$

and

$$P_i V_i^{\gamma} = P_f V_f^{\gamma} \qquad \text{Equation 5}$$

where γ is the ratio of constant pressure specific C_p heat to constant volume specific heat C_v, i.e., $\gamma = \frac{C_P}{C_v}$.

See section 21.2.

Procedure:

Part 1- Ideal gas laws and Balloons - Estimation of temperature of liquid nitrogen and absolute zero

1. You will have a total of four liquid nitrogen setups. The balloons will be filled with either air or nitrogen. You will observe and record the behavior of the balloons.

2. If the length (height) of a balloon filled with nitrogen gas is measured before it is put into liquid nitrogen and the length is measured after some sufficient time one can estimate both the temperature of liquid nitrogen and absolute zero. See questions 2 and 3.

Part 2- Adiabatic processes

2. Unscrew the blue cap of fire syringe (or fire piston) and **gently** measure the outside diameter of the o-rings on the piston. Record this value.

Figure 1- Fire syringe

3. Next, use **10 cm as the height** of the inside of the cylinder (the part that the piston goes into). **This is your initial height.**

4. Making sure your hands are dry tear off a small piece of cotton and gently place (slide) it to the bottom of the cylinder using a coffee stirrer. Be careful not to compress it too much. Screw blue cap back on.

5. Using a rapid motion compress the piston handle downward and observe what you see. You will probably have to make more than one compression- one quickly after another. What do you observe?

6. You will need to estimate how far down you were able to compress the cylinder **in order to determine your final height.** See question 5.

Part 3- Isothermal compression

7. Using the fire syringe and masses you will now plot **mass vs. height** (this is essentially the same as pressure vs. volume since the cylinder is uniform.)

8. Use the ruler stuck to the side of the cylinder to measure height. See figure 2 below. Add 200 g to mass to account for the mass of the piston and friction. Plot mass vs. height and curve fit.

Part 4- Ideal gas Law

9. In the front of the lab is a container of compressed gas with approximately 100 psig (i.e., pounds per square inch *gauge*). You will measure its weight and then leak the air out. You will then measure its weight again to determine the weight of air that was in the tank initially.

10. While the tank is leaking someone will point a remote temperature gauge at the tank to see if the temperature of the tank changes while the tank is leaking.

Questions

1. What did you observe about the behavior of the balloon filled with air compared to the balloon filled with nitrogen. What did you observe about the balloon filled with nitrogen at a higher pressure (and the one you squeezed) compared to the the one at the lower pressure. How do you account

fo for the difference?

2. Using the ideal gas law we know that

$$\frac{P_{initial}V_{initia}}{T_{initia}} = \frac{P_{final}V_{final}}{T_{final}} \qquad \text{Equation- 6}$$

For a balloon filled with nitrogen gas starting at low

pr pressure (i.e., slightly above atmospheric pressure) we can assume $P_1 = P_2$ - **Charles's law is applicable**. Also, since the balloons used in this experiment are cylindrical (with volume given by $\pi r_{balloon}^2 h_{balloon}$) we can write Equation 6 as

$$\frac{P_{initial}\pi r_{balloon}^2 h_{bal\_initial}}{T_{initial}} = \frac{P_{final}\pi r_{balloon}^2 h_{bal\_final}}{T_{final}} \qquad \text{or}$$

$$\frac{h_{bal\_initial}}{T_{initial}} = \frac{h_{bal\_final}}{T_{final}} \qquad \text{Equation 7}$$

Using Equation 7 and a starting and final height as determined (or given) in lab calculate temperature of liquid nitrogen (convert temperature to Kelvin.) give the results assuming 10% uncertainty.

3. Since volume of a gas approaches zero as the temperature of the gas approaches absolute zero, if volume vs. temperature is plotted, the x intercept (i.e., temperature) is absolute zero.

Plot height of balloon for room temperature and the height of the balloon at the boiling point of liquid nitrogen (using true value). Use x intercept to determine absolute zero.

4. From part 3 **plot mass vs. one over height** (i.e., 1/height) and perform a linear fit. How well does your data support Equation 1 (i.e., is your plot linear and is $PV = \text{constant}$ for your data?) Do you think that the assumption that the compression is isothermal is a valid assumption? Why or Why not.

5. Using example 21.3 on page 608 of text (Serway & Jewett) as a guide, calculate both the final pressure and final temperature in part 2 (**Adiabatic compression**). To simplify calculations assume $\frac{V_{initial}}{V_{final}} = \frac{H_{initial}}{H_{final}}$. Also, assume starting temperature was 22 degrees Celsius (if you did not measure).

6. The volume of the green army tank is approximately **one cubic foot**. Using the starting pressure (~100 psi) and the ideal gas law $PV = nRT$ compare the left hand side of the equation to the right hand side.

The easiest way of getting energy of the left hand side is to multiply starting volume times pressure by 144 since $\frac{1\ lb_f}{in^2} \times \frac{144 in^2}{1\ ft^2}$. This will give you units of $lb_f\text{-}ft = ft\text{-}lb_f$. Show all units of calculations.

The number of moles can be determined by changing pounds of air to grams and using the gram molecular mass of air (29 grams/mole). Answer in SI units is Joules.

Convert joules to ft-lbs of energy (use 1 joule = 0.7376 $ft\text{-}lb_f$) and compare to the some of the energies given on the following website

http://en.wikipedia.org/wiki/Muzzle_energy