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# LabWorks

## An Introduction to Thermodynamics and the Gas Laws

After completing this module you will be able to:

1. Define the concepts of calories and specific heat.
2. Compare the energy changes involved in chemical and physical changes
3. Perform heat-gain and heat-loss calculations
4. Better appreciate real-world examples using heat loss and heat gain.
5. Define the effects of temperature, pressure, and volume on gases.
6. Determine, from experimental data, the effect of temperature and pressure on the volume of a gas.
7. Graphically deduce Boyle's and Charles' Laws.
8. Apply the Gas Laws to real-world examples.

Thermistor Calibration	Heat Associated with Chemical Change 1	Chemistry of Antifreeze: Density and Freezing Pt. 2
Boiling-Point Elevation by Dissolved Substances 3	Factors Affecting Freezing- Pt. Depression in H <sub>2</sub> O 4	Molar Heat of Reaction For HCl and NaOH 5
Charles' Law 6	Boyle's Law 7	

Experiment	Title	Math Skills
1	Thermistor Calibration	LabPro Interface
2	Heat Associated with Chemical Change	Single unknown equations
3	Chemistry of Antifreeze: Density and Freezing Pt.	Linear regression, scaling
4	Boiling-Point Elevation by Dissolved Substances	Programming, interpreting
5	Factors Affecting Freezing-Point Depression in H <sub>2</sub> O	Programming, interpreting
6	Molar Heat of Reaction for HCl and NaOH	Programming, interpreting
7	Charles' Law	Ratios
7	Boyle's Law	Graphing, linear regression

# Thermistor Calibration

## 1.0 Introduction

This tutorial explains how to do temperature measurements using the DataMate program in Time Graph mode on a TI graphing calculator connected to a LabPro interface with a temperature probe. The LabPro interface reads signals from a variety of sensors and transmits those signals electrically to a calculator or computer.

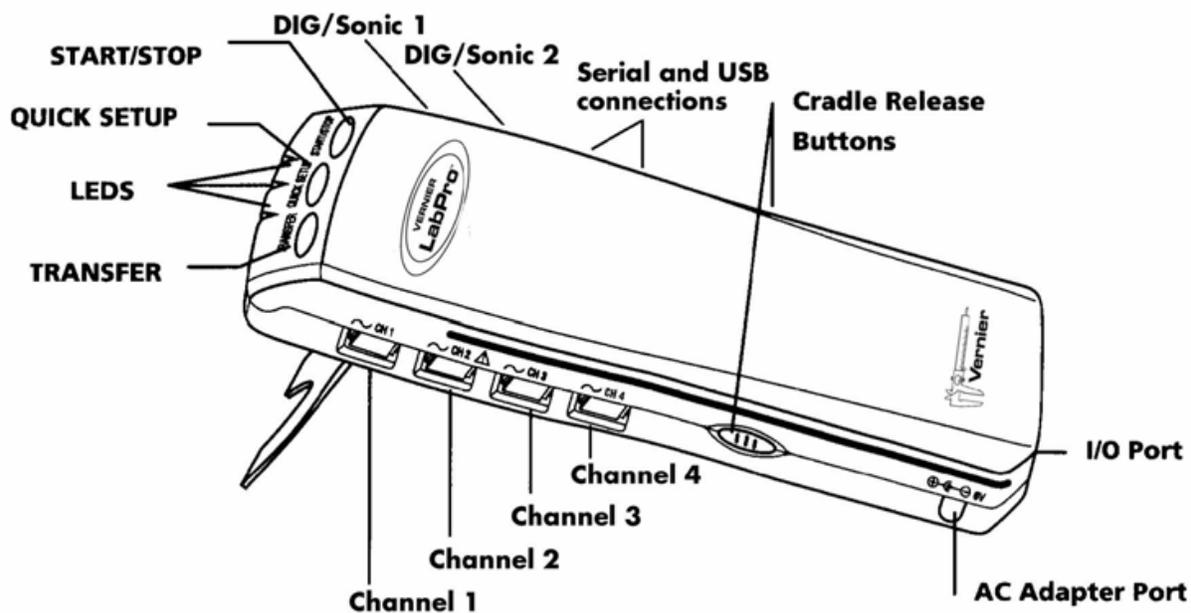
Sensors can be divided into two basic types—analogue and digital. Examples of analogue sensors are temperature probes, pH sensors, force sensors, oxygen gas sensors, *etc.* Up to four analogue sensors can be connected to LabPro. Four jacks for the analogue sensors (CH 1 – CH 4) are located on the same side as the AC adapter port. The analogue ports accept British Telecom-style plugs with a right-hand connector. Examples of digital sensors are motion detectors, radiation monitors, photogates, and rotary motion sensors. Up to two digital sensors can be connected to LabPro. The digital ports (DIG/SONIC), which accept British Telecom-style plugs with a left-hand connector, are located on the same side as the serial and USB computer connections.

## Materials and Equipment

1. LabPro interface (see next page for drawing thereof)
2. TI graphing calculator (TI-84 Plus provided by us, but others will work)
3. Temperature probe (Vernier)
4. DataMate program
5. Silver graphing link cable (to connect LabPro Interface to computer)
6. Computer with Logger Pro 3.3 software

## LabPro Interface

The LabPro interface has three buttons, three LEDs, four analogue channels (CH 1, CH 2, CH 3, and CH 4), two digital channels (DIG/SONIC 1 and DIG/SONIC 2), a serial computer connection, a USB computer connection, a piezo speaker, and a calculator I/O port.



## Procedure

### 1. Plug the link cable into the link port of the TI graphing calculator and LabPro

You will collect data on TI graphing calculator and use LabPro only as an interface (I/O port).

### 2. Plug the stainless-steel temperature probe into Channel 1

Connect the sensor to the appropriate port (Channel 1 for temperature probe). An important feature of LabPro is its ability to detect auto-ID sensors, and automatically set up an experiment. Remember that the temperature probe does not need calibration, so skip any instructions asking you to calibrate the sensor!

### 3. Provide power to LabPro (AC adapter)

To use LabPro with the AC power supply, plug the round plug on the 6-volt power supply into the side of the interface. Shortly after plugging the power supply into the outlet, the interface will run through a self-test. You will hear a series of beeps and blinking lights (red, yellow, then green) indicating a successful power up. If the self-test is not successful, unplug everything and plug it in again. You cannot proceed until you hear the beeps!

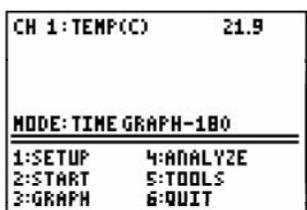
### 4. Start the DataMate program on TI-84 Plus graphing calculator

Use the following steps to start the DataMate program on your calculator:

1. Turn the calculator ON
2. Press APPS (Applications)
3. Press number under DataMate program
4. This temporary screen will appear:



After the program begins, the calculator will try to communicate with the interface and check for auto-ID sensors. If this communication is successful, the main screen (shown below) will appear.



In this example a temperature probe was connected and DataMate automatically identified and displayed the channel it was connected to, giving the current temperature reading. DataMate also sets up a default time graph experiment for auto-ID sensors. In this example, an experiment was set up to collect data for 180 seconds.

To change time settings:

1. Press 1 (for Setup in Main Menu)
2. Select Mode and press Enter
3. In Select Mode menu press number 2 (Time graph)
4. Select Options and then Change time settings
5. Pick time interval between the points and total time of the experiment (should be the same for all three experiments!)
6. You are now ready to collect data

### **5. Collect data with DataMate on a TI graphing calculator**

Use the following steps to collect data on the DataMate program. Note that the TI calculator can only hold a maximum of three sets of data from a single auto-ID sensor such as a temperature or pressure probe:

1. Press 2 (for Start in Main Menu)
2. When experiment is done (graph will automatically appear) press Enter (to go back to Main Menu). If you want to stop experiment run earlier click STO>
3. Press 5 (for Tools in Main Menu)
4. Press Store Last Run
5. Repeat it for the second set of data
6. At the end of the third set of the data there is no need for Store Last Run step, instead press 6 (for Quit in Main Menu) (last measurement will be stored automatically)

## **6. Transmit data to computer**

Use the following steps to retrieve data from the TI calculator by the computer:

1. All computers have Logger Pro 3.3 software on them
2. Connect the calculator and the computer I/O (calculator) - USB (computer) silver graphing cable
3. Start Logger Pro 3.3 program by double-clicking on the icon
4. Click on the little calculator icon in the top line on the program menu (Import from TI device should appear when you put mouse on little calculator icon)
5. Pick the TI-GRAPH LINK™-USB from the Port pull down menu. Skip step 6 if TI-GRAPH LINK™-USB is already present.
6. Click on Scan for device
7. From Choose data to import, highlight L1 (time) and L2, L3, and L4 lines (Temperature) by holding Ctrl button
8. Click OK
9. You can make:
  - a. Single graph (for example only L2 vs. L1)
  - b. All three measurements appear on the screen at the same time by clicking on the temperature axes (y-axes), choose option More and then check the three columns to be displayed (L2, L3, L4)
  - c. Click OK
10. To save graph in format that you can later use with Excel, Origin, or MATLAB go to:
  - a. File
  - b. Export as
  - c. Text
11. Save that text file (\*.txt) on your flash drive
  - a. Click No for “Do you want to save changes you made to Untitled.cmb1”
12. Repeat another set of three measurements if needed.
13. With Excel
  - a. Open the Excel program
  - b. Under the File pick Open, mark the text file, click OK
  - c. Importing text file: click Finished
14. With MATLAB
  - a. Open the MATLAB program
  - b. At the > prompt, type load filename
  - c. The variable with the data is now in memory with the name ‘filename’

## 2.0 Thermodynamics Background

Almost all chemical reactions either absorb or release heat. These experiments are designed to introduce you to three concepts of thermodynamics:

- 1) The measurement of heat energy
- 2) In a closed system, the heat gained by one object must equal the heat lost by another.
- 3) Energy is liberated or absorbed (as heat) during physical and chemical changes.

Heat and temperature are not the same. When heat is added to a substance, the temperature rises, but the amount of temperature rise depends on the type and quantity of substance. Temperature depends on heat, but is not directly or simply a measure of the amount of heat. You feel that something is hot or cold, because either it transfers heat to you or you to it.

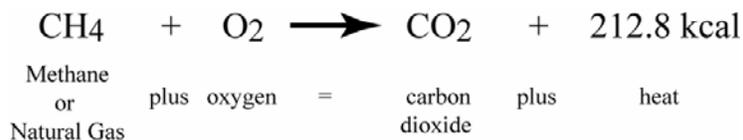
Temperature is a measure (in degrees) of the average rate of motion of molecules in a substance. Heat energy is often measured in calories. The calorie is the amount of heat required to raise the temperature of exactly 1 gram of water from 14.5°C to 15.5°C. For all practical purposes, the calorie is simply defined as the amount of heat required to raise the temperature of 1g of water 1°C.

The amount of heat required to raise 1 g of a substance 1°C in temperature is called the specific heat of the substance. Specific heat is a characteristic property, just like density or boiling point.

## 2.1 Examples of How Heat Loss and Heat Gain Are Used in Everyday Life

### 2.1 Combustion of Natural Gas in a Water Heater or Furnace

The combustion of natural gas in a water heater, furnace, or modern power plant is a common heat-producing (exothermic) reaction. Natural gas is mostly methane, CH<sub>4</sub>. The combustion of methane results in the following reaction:



According to the reaction, 212.8 kcal (or 212,800 calories) of heat are produced by the oxidation of 16 grams (one gram-molecular weight, or one mole) of methane (natural gas).

### 2.2 Energy for Living Organisms

Not all oxidation reactions proceed rapidly and vigorously by combustion. In living organisms, oxidation reactions proceed much more slowly and are controlled by enzymes. Some of the energy liberated by these reactions is used to raise the body temperature in warm-blooded animals, and some of it is used for other vital processes, such as muscle contraction,



## 3.0 The Gas Laws

### 3.1 Background

Of the three states of matter (solid, liquid, and gas) gases are the most affected by the surrounding environment. Most solids undergo very little or no variation in shape and volume due to changes in temperature, pressure, or the type of container, *etc.* Although liquids assume the shape of their container, they tend to maintain a constant volume. While thermal expansion / contraction is noticeable in the case of both solids and liquids, it is relatively insignificant. External pressures do not affect matter in either of these states appreciably. However, a given amount of gas placed in an empty (evacuated) container of any shape will immediately expand to fill the container completely and uniformly. If pressure is exerted on a sealed, flexible container filled with gas, the gas will become compressed until the pressure exerted by the gas on the walls of its container equals the outside pressure applied. If the gas is heated, it will either apply more pressure on its container or expand to a greater volume.

### 3.2 Boyle's Law

Robert Boyle, a British chemist, was one of the first scientists to study gases quantitatively. In one set of experiments, Boyle established a relationship between the pressure and volume of a gas at constant temperature. In order to arrive at a physically significant law, Boyle simplified the problem by doing his experiments under controlled conditions. He kept the mass  $m$  of gas constant (*i.e.*, no leaks in the container), and he kept the temperature constant. Under such circumstances, the relationship between pressure and volume is known as Boyle's Law. At a constant temperature, the volume ( $V$ ) occupied by a finite mass of gas is inversely proportional to the applied pressure ( $P$ ). Mathematically, this relationship can be stated as follows:

$$(1) \quad V \propto 1/P$$

Where the  $\propto$  symbol represents proportionality. Another way of stating Boyle's Law is: for a given amount of gas at a constant temperature, the product of the pressure and the volume is a constant. Thus, by replacing the proportionality constant in the previous equation with a constant  $k$ , we will have another form of Boyle's Law.

$$(2) \quad V = k \cdot \left(1/P\right)$$

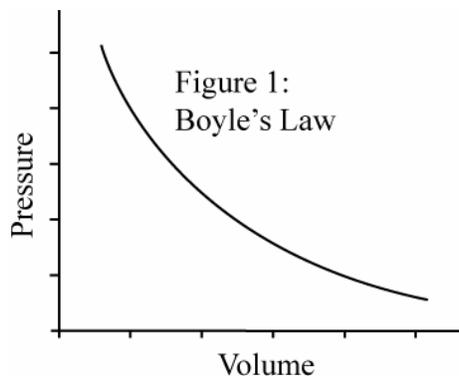
$$(3) \quad P \cdot V = k$$

When using two sets of pressure/volume conditions (1 and 2), Boyle's Law can be expressed:

$$(4) \quad P_1V_1 = P_2V_2$$

where the subscript 1 refers to the initial state and the subscript 2 refers to the final state. In the Boyle's Law experiment, you will examine the relationship between pressure and volume

for a gas. Graphing this relationship for an ideal gas, shown in Fig. 1, results in a hyperbola. As we increase the pressure, passing from state 1 to state 2, the product  $PV$  is constant.



### 3.3 The Effect of Mass

The value of the constant,  $k$ , in Boyle's Law depends on several factors. First, imagine that the quantity of gas changes while the temperature and volume remain constant. Pressure is caused by the impact of molecules as they strike the walls of the container. Therefore, doubling the mass (*i.e.*, doubling the number of gas molecules) at a constant temperature will cause the number of impacts per second to double, and the pressure will be doubled.

Experimental verification shows that the product  $PV$  is proportional to the mass of gas.

$$(5) \quad \frac{P_1 V_1}{m_1} = \frac{P_2 V_2}{m_2} \text{ (at constant temperature)}$$

In other words, if  $m$  is increased, then so is  $PV$ ; and  $PV/m$  remains the same.

In general, when dealing with different kinds of gases, such as hydrogen and oxygen, it is important to use equal number of molecules rather than equal masses. For instance, there are actually more molecules in 8 g of  $H_2$  (whose molecular weight is 2.0) than in 80 g of  $O_2$  (whose molecular weight is 32). For  $H_2$ ,  $8 \text{ g} \times 1 \text{ mol} / 2 \text{ g} = 4 \text{ moles}$ ; therefore, it has  $4 \times 6.02 \times 10^{23}$  molecules. Similarly, 80 g of  $O_2$  is only  $80/32 = 2.5 \text{ moles}$ , and contains only  $2.5 \times 6.02 \times 10^{23}$  molecules. All things being equal, we expect that the pressure exerted by hydrogen would be greater by a ratio of 4 to 2.5. Thus, the following equation applies:

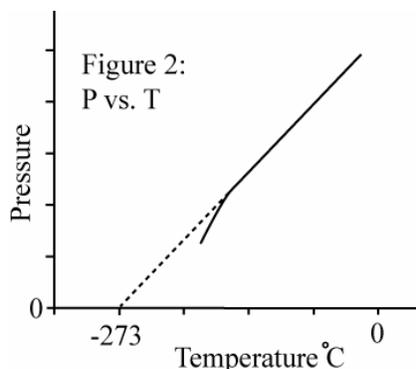
$$(6) \quad \frac{PV_1}{n} = \text{constant (at constant temperature)}$$

where  $n$  is the number of moles of gas present.

### 3.4 Charles' Law

Around 1800, French scientist and balloonist Jacques Charles began studying the effect of increasing temperature on gases. He observed that the rate of thermal expansion is constant

and is the same for all gases as long as the pressure is constant. Charles examined the effects of temperature upon the pressure exerted by a confined gas with volume and mass remaining constant. Fig. 2 shows that the pressure increases uniformly as the temperature increases. Thus, the graph of pressure vs. temperature is a straight line.



In 1848, Lord Kelvin, a British physicist, noted that by extending different temperature/volume lines (for gases at various but constant pressures) back to zero volume, he always found the same intercept. The intercept on the temperature axis is  $-273.15^{\circ}\text{C}$ . Kelvin named this temperature absolute zero. The Kelvin absolute temperature scale, in which  $\text{K} = ^{\circ}\text{C} + 273.15$ , is named in his honor. The volume/temperature relationship for gases using the absolute temperature scale is known as Charles' Law. At a constant pressure, the volume of a finite amount (mass) of gas is directly proportional to the temperature of the gas. Mathematically, Charles' Law can be stated as follows (using the absolute temperature scale)

$$(7) \quad V \propto T$$

$$(8) \quad V = k \cdot T$$

This relationship leads to:

$$(9) \quad \frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Note that in Figure 2, the graph of  $P$  vs.  $T$  is shown as a dashed line below a certain temperature. This is to indicate that the pressure would become zero at some temperature if it continued to decrease at the same rate as it does near room temperature. Any ideal gas is assumed to behave this way; but in actual practice, before the gas pressure becomes zero, the gas will liquefy at some temperature and then, of course, it will no longer even be a gas. As the temperature of the gas decreases, the pressure deviates somewhat from the straight-line relationship. We expect such a behavior on the basis of what we know about molecules. The slower-moving molecules are more affected by mutual cohesive forces, and the pressure exerted by them will drop off a little. Finally, when the temperature is low enough and the

molecules slow enough, the cohesive forces causes the molecules to stick together as a liquid. The mathematical equation of the straight line in Fig. 2 can be written as

$$(10) \quad P_T = P_0 \cdot (1 + b \cdot T)$$

Therefore, using the assumption that gases are nearly ideal, we can rework equation (10) into a form that makes calculations very simple. Using the numerical value of  $0.00366 / ^\circ\text{C}$  for  $b$ :

$$(11) \quad P_T = P_0 \cdot (1 + 0.00366 \cdot T)$$

At what temperature would the pressure be zero, if the straight line continued to be a true description of the course of events? From equation (11), it is evident to make  $P \cdot T = 0$ , we must have  $(1 + 0.00366 \cdot T) = 0$ . That is, the required temperature is:

$$(12) \quad T = -\frac{1}{0.00366/^\circ\text{C}} = -273^\circ\text{C}$$

Where  $P_T$  is the pressure at temperature  $T$ ,  $P_0$  is the pressure at  $0^\circ\text{C}$ , and  $b$  is the pressure coefficient for the gas. Experimental data shows that the constant  $b$  is practically the same for all gases and is about  $0.00366 / ^\circ\text{C}$ . for those gases, such as hydrogen, which are nearly "ideal." The pressure coefficients for helium and oxygen are  $0.00367 / ^\circ\text{C}$ . Such differences are minor and show that no gas is totally ideal.

Thus it can be seen that the temperature at which the pressure would become zero is the same for all ideal gases if the gas remained an ideal gas all the way down to zero pressure. Again, we call this temperature absolute zero. In the laboratory, we can measure the pressure of a gas at different temperatures and extrapolate the graph down to zero pressure, see Fig. 3. Precise experiments have given  $-273.15^\circ\text{C}$  as the value of this important constant, but we can usually round the value to  $-273^\circ\text{C}$ .

In the Charles' Law experiment you will examine the effect of temperature on the volume of a gas. You will also extrapolate the temperature/volume relationship to obtain a temperature value for absolute zero.

### 3.5 Ideal Gas Equation

From Charles' Law, we know that volume varies proportionally with absolute temperature. From Boyle's Law, we know that volume varies with an inverse proportionality with pressure. Combining the two laws, we obtain:

$$(13) \quad V = k' \cdot \left(\frac{T}{P}\right) \quad \text{or} \quad PV = k' T$$

where the constant  $k'$  represents the number of gas molecules present. We can replace  $k'$

with  $n$  and  $R$  to denote the number of gas molecules and the gas constant:

$$(14) \quad k' = nR$$

Combining equations (13) and (14), we get:

$$(15) \quad PV = nRT$$

which is referred to as the ideal gas equation of state, where  $R = 8.314 \text{ dm}^3 \text{ kPa} / \text{mol} / \text{K}$ . Additionally, the number of moles,  $n$ , is equal to the mass  $m$  divided by the molecular mass  $M$ , and can be inserted into equation (15) to yield:

$$(16) \quad PV = \frac{mRT}{M}, \text{ or } M = \frac{mRT}{PV}$$

### 3.6 Dalton's Law of Partial Pressure

As we have already discussed, gases exert pressure on their enclosures. Mixtures of gases exert pressure on their enclosures as well, but each component of the gas mixture is responsible for only a fraction of the total exertion. John Dalton was the first scientist to form a hypothesis about how much each component of a mixture of gases contributes to the overall pressure exerted. We call the pressure exerted by a single type of gas a partial pressure. Dalton's hypothesis states that the total pressure in a container is the sum of the partial pressures of the gases in the container.

The contribution by a single gas to the overall pressure exerted on a container depends largely on the amount of gas present. For equal volumes of gases in a container, there are different numbers of molecules present for each gas. Since each gas is at the same temperature in a mixture, the pressure due to a single gas is due to the number of molecules of the gas in relation to the overall number of molecules.

## **4.0 Examples of How Boyle's and Charles' Laws Are Used in Everyday Life**

The measurement and effects of temperature and pressure are very important to our everyday life. Let us first discuss the significance of atmospheric pressure to the forecasting of weather, creation of winds, the measurement of altitude for aircraft, and the cause of atmospheric circulation.

### **4.1 Significance of Atmospheric Pressure**

The average pressure exerted by the atmosphere is approximately 14.7 pounds per square inch at sea level (also known as 1 atmosphere). This means that a column of air of one inch square extending from sea level to the top of the atmosphere would weigh about 14.7 pounds. The actual pressure at a given place and time, however, depends upon several factors, including altitude, temperature, and the density of the air.

### **4.2 Measurement of Atmospheric Pressure**

A barometer generally consists of a column of mercury in a glass tube. It is sealed at one end and calibrated in inches of mercury. An increase in pressure forces the mercury higher in the tube; a decrease allows some of the mercury to drain out, reducing the height of the column. In this way, changes of pressure are registered in inches of mercury. The standard sea-level pressure expressed in these terms is 29.92 inches at a standard temperature of 15°C (59°F).

If all weather stations were at sea level, the barometer readings would give a correct record of the distribution of atmospheric pressure at a common level. To achieve a common level, each station translates its barometer reading in terms of sea level pressure. A change of 1,000 feet of elevation makes a change of about one inch on the barometer reading. Thus, if a station located 5,000 feet above sea level found the mercury to be 25 inches high in the barometer tube, it would translate and report this reading as 30 inches. Because the rate of decrease in the atmospheric pressure is fairly constant in the lower layers of the atmosphere, the approximate altitude can be determined by finding the difference between pressure at sea level and the pressure at the given atmospheric level. In fact, the aircraft altimeter is an aneroid barometer with its scale in units of altitude instead of pressure.

### **4.3 Effects of Altitude on Atmospheric Pressure**

We can conclude that atmospheric pressure decreases as altitude increases and that the pressure at a given point is a measure of the weight of the column of air above that point. As altitude increases, pressure decreases along with the weight of the air column. This decrease in pressure has a pronounced effect on aircraft takeoffs, rates of climb, and landings. An airplane that requires a 1,000 foot runway for takeoff at a sea level airport will require a runway almost twice as long to takeoff at an airport 5,000 feet above sea level. The purpose of the takeoff run is to gain enough speed to generate lift from the passage of air over the wings. When the air is less dense, more speed is required to obtain enough lift for takeoff; hence, a longer ground run is necessary. It is also true that the engine is less efficient in thin air, and the thrust of the propeller is less effective. Generally, atmospheric pressure drops by half every 5 km increase in altitude.

#### **4.4 Effects of Differences in Air Density**

Differences in air density caused by temperature changes result in pressure changes. This, in turn, creates motion in the atmosphere, both vertically and horizontally (currents and winds). This action, when mixed with moisture, produces clouds and precipitation. In fact, these are all the phenomena called weather. In general, a marked fall of the barometer indicates the approach of bad weather and marked rise indicates the clearing of the weather.

#### **4.5 The Cause of Atmospheric Circulation**

Pressure and temperature changes produce two kinds of motion in the atmosphere: vertical movement of ascending and descending currents, and horizontal flow called wind. The atmosphere tries to maintain an equal pressure over the entire earth, just as the ocean tends to maintain a constant level. When the equilibrium is disturbed, air begins to flow from areas of higher pressure to areas of lower pressure. The factor that upsets the normal equilibrium is the uneven heating of the earth by the sun.

## 5.0 Thermodynamic Experiments

### 5.1 Background

Heat can best be visualized in terms of what it does. When water is heated, what is happening to the water molecules? The warmer molecules are moving at a greater velocity and are colliding with their neighboring molecules. Energy is being transferred from the heat source to the water molecules, causing them to move faster and collide with other water molecules more frequently. The rate of evaporation of the water molecules increases as a sample warms, and the vapor pressure of the sample (determined by the number of water molecules in gaseous form above the liquid) also increases. Therefore, visualizing the effects of heat at the molecular level makes it easier to comprehend.

Since frozen and liquid water have different physical properties, their models must be different to. Water molecules in the liquid form can be visualized as being extremely small V-shaped objects that are packed closely together. These molecules are moving extremely rapidly in all directions and are colliding with great frequency. In changing from a liquid to a solid (as in freezing), much of the molecule's energy is lost as heat is removed, causing the movement and frequency of collisions to diminish. In ice, the intermolecular forces (forces between molecules) of attraction are strong enough to hold the water molecules relatively immobile. The motion of molecules should cease if the temperature of the ice were reduced to a temperature known as absolute zero (-273.15°C or 0 K).

When heat is applied to ice, part of the heat goes toward breaking the bonds that hold the water molecules together as a solid. The heat of fusion, which is the amount of energy required to melt one gram of solid, is another identifying characteristic value for each compound. The objective of the second portion of this module is to determine the amount of heat required to melt one gram of ice.

Density and specific heat are also frequently used to identify unknown liquids and solids. As mentioned previously, the specific heat of a substance is the amount of heat (calories) required to raise 1 g of the substance by 1°C in temperature. Density is a measure of the mass per unit volume for a substance or object (measured in g/cm<sup>3</sup> for instance).

#### A Note on Energy Units

The calorie is a fundamental and easily understood unit. However, there are many other units of heat, and you will need to be able to convert to and use these other units in laboratory and classroom discussions. The unit of heat most commonly used is the joule, which is equal to 0.2390 calories. This is more commonly stated and remembered with the equation

$$1.000 \text{ cal} = 4.184 \text{ joules.}$$

A joule and a calorie are both measures of heat energy. The only difference between the two is the measurement system you are using (*i.e.*, metric or British). The Joule is a unit in the International System (or metric system).

## 5.2 Safety Precautions

A chemistry laboratory is a safe place to work if you follow appropriate safety procedures. The most important piece of safety equipment is eye protection. The following safety items are mandated:

- Wear eye protection at all times!
- DO NOT eat, drink, chew, or smoke in the lab.
- DO NOT wear shorts or short skirts in the lab.
- DO NOT wear open-toe sandals or high heels in the lab.
- Clean up spills and accidents immediately and correctly. If you're not sure how, ask.
- NEVER pipet anything by mouth; use a rubber bulb.
- Read the label on the reagent bottles each time before you use them.
- Dispose of waste in proper containers; when in doubt, ask!
- Assume all chemicals are toxic and possibly corrosive. Read the label on reagent bottles and on experimental containers each time before you use them (afterward may be too late). Some chemicals are toxic; some are corrosive; some are flammable. If you don't know anything about a compound's properties, it's best to assume it might be all three!
- Know where the safety shower, fire extinguisher, and emergency eye wash stations are and how to use them.
- Flush acids and bases from the skin with lots of water.

Be your brother's keeper. Watch out for others' safety as well as your own. Point test tubes being heated away from yourself and your neighbors. Wait until your glassware is cool before you ask anyone to pick it up. If you see anything that happens to be unsafe, report the situation to the T.A.

Bunsen Burner - It is sometimes difficult to see at a glance whether or not your Bunsen burner is operating. Always exercise caution when working around a Bunsen burner by making sure your cloths, hair, and papers are clear from the flame area.

### 5.3 Materials Required

#### Experiment 1

Equipment:	Styrofoam cup or insulated plastic mug	Metal Cup
	Ring Stand	Clay Triangle
	Candle	Bunsen Burner or hot plate
	Ice	Thermometer ( for thermistor calibration)
	Thermistor	

#### Experiment 2

Equipment:	Two Styrofoam cups or insulated plastic mugs	
	Pasteur pipet with bulb	Ice
	Plastic gloves	Thermistor
	Thermometer (for thermistor calibration)	

Reagents: Antifreeze (or ethylene glycol)  
*t*-Butanol (*tert*-butyl alcohol)

## **Experiment 1**

### **Heat Associated with Chemical Change**

In this experiment, you will determine the heat of combustion for candle wax. This combustion results in a chemical change. During the experiment, some wax will melt as well as combust. To avoid losing the mass of wax drops, it is a good idea to weigh the candle with a paper towel. Put the paper towel under the candle while it is burning to catch any drips. At the end of the experiment, weigh the candle and the towel. In this way, you account for the mass of any unburned wax that has dripped from the candle.

Step 1 - Set up a ring stand with a clay triangle to hold the aluminum can. The can will serve as a water container. Heat transfer from the flame to the can is facilitated by the use of a clay triangle, as opposed to a wire mesh (see figure below).

Step 2 - Clean as much carbon off the metal can as possible, then find its mass using a balance. Place ~100 mL of cold water in the can and determine the mass of the metal can with the water. The temperature of the water need not be any particular value, but the temperature should be correctly measured.

Step 3 - Use a balance to accurately determine the mass of the candle (if you choose to use a paper towel, include its mass in the measurement). By means of a clamp, position the candle with the wick about 1 inch below the cup.

Step 4 - Connect the thermistor to the LabWorks™ interface. Calibrate the thermistor if it has not been calibrated already.

Step 5 - Start your program for measuring temperature or start a pre-written program (THERM2, for instance).

Step 6 - Record the initial temperature of the water.

Step 7 - Light the candle. Stir the water gently with the thermistor until a 20°C temperature rise is noted.

Step 8 - Blow out the candle. Allow the molten wax to solidify so that drops are not lost when removing the candle from its holder

Step 9 - Record the final temperature of the water.

Step 10 - Find the mass of the candle and towel to determine the grams of candle burned. You may also need to remove any wax drippings from the candle holder and include them when measuring the mass.

Step 11 - Data Analysis - Perform the following calculations on your Data Summary Sheet:

a) The specific heat of water is 1.00 cal/g/°C. The equation which allows you to calculate the amount of heat absorbed by water is:

$$q = s \times g \times \Delta T$$

where  $q$  is the number of calories absorbed by water,  $s$  is the specific heat (for water,  $s$  is 1.00 cal/g/°C,  $g$  is the mass of the water, and  $\Delta T$  is the change in temperature (e.g., 20°C). Calculate the amount of heat,  $q$ , absorbed by the water in your experiment. Report this value on your Data Summary Sheet.

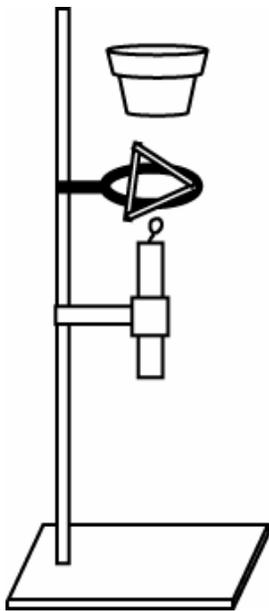
b) The value of  $q$  also represents the amount of heat given off by the combustion of the candle, assuming all the heat given off was absorbed by the cup of water. Actually, there would be some heat loss to the atmosphere, clamp, etc. The amount is relatively small and may be disregarded in this experiment.

c) The amount of heat liberated per gram of candle burned ( $H$ ) can be calculated by the following expression:

$$H = q / G$$

Where  $G$  refers to the grams of candle burned.  $H$  is a measure of the heating ability of the candle and is also a measure of the energy change involved in a chemical change (the reaction of the candle wax with oxygen from the air). Calculate  $H$  for your candle. Report this value on you Data Summary Sheet.

### Experimental Setup:



# Experiment 1

## Data Summary Sheet

Name \_\_\_\_\_

Date \_\_\_\_\_

Lab Section \_\_\_\_\_

Mass of can and water \_\_\_\_\_ g

Mass of empty can \_\_\_\_\_ g

Mass of water \_\_\_\_\_ g

Temperature before heating \_\_\_\_\_ °C

Temperature after heating \_\_\_\_\_ °C

Temperature change \_\_\_\_\_ °C

Mass of candle before burning \_\_\_\_\_ g

Mass of candle after burning \_\_\_\_\_ g

Mass of candle burned \_\_\_\_\_ g

Calculate calories of heat absorbed by water:

Calculate the calories of heat produced per gram of candle burned:

## Experiment 2

### Chemistry of Antifreeze: Density and Freezing Point

The objective of this experiment is to determine the freezing points of both *t*-butanol and a solution of automobile antifreeze in *t*-butanol. In addition, you will determine the densities of both the solution and the pure solvent using the LabWorks™ Interface system.

Most of you know that in northern climates you cannot simply put water in your car's radiator - it will freeze in the winter and damage or destroy the engine. Antifreeze must be mixed with the water to prevent this from happening. If you look at the ingredients list on any antifreeze container, you will see that it is more than 95% ethylene glycol.

The fluid in your radiator is therefore a homogeneous liquid mixture: it is "homogeneous" because it looks the same (even under a microscope) and a "mixture" because it contains two components which can be separated (if you boil the mixture, the water will evaporate leaving the ethylene glycol behind). This kind of mixture is very important in chemistry; it is called a solution.

In a solution there are always two or more components, one of which there is a lot of and very little of the other. Think of the ocean; it has some salt, but mostly water. Cola drinks are a mixture of several different components with water, yet they are still mostly water. The major component is called the solvent, and the minor one the solute.

There are several ways of describing the composition of such a mixture. In this experiment, we will use % composition as the measure. % Composition is defined as:

$$\% \text{Composition} = \frac{\text{Weight of Component}}{\text{Total Weight of Solution}} \times 100$$

For example, if we take 5.00 g ethylene glycol and mix it with 95.00 g water, this gives a total weight of 5.00 g + 95.00 g = 100.00 g and a percent composition of glycol:

$$\frac{5.00\text{g}}{100.00\text{g}} \times 100 = 5.00\%$$

To make such calculations in this experiment, we will use densities, and make some assumptions about how big the drops of antifreeze are.

The solution in your car radiator freezes at much lower temperature than pure water does. This is a general property of solutions: any solution freezes at a lower temperature than the pure solvent. It would be nice if we could do this experiment with water, but to do so we would need access to temperatures down to -30°C. As we do not have the equipment to do this, we will use *t*-butanol in its place as it has a higher freezing point.

You will begin by determining the freezing point of *t*-butanol. Then each group will be assigned a specific mixture of ethylene glycol and *t*-butanol for measuring the change in the freezing point. You will mix this solution, and determine the density and freezing point. At the end of the class period, the density and freezing point data will be collected and correlated for the whole class. It is important to work slowly and carefully; sloppy work usually shows when all the class results are compared.

### **Part 1. Density and Freezing Point of Pure *tert*-Butanol.**

Step 1 - Connect the thermistor to the LabWorks™ interface with one lead on “DAC1” and the other lead on “I1”.

Step 2 - Calibrate the thermistor, if needed. You may use ice water and warm tap water (about 40-50°C). The thermistor should either be calibrated or checked prior to use to ensure accurate readings.

Step 3 - Put on a pair of plastic gloves, and pour about 10 mL of *t*-butanol from the stock bottle into a small beaker. Use plastic gloves whenever you handle *t*-butanol or the ethylene glycol. Spills on the skin are usually harmless but you may have an allergic reaction.

Step 4 - Check to see that your 10 mL graduated cylinder is clean and dry. If it is not, wash it out with soap and water and dry it carefully with a wipe. Now carry out the following sequence of operations to determine the density of your solution, recording the raw data in your notebook.

- a. Weigh the empty graduated cylinder and record its mass.
- b. Transfer about 1.5 mL of *t*-butanol from the beaker to the graduated cylinder and record the volume (read the bottom of the meniscus accurately to within two decimal places).
- c. Weigh the graduated cylinder and its contents and record the mass.
- d. Repeat steps b and c at least 4 more times. The volume and the mass should increase in each step; at the end of the process, you should have ~10 mL in the graduated cylinder.

This whole operation should take no more than ~5 minutes. Do it at the balance, but be careful not to spill any *t*-butanol. If you do have a spill, clean it up immediately with a wipe. Leave the *t*-butanol in the graduated cylinder and take it back to your station. Keep the extra *t*-butanol you have in the beaker as it is needed in the next step.

Step 5 - Fill a clean, dry 10 cm disposable test tube half full with *t*-butanol from the graduated cylinder. Dry your thermistor off carefully with a wipe, place it into the liquid in the test tube and put the test tube into a beaker of hot water to warm it well above the freezing point. Fill the second beaker with ice, and add just enough water to make the mixture slushy.

Step 6 - Now you are ready to start your LabWorks™ Interface program; (BUTANOL.EXP is a pre-written program for this experiment). You can call the output file BUTANOL1.DAT, for example). Once you can see the graph plotting ( $T$  should be 40 °C or higher) insert the test tube and its contents into the ice bath. You are collecting data for a cooling curve, and you are also sending it to a spreadsheet so you can have a permanent copy.

This experiment is very sensitive to external influences, and the *t*-butanol needs to be stirred vigorously as the cooling proceeds. Hold the test tube with one hand, and use the thermistor to stir the liquid with a circular motion as the cooling curve is plotted. Continue to stir and watch both the liquid and the graph until you are down to about 10°C. Be careful to keep all of the liquid submerged in the ice bath, and do not let the thermistor tip rest on the bottom of the test tube. The temperature will drop, but there will be some parts that are not smooth. Notice what the curve does when you see crystals beginning to form in the test tube; be prepared to describe what you see.

## **Part 2. Density and Freezing Point Determination for a Solution of Ethylene Glycol in *t*-Butanol.**

Step 7 - You are now ready to determine the freezing point of a solution in *t*-butanol. Pour 10 mL of *t*-butanol into the beaker (if you need to, melt the sample you just used in step 6 in warm water). Put a few mL of ethylene glycol from the stock bottle into another clean and dry test tube. Your T.A. will assign you a specific number of drops (between 5 and 30) to add for your measurement of the change in the freezing point.

The amount of ethylene glycol added will vary across the class so you can see what happens to the freezing point over a range of densities. Just 1 or 2 mL will be plenty. Measure out your assigned number of drops using a Pasteur pipet (everyone needs to use the same kind of dropper so that the results can be compared). Add the ethylene glycol to the *t*-butanol and mix thoroughly with a stirring rod.

Step 8 - Now repeat steps 4-6 to find the density and freezing point of your solution (be sure to name your output data file differently than before). You will find that your solution freezes at a lower temperature than pure *t*-butanol; how much lower, and how this relates to the density, are the questions that we need to answer. When you have determined the freezing point and the density of your assigned solution, write your results in a table. A copy of this table will need to be included in your laboratory write-up.

Step 9 - Before performing the data analysis, you should investigate the sensitivity of the experiment. Follow steps 4-6 to obtain two more cooling curves with pure *t*-butanol, one in which you stir the solution by moving the thermistor gently up and down while the cooling curve is collected, and one in which you do not stir with the thermistor at all. You can use the same program (BUTANOL.EXP) to collect these curves. Be sure you give the data files different names so that you can refer to all of the data when writing your report.

## Data Analysis

Step 10 - You should now have four cooling curve data files on disk; one each for pure *t*-butanol, the solution with vigorous stirring, moderate stirring, and no stirring at all. Make graphs of your cooling curves using any software available to you. You need to determine the freezing point temperature to within 0.1°C; this may be difficult from the raw graph. To increase the readability of the graph, scale the data to within 1°C above and below what you think is the freezing point which should make it possible to read your freezing point to within  $\pm 0.1^\circ\text{C}$ .

Step 11 - Repeat step 10 to determine your freezing point for the curves for the vigorously stirred solvent and solution. These should have a recognizable feature which you can use to measure the freezing temperature accurately. Then compare the vigorously stirred pure *t*-butanol with the others that were not so vigorously stirred. Describe the differences between them in a short paragraph in your report. The explanation for these differences is not very complex. Suggest some possibilities yourself.

Step 12 - The calculation of the density requires the use of a spreadsheet. Using your program of choice, label two columns: MASS and VOLUME. Now enter your measurements of mass and volume for each of the additions of liquid from step 4.

Step 13 - The entries in the MASS column should be the mass of both the solvent and the mass of the graduated cylinder (see step 4 if you forgot about this). To obtain the mass of the solvent, the mass of the graduated cylinder needs to be subtracted from each measurement. To do this, you can simply add to each cell “-##” where ## is the mass of the empty, dry graduated cylinder.

Step 14 - To get the “best” value for the density of pure *t*-butanol, graph the volume on the *x*-axis and the mass (calculated in step 13) on the *y*-axis. Now calculate the best fit line to the data; if you’re using a spreadsheet, you can graph the data while choosing “linear regression” as the type of graph to obtain these data. The parameters of the best fit line will be in the form of  $y = mx + b$ ; further, you should calculate the correlation coefficient or “goodness of the fit” (sometimes referred to as  $R^2$ ) A correlation coefficient of 1.0 represents a perfect fit. Which of the parameters in the regression is the density that you should report?

Step 15 - Save this file to a flash drive, give it a name such as “PURE”.

Step 16 - Repeat steps 12-15 for the *t*-butanol and ethylene glycol solutions.

Step 17 - Cleanup. The T.A. will tell you where to dispose all the chemicals.

At the end of the lab period, your data will be put together and used to discuss precision and accuracy in the laboratory measurements. Prepare a table which includes the densities and freezing points for both pure *t*-butanol and for the solution. These will be compared across the whole section.

## Experiments 3, 4, and 5

### Projects for Further Investigation

Each of the following projects defines a problem. Your assignment for each problem is to design an experiment (or several experiments) that might be useful in solving the problem or explaining the phenomenon. Select a project that interests you the most.

Experiment 3. Boiling-Point Elevation by Dissolved Substances. When solutes are dissolved in solvents, the boiling point of the resultant mixture is different than the pure solvent. Use pentane and diethyl ether as solvents to investigate this phenomenon. Use camphor (moth repellent) as the solute and be careful not to expose yourself to fumes.

Experiment 4. Factors Affecting Freezing-Point Depression in Water. You looked at the effect of antifreeze on the freezing point of *t*-butanol in Experiment 2. Use water as the solvent this time, and explore the effects of antifreeze, sugar, and salt. Use a mixture of ice and salt as your cooling bath.

Experiment 5. Molar Heats of Reaction for Hydrochloric Acid (HCl) and Sodium Hydroxide (NaOH). Measure the heat of reaction with 1 mole of aqueous HCl with one mole of aqueous NaOH. Find out if the heat you measure is concentration dependent. Note: you do not need to react one full mole of each substance to calculate their molar heats of reaction.

#### Individual Report Requirements

Reports should be submitted on letter sized paper and must be typed. You and your lab partner will submit separate reports, although we know that the data will be identical.

As part of this project, you and your lab partner will be required to give a 5-10 minute presentation to your section. You and your lab partner may divide the tasks any way you wish, but there must be a clear level of equal participation.

Include the following sections in your report. Grammar, spelling and presentation all count.

**A. Introduction / Purpose** – Objectives/background briefly given in your own words.

**B. Experimental Measurements and Data** – Principles of experiment, apparatus, all data you used in calculations, discussions, and conclusions. Large amounts of related data should be presented in tabular form. Include labeled graphs and notes made on the experiment.

**C. Data Analysis / Calculations** – Explain how data was analyzed including all formulas and calculations you used. Discuss relation of analysis to theory and experiment. If a calculation was repeated several times, the multiple results should be shown in tabular form with a single example of the calculation shown once. Include an error analysis.

**D. Discussion / Conclusions** - Briefly and clearly state your interpretations of the results and mention any problems you had in doing the experiment. If you didn't reach any conclusions from the experiment, you should state why and support your explanation with data.

**E. References** - Document any outside primary references (**not** the internet!) you used in preparing your report. You should use text books, and perhaps journals, as references. You do not need to cite your lab partner or others in your lab (including your T.A.) as references.

## 6.0 Gas Law Experiments

### 6.1 Background

In several of these experiments, you will not need the LabWorks™ interface. While the interface can be used to monitor the temperature and pressure, it cannot be used to directly monitor volume.

### 6.2 Safety Precautions

A chemistry laboratory is a safe place to work if you follow appropriate safety procedures. The most important piece of safety equipment is eye protection. The following safety items are mandated:

- Wear eye protection at all times!
- DO NOT eat, drink, chew, or smoke in the lab.
- DO NOT wear shorts or short skirts in the lab.
- DO NOT wear open-toe sandals or high heels in the lab.

### 6.3 Materials Required

#### Experiment 6

Equipment:	400 mL beaker	crucible tongs
	125 mL Erlenmeyer flask with rubber stopper	Bunsen burner or hot plate
	large beaker or battery jar	thermistor
	100 mL graduated cylinder	thermometer (for thermistor calibration)
	ice	

#### Experiment 7

Equipment: TI Graphing Calculator  
Vernier Gas Pressure Sensor with 20 mL syringe  
LabPro Interface  
DataMate program

## Experiment 6

### Charles' Law

In this experiment, you will explore the effect of temperature on a gas which remains at constant pressure. You will not use the LabWorks™ interface to make measurements for this experiment.

The dependence of the volume of a gas on its temperature was first described by Jacques Charles in 1787. He was among the first people in history to fly in a hot air balloon and was, therefore, interested in the relationship between the volume of a gas and its temperature. The relationship is that the ratio of volume to the absolute temperature is a constant if the pressure remains unchanged. Mathematically, this is generally expressed as:

$$(1) \quad \frac{V}{T} = k$$

or

$$(2) \quad \frac{V_1}{T_1} = \frac{V_2}{T_2}$$

#### Procedure

Step 1 - Place about 200 mL of water in a 400 mL beaker. Place the beaker on a hot plate or over a Bunsen burner and heat it to the boiling point.

Step 2 - While the water is heating, prepare a cold water bath by putting cold water in a heavy container (e.g. battery jar or a large beaker) and adding ice. The water level for the cold water bath should be within about 3 cm of the rim.

Step 3 - Thoroughly dry a 125 mL Erlenmeyer flask and fit it with a rubber stopper with a single hole in it.

Step 4 - Using crucible tongs, immerse the flask in the boiling water up to its neck. Keep the hole in the stopper open to the atmosphere and do not allow water to enter the flask during the heating process. Hold the flask in the boiling water for at least six minutes.

Step 5 - Record the temperature of the boiling water with a thermometer.

Step 6 - Using the tongs, quickly invert the flask and insert it into the cold water bath. Keep the stoppered end pointed downward and immerse the flask as completely as possible for five minutes. Water will be drawn into the flask.

Step 7 - When the flask is cool, adjust its position (still upside down) so the level of the water inside the flask is the same as that in the cold water bath. While the flask is in this position, place your finger over the hole in the stopper and remove the flask, placing it upright on the lab table.

Step 8 - Record the temperature of the cold water bath, using the thermistor or thermometer.

Step 9 - Pour the water from the flask into a 100 mL graduated cylinder and determine the volume of the water that was drawn into the flask while it was under water.

Step 10 - Determine the volume of the full flask with the stopper in place. Fill the flask with water, insert the stopper as it was initially, and pour the water into the graduated cylinder to determine the amount of water in the full flask.

### **Calculations**

We know that the volume decreases with decreasing temperature. We would like to quantify this relationship.

Step 11 - Determine the ratio of the volume of the air to its absolute (Kelvin) temperature when it was hot. The volume is the same as that of the full flask and the temperature is the same as the hot water temperature.

Step 12 - Determine the ratio of the volume of the air to its absolute (Kelvin) temperature when it was cold. The volume of the cold air is the volume of the full flask minus the volume of the water that was drawn in when the flask was upside down, and the temperature is the same as the cold water temperature.

Is there a relationship between these two ratios? What does this tell you about the dependence of the volume on its absolute temperature?

## Experiment 6 Data Summary Sheet

Name \_\_\_\_\_

Date \_\_\_\_\_

Lab Section \_\_\_\_\_

Temperature of the hot water bath \_\_\_\_\_ °C

Volume of water in the flask after cooling \_\_\_\_\_ mL

Temperature of the cold water bath \_\_\_\_\_ °C

Volume of the full flask \_\_\_\_\_ mL

Absolute temperature of the hot water bath \_\_\_\_\_ K

Absolute temperature of the cold water bath \_\_\_\_\_ K

Volume of cold air \_\_\_\_\_ mL

Volume / temperature ratio for hot air \_\_\_\_\_ mL/K

Volume / temperature ratio for cold air \_\_\_\_\_ mL/K

1. What is the mathematical expression for Charles' Law?
  
2. A sample of gas had a volume of 265 mL when its temperature was 298 K. If the pressure remains constant, what will be its volume when the temperature decreases to 225 K?
  
3. A sample of argon had a volume of 19.78 L at 25°C. At what temperature would its volume be 24.0 L? (Assume constant pressure).

## Experiment 7

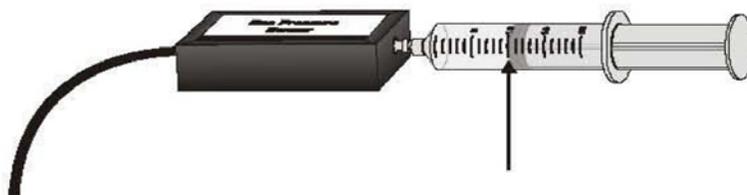
### Boyle's Law

The LabPro interface has an “Events with Entry” mode which collects one data point each time you press [ENTER] and then asks you to enter a corresponding value for that point. This tutorial will guide you through the setup procedure of an experiment using the Events with Entry to investigate the relationship between the pressure and volume of a gas.

#### Procedure

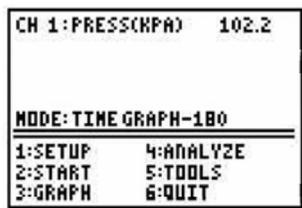
Step 1 - Prepare the Gas Pressure Sensor and an air sample for data collection.

- Plug the Gas Pressure Sensor into Channel 1 of the LabPro. Use the link cable to connect the TI graphing calculator to the interface. Firmly press in the cable ends.
- With the 20 mL syringe disconnected from the Gas Pressure Sensor, move the piston of the syringe until the front edge of the inside black ring (indicated by the arrow) is positioned at the 10 mL mark.



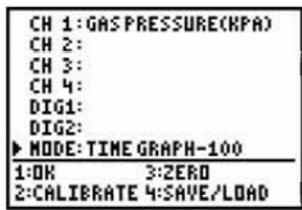
- Attach the 20 mL syringe to the valve of the Gas Pressure Sensor. The Vernier Gas Pressure Sensor has a white stem protruding from the end of the sensor box - attach the syringe directly to the white stem with a gentle half-turn.

Step 2 - Turn on the calculator and start the DataMate program. Notice how the program automatically identified the Pressure Sensor (an auto-ID sensor). It displays the channel it is connected to and the current reading. Press [CLEAR] to reset the program.



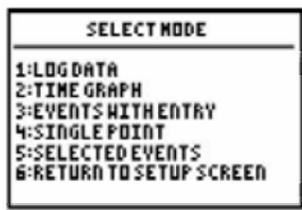
Step 3 - To collect pressure vs. volume data for Boyle's law, you need to collect the data in the Events with Entry mode. The current mode is time graph, so select SETUP.

Step 4 - Below is the Setup screen. Here you can change the data collection mode. To select MODE, press up once and press [ENTER].



```
CH 1: GAS PRESSURE(KPA)
CH 2:
CH 3:
CH 4:
DIG1:
DIG2:
▶ MODE: TIME GRAPH-100
1:OK          3:ZERO
2:CALIBRATE  4:SAVE/LOAD
```

Step 5 - A list of modes will appear in the SELECT MODE screen. Press the number preceding EVENTS WITH ENTRY. You will be returned to the Setup screen. Then select OK to return to the setup screen.



```
SELECT MODE
1:LOG DATA
2:TIME GRAPH
3:EVENTS WITH ENTRY
4:SINGLE POINT
5:SELECTED EVENTS
6:RETURN TO SETUP SCREEN
```

Step 6 - Select OK to return to the Main screen. You are now ready to collect pressure and volume data. It is best for one person to take care of the gas syringe and for another to operate the calculator.

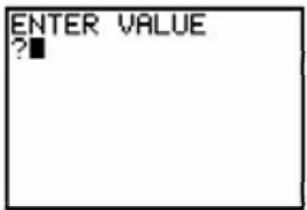
Step 7 - Select START. Notice that a graph is not initially displayed; instead, it is going to wait for you to collect at least two data points.

Step 8 - Move the piston so the front edge of the inside black ring is positioned at the 5 mL line on the syringe. Hold the piston firmly in this position until the pressure value displayed on the calculator screen stabilizes (in this example, the number to the right, 194.32 kPa). The number to the left indicates the sample number.



```
PRESS [ENTER] TO COLLECT
OR [STOP] TO STOP
1    194.32
```

Step 9 - Press [ENTER], which tells the DataMate to record the pressure. Type in “5”, the gas volume (in mL) on the calculator. Press [ENTER] to store this pressure-volume data pair. To collect another data pair, move the syringe to 7.5 mL. When the pressure reading stabilizes, press [ENTER] and enter “7.5” as the volume. Repeat the above procedure for volumes of 10.0, 12.5, 15.0, 17.5, and 20.0.

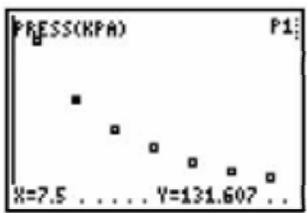


As you work your way through the experiment, you get to view the data being collected. Notice the sample number (5), the value given (68.9 kPa) and the graph in progress.



Step 10 - When you are done, press [STO] to end data collection.

Step 11 - Move the cursor right or left (using the arrow keys) to examine the data pairs on the displayed graph.



Step 12 - Press [ENTER] to return to the Main menu, then select QUIT to quit the program.

## Experiments 8 and 9

### Projects for Further Investigation

Each of the following projects defines a problem. Your assignment for each problem is to design an experiment (or several experiments) that might be useful in solving the problem or explaining the phenomenon. Select a project that interests you the most.

Experiment 8. Adiabatic Gas Expansion. When compressed gases expand, they must perform work to overcome the attractive forces between molecules. Design and test an experimental setup to determine the amount of work (as heat) developed by such an expansion.

Experiment 9. Boiling Point Elevation Due to Pressure. When a liquid boils, its vapor pressure has reached the pressure of the atmosphere. By increasing the pressure surrounding the liquid, the temperature required to reach boiling increases. Design and test an experimental setup which allows you to determine the boiling point for water at several pressures and compare these values to a phase diagram.

#### **Individual Report Requirements**

Reports should be submitted on letter sized paper and must be typed. You and your lab partner will submit separate reports, although we know that the data will be identical.

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**C. Data Analysis / Calculations** – Explain how data was analyzed including all formulas and calculations you used. Discuss relation of analysis to theory and experiment. If a calculation was repeated several times, the multiple results should be shown in tabular form with a single example of the calculation shown once. Include an error analysis.

**D. Discussion / Conclusions** - Briefly and clearly state your interpretations of the results and mention any problems you had in doing the experiment. If you didn't reach any conclusions from the experiment, you should state why and support your explanation with data.

**E. References** - Document any outside primary references (**not** the internet!) you used in preparing your report. You should use text books, and perhaps journals, as references. You do not need to cite your lab partner or others in your lab (including your T.A.) as references.