CHEMISTRY MODULE 1
AN INTRODUCTION TO THERMODYNAMICS
Heat, Calories, and Specific Heat

MODULE PURPOSE — After completing this module students will be able to:

1. Define the concepts of calories and specific heat.
2. Compare the energy changes involved in chemical and physical changes.

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1.0 Background

Almost all chemical reactions either adsorb or release heat. These experiments, although simple ones, 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 amount 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 it either 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 1 g 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.

Many chemical reactions spontaneously give off or absorb heat as they occur. Scientists, in fact, are often more interested in the heat liberated in reactions than in the resulting product. This heat can be used directly or can be converted into useful forms of work or energy.

2.0 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, CH₄. The combustion of methane results in the following reaction:

\[
\text{CH}_4 + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O} + 212.8 \text{ kcal}
\]

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). One kcal of energy is equal to one food Calorie or 1000 heat calories.

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, transmission of nerve impulses, and synthesis of essen-
Thermodynamics: Heat, calories, and specific heat

tial compounds. For example, when glucose (a type of sugar) is used in metabolism, the following oxidation reaction occurs:

\[
\text{C}_6\text{H}_{12}\text{O}_6 + 6\text{O}_2 \rightarrow 6\text{CO}_2 + 6\text{H}_2\text{O} + \text{688 kcal}
\]

or

\[
\text{glucose} + \text{oxygen} \rightarrow \text{carbon dioxide} + \text{water} + \text{heat}
\]

We often refer to the number of calories in a given portion of food. Accordingly, we would say there are 688 kcal of energy in 180 g (one mole) of glucose; meaning when this amount of glucose is oxidized in metabolism, there are 688 kcal of energy liberated. The minimum daily requirement for an adult human is between 1800 to 3000 kcal depending upon the individual’s size and activity level.

If our diet provides more energy than our bodies can use, the surplus may be stored as lipids (fats), resulting in excess body weight. Approximately 3500 kcal equals one pound of body weight. If our diet does not provide us with enough energy, we must either use some stored fat or suffer a lack of energy.

3.0 Thermodynamic Experiments

3.1 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 too. Water molecules in the liquid form can be visualized as 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 molecules' energy is lost as heat is removed, causing the movement and frequency of collision 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 the molecules should cease if the temperature of the ice were reduced to a temperature known as absolute zero (−273.15°C or –459.67°F).

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/cm³, for instance).

You can find the specific heat of an unknown metal using principles presented in experiment 1-2, Heat Loss & Heat Gain. The density of the metal can be determined by its mass and the volume displacement of water in a graduated cylinder. You can then refer to a table of known physical properties for specific heat and density.
to identify your unknown metal.

**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. British or metric). The Joule is a unit in the International System (metric system).

### 3.2 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 recommended:

- **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-toed sandals, thongs, 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 reagent bottles each time before you use them.
- Dispose of wastes 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 (Afterwards 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 is best to assume it might be all three!
- Know where the safety shower, fire extinguisher and emergency eye wash stations are how to use them.
- Flush acids and bases from 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 appears unsafe, report the situation to your instructor.

**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 clothes, hair, and papers are clear from the flame area.

### 3.3 Materials Required

**Experiment 1-1**

**Equipment:**
- Styrofoam coffee cup or insulated plastic coffee mug
- 6-inch test tube
- Filter paper
- Boiling chip
- Thermistor
- Thermometer (for thermistor calibration)
Reagents: Pentane  
Acetic Acid (1 M)  
Ethyl Alcohol  
Isopropyl Alcohol

Experiments 1-2 through 1-5

Equipment: Thermometer (for thermistor calibration)  
Styrofoam coffee cup or insulated plastic coffee mug  
Metal cup  
Ring stand  
Clay triangle  
Candle  
Ice  
Thermistor  
Bunsen burner or hot plate

Experiments 1-5 through 1-6

Equipment: Two styrofoam coffee cups or insulated plastic coffee mugs  
400 ml beaker  
Two thermistors  
Thermometer (for thermistor calibration)

Unknown: Obtain unknown metals per directions from your lab instructor.  
**Instructor Note:** You will need each unknown in both slugs and in small pieces (like shot). Hook a small wire or stirrup to the slugs, using wire of the same material as the unknown, if possible.

Experiment 1-7

Equipment: Two styrofoam coffee cups or insulated plastic coffee mug  
Stirring rod  
Thermistors  
Thermometer (for thermistor calibration)  
4-inch or 10 cm test tubes  
Labeling tape  
Spatula

Reagents: Potassium nitrate  
Calcium Chloride  
Sodium acetate  
Ammonium thiocyanate

Experiments 1-8

Equipment: Two styrofoam coffee cups  
Antifreeze or ethylene glycol  
t-butanol  
Pasteur pipette with bulb  
Ice  
Plastic gloves  
Thermistors  
Thermometer (for thermistor calibration)
### 3.4 Experimental Procedures

The following calibration procedure and experiments are provided:

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Thermistor Calibration

A thermistor is made of a material known as a semiconductor in which almost all of the electrons are bound closely to their parent electrons and will not move. Application of heat to the thermistor promotes some electrons to a conduction band, which is shared among the atoms in the semiconductor crystal, and in which electrons are free to move from atom to atom. Thus the number of mobile electrons increases as the temperature of the semiconductor increases. If a small voltage is applied to one side of the semiconductor, the number of electrons forced out of the other side will be related to the temperature of the material. We can measure this movement of electrons (electrical current). If we have accurately measured the current for at least two known temperatures, the LabWorks™ interface and software can compute and display the temperature of the thermistor.

Because the Interface can determine a value to one part in about 8000, it is possible to use a thermistor to resolve changes of about 0.01 degree Celsius in the room temperature range. The sensitivity of the thermistor decreases at higher temperatures.

To be of value, the thermistor must be calibrated. To calibrate a sensor, we place it in a known condition and record the current or voltage produced. By repeating this two or more times, we can create a graph that relates the sensor output to the measure of value. If the graph forms a smooth line as shown in Figure 1-1, the computer can develop an equation that relates sensor output to the measured value. By solving this equation, the computer can calculate the true value that corresponds to the measured voltage or current. It can then display this value.

Step 1 — Connect the Thermistor Sensor leads to DAC1 and I1 if you are using a single thermistor. See page R1-6 for an illustration of this connection, if needed. If you are using two thermistors, connect the leads of the second thermistor to DAC2 and I2.

Step 2 — Select the CALIBRATE SENSORS function from the LabWorks™ main menu. You will immediately see a sub-menu asking which sensors you wish to calibrate. You can calibrate Temp1, Temp2, or both. This generalized calibration procedure is useful for any sensor that produces either current or voltage, and is either linear or logarithmic. Dissolved oxygen probes and light sensors are examples of probes that can be calibrated with this routine, in addition to temperature and pH.

Step 3 — Calibration Procedure: Place your thermistor and a laboratory thermometer into a cold solution (such as ice-water) and watch the display of current through the thermistor. Note that it gradually decreases as the thermistor cools and ultimately stabilizes when the thermistor reaches equilibrium with the solution. At this point the current should be relatively constant (fluctuating by less than 1 micro-ampere).

*The critical factor involved in calibration is that the thermistor and the laboratory thermometer MUST be at the same temperature when the calibration point is entered.*
Step 4 — Repeat step 3 with a hot solution, again waiting for both the thermistor and the thermometer to reach equilibrium with the solution. Follow the directions on screen.

Consider the problem caused in a two-point calibration if one of the points is in error. What would happen to the curve of the line illustrated in Figure 1-1 if only two calibration points were used, and one was incorrect? Your measurements will be most accurate if the calibration standards are above and below the measurements you expect to make. You may wish to use ice water for the cold solution and near boiling water for the hot solution.
Level 1  
Experiment 1-3  

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 this 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 metal cup (see Figure 1-3). The cup will serve as a water container. Heat transfer from flame to cup is facilitated by the use of a clay triangle, as opposed to a wire mesh.

Step 2 — Clean as much carbon off the metal cup as possible, then find its mass using a balance. Place about 100 ml of cold water in the cup and determine the mass of the metal cup with water. The temperature of the water need not be any certain value, but the temperature should be carefully measured.

Step 3 — Use a balance to accurately determine the mass of a 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 a 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

\[
q = s \cdot g \cdot T
\]

where \(q\) is the number of calories absorbed by the water, \(s\) is the specific heat (for
water, \( s = 1.00 \text{ cal/g}^{\circ}\text{C} \), \( g \) is the mass of the water, and \( T \) is the change in temperature (e.g. 20 \(^{\circ}\text{C} \)). 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 heat given off was absorbed by the cup of water. Actually, there would be some heat loss to the atmosphere, clamp, etc., but 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:

\[
H = \frac{q}{G},
\]

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 your Data Summary Sheet.
# Experiment 1-3

## Data Summary Sheet

### Heat Associated with Chemical Change

<table>
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<tr>
<th>Description</th>
<th>Value</th>
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<tbody>
<tr>
<td>Mass of cup and water</td>
<td>_____ g</td>
</tr>
<tr>
<td>Mass of empty cup</td>
<td>_____ g</td>
</tr>
<tr>
<td>Mass of water</td>
<td>_____ g</td>
</tr>
<tr>
<td>Temperature before heating</td>
<td>_____ °C</td>
</tr>
<tr>
<td>Temperature after heating</td>
<td>_____ °C</td>
</tr>
<tr>
<td>Temperature change</td>
<td>_____ °C</td>
</tr>
<tr>
<td>Mass of candle before burning</td>
<td>_____ g</td>
</tr>
<tr>
<td>Mass of candle after burning</td>
<td>_____ g</td>
</tr>
<tr>
<td>Mass of candle burned</td>
<td>_____ g</td>
</tr>
</tbody>
</table>

Calculate calories of heat absorbed by the water:

Calculate the calories of heat produced per gram of candle burned:
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 the northern climates you cannot simply put water in your car’s radiator — it will freeze in winter temperatures and damage or destroy the engine. Antifreeze must be mixed with the water to prevent this freezing. If you look at the ingredients list on any antifreeze jug, 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 all 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 and leave the ethylene glycol behind). This kind of mixture is very important in chemistry, and is called a solution.

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

\[
\text{Weight Component} \times 100 = \% \text{ Composition} \\
\text{Total Weight of Solution}
\]

For example, if we take 5.00 g of ethylene glycol and mix it with 95.00 g of water, this gives

\[
\text{Total weight of solution} = 5.00 \text{ g Glycol} + 95.00 \text{ g H}_2\text{O}
\]

\[
\text{5.00 g Glycol} \times 100 = 5.00 \% \text{ Glycol} \\
\text{100.00 g Total}
\]

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

The solution in your car radiator freezes at a 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. Such temperatures are difficult (expensive) to reach in most school laboratories. Instead of water we will use t-butanol, which has a higher freezing point.

You will begin by determining the freezing point of t-butanol. Then each group of students will be assigned a specific mixture of ethylene glycol and t-butanol for measuring the change in freezing point. You will mix this solution, and

**SAFETY PRECAUTIONS:** Ethylene glycol and t-butanol are both mildly poisonous, but both are safe to handle as directed in this experiment. Both of them can cause allergic reactions in some people, and you should wear plastic gloves whenever you are pouring them from one container to another. If you should spill any on your hands or anywhere else, wash it off at once with lots of water.
determine its 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 T-butanol

Step 1 — Connect the thermistor to LabWorks™ 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 the plastic gloves whenever you handle the t-butanol or the ethylene glycol. Spills on the skin are usually harmless, but some people have allergic reactions.

Step 4 — Check to see that your 10 ml graduated cylinder is clean and dry. If it is not, wash it out with soap solution and dry it carefully with a paper towel. 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 accurate to two decimal places!).

c. Weigh the graduated cylinder and its contents and record the mass.

d. Repeat steps b and c at least 4 times more. The volume and the mass will increase each time, and, at the end of the process, the graduate should contain almost 10 ml.

This whole operation should not take 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 at once with a paper towel. Leave the t-butanol in the graduated cylinder to take back to your station. Also, keep the extra t-butanol that you have in the beaker as well; you will be using both of these again in a few minutes.

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 paper towel, put it into the liquid in the test tube, and put the test tube in a beaker of hot water to warm it well above the freezing point. Fill a second beaker with ice, and add just enough water to make the mixture a little 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 that you have a permanent copy.

This experiment is very sensitive to external influences, and the 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 dry test tube. Your instructor will assign you a specific number of drops (between 5 and 30) to add for your measurement of the change in 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 in density. Just 1 or 2 ml will be plenty. Measure out your assigned number of drops using a Pasteur pipette (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 them 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 we need to answer. When you have determined the freezing point and density of your assigned solution, write your results in the table on the blackboard. 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 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, t-butanol with moderate stirring, and t-butanol with no stirring. Make graphs of your cooling curves using the Spreadsheet. You need to determine the temperature of the freezing point to about 0.1°C, and this may be difficult from the raw graph. To increase the readability of the graph, use the SCALE function to plot just a part of the graph. Assign a minimum for the Y1 axis about 1 degree below your estimated freezing point. Assign the maximum about 1 degree above. This graph should make it possible to read your freezing point to about ±0.1°C. Scaling can be very useful in many experiments.

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 temperature of freezing accurately. Then compare the vigorously stirred pure t-butanol with the others which were not so strongly stirred. Describe the differences between them in a short paragraph in your report. The explanation for these differences is not very complex. Maybe you can suggest some possibilities yourself. (Try using your textbook as a reference.)
Step 12 — The calculation of density requires the use of the spreadsheet. Start with a new or cleared spreadsheet. First, label two columns: MASS and VOLUME. Now enter your measurements of mass and volume for each of the additions of liquid.

Step 13 — The entries in the MASS column should be the mass of both the solvent and graduate. To obtain the mass of the solvent, the mass of the graduate needs to be subtracted from each measurement. The spreadsheet can do this for you. Enter a formula (function) of A - ## where A designates the column of mass entries, and ## indicates the mass of the empty graduated cylinder.

Step 14 — To get the "best" value for the density of the pure t-butanol, graph the volume on the X-axis and the mass (calculated in step 13) on the Y-axis. Choose "Linear Regression" as the type of graph in this case. The graph should look like a very good best-fit line, and the coefficient of x in the regression equation (slope) is the density. Record this density as "raw" density because it may be necessary to correct it later. The regression equation gives you the equation of the best-fit line in the form \( Y = mx + b \), as well as a correlation coefficient, which tells you something about the "goodness of fit". Correlation coefficients are often 0.99 or better (1.00 is a perfect fit), even when the data is not very good, so you need to be a little skeptical about reading this item. Linear regression is used every day in laboratories to fit data to equations, and we will make frequent use of it.

Step 15 — Save this file to disk, giving it a name, such as "PURE."

Step 16 — Repeat steps 12-15 for the t-butanol and ethylene glycol solution.

Step 16 — Clean up. Your lab instructor will tell you where to dispose of your chemicals.

At the end of the lab period, all of your data will be put together on the blackboard and used to discuss precision and accuracy in laboratory measurements. Prepare a table which includes the densities and freezing points for both the pure t-butanol and for the solution. These will be compared across the whole class.
Level 3
Experiments 1-9, 1-10, and 1-11
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 1-9  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. Don't inhale the fumes. Camphor (moth repellent) can be used as the solute.

Experiment 1-10  Factors Affecting Freezing-Point Depression in Water. You looked at the effect of antifreeze on the freezing point of t-butanol in Experiment 1-8. Use water as the solvent this time, and explore the effects of antifreeze, sugar, and salt (NaCl). Use a mixture of ice and salt as your cooling bath.

Experiment 1-11  Molar Heat of Reaction for Hydrochloric Acid (HCl) and Sodium Hydroxide (NaOH). You made several different kinds of heat measurements in Experiment 1-5. Apply those techniques to measure the heat of reaction of 1 mole of aqueous solution of HCl with one mole of aqueous solution of NaOH. Find out if the heat you measure is concentration dependent. Note: You don't have to react 1 mole each to find their molar heat of reaction.

Individual Report Requirements

Reports should be submitted on 81/2 x 11 inch paper. Typing the text portions is recommended. You and your lab partner will submit separate reports, although your data will be identical. The due date is up to your instructor.

As part of this project you will be required to give a five to ten minute presentation to the class. You and your partner may divide this up any way you wish, but you both must participate in the presentation.

The following sections should be included in your report. Grammar, spelling and neatness do count!

A. Purpose — This should be in your own words, and brief; no more than two or three sentences.

B. Experimental Measurements and Data — Include all data that you used in calculations, discussion, and conclusions. Present your data as tables wherever you can. Spreadsheet graphs and your raw notes of observations during the experiments should also be included.

C. Data Analysis/Calculations — Any formula calculations you used to analyze data must be shown here. If a calculation was repeated several times the multiple results should be shown as a table, but the calculation only needs to be shown once. In many cases a composite table with the raw data and the results all together is necessary, and your lab instructor may require you to do this.

D. Discussion/Conclusions — Briefly state your interpretations of the results and mention any problems you had in doing the experiment. Your did not make any conclusions from your experiment, correct or incorrect, explaining which support that explanation with your data.

E. References — Document any outside references you used in preparing your report. You should use books, journals, and people as references. The people in your lab and your instructor do not have to be documented as references.
MODULE PURPOSE — After completing this module, students will be able to:

1. Define the effects of temperature, pressure, and volume on gases.
2. Determine, from experimental data, the effects of pressure and temperature on a volume of gas.
4. Apply Gas Laws to real-world examples where the Gas Laws are important.

### Level 1
Basic Principles

- Pressure Sensor Calibration
- Charles’ Law—1 6-1
- Molar Volume of Hydrogen 6-2

### Level 2
Applications

- Boyle’s Law 6-3
- Charles’ Law—2 6-4

### Level 3
Further Investigation

- Adiabatic Gas Expansion 6-5
- Boiling-Point Elevation Due to Pressure 6-6

### Math Skill Development:

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1.0 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, 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 a 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 be 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.

1.1 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:

\[
V \propto \frac{1}{P}
\]

where the symbol "\( \propto \)" 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 placing a proportionality constant in the previous equation, we will have another form of Boyle's Law.

\[
V = k \left( \frac{1}{P} \right)
\]

\[
P V = k
\]

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

\[
P_1 V_1 = P_2 V_2
\]

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. A graph of this relationship for an ideal gas, shown in Figure 6-1, results in a hyperbola. As we increase the pressure, passing from state 1 to state 2, the product \( PV \) is constant.

1.2 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 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.

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

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

In general, when dealing with different kinds of gas, such as hydrogen and oxygen, it is important to use equal numbers 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 g) (1 mole/2 g) = 4 moles; therefore, it has $4 \times 6.02 \times 10^{23}$ molecules. (One mole of any substance has $6.02 \times 10^{23}$ molecules, which is Avogadro's number.) Similarly, 80 g of O$_2$ is only $80/32 = 2.5$ moles, and contains only $2.5 \times 6.02 \times 10^{23}$ molecules. All else being equal, we expect that the pressure exerted by the hydrogen would be greater by a ratio of 4 to 2.5. Thus, the following equation applies:

\[
\frac{PV}{n} = \text{constant (at constant temperature)}
\]

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

1.3 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. Figure 6-2 shows that the pressure increases proportionally 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 pressure) back to zero volume, he always found the same intercept. The intercept on the temperature axis is -273.15°C. Kelvin named this temperature absolute zero. The Kelvin absolute temperature scale, in which °K = °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 an absolute temperature scale):

\[
V \propto T
\]

(8) \hspace{1cm} V = kT

or

(9) \hspace{1cm} \frac{V_1}{T_1} = \frac{V_2}{T_2}

Note that in Figure 6-2, 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 in 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 force cause the molecules to stick together as a liquid.

The mathematical equation of the straight line in Figure 6-2 might be written as

$$P_t = P_0 \left(1 + bT\right)$$

where $P_t$ is the pressure at a temperature $T$, $P_0$ is the pressure at 0°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 $(°C)^{-1}$ for those gases, such as hydrogen, which are nearly "ideal." The pressure coefficient for the mixture of gases we breathe as air matches hydrogen's, but for carbon dioxide, it is 0.00371. The pressure coefficients for helium and oxygen are 0.00367 $(°C)^{-1}$. Such differences are minor and show that no gas is totally ideal.

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 $(°C)^{-1}$ for $b$:

$$P_t = P_0 \left(1 + 0.00366T\right)$$

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 that to make $P_t = 0$, we must have $(1 + 0.00366T) = 0$. That is, the required temperature is:

$$T = -\frac{1}{0.00366(°C)^{-1}} = -273° C$$

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 Figure 6-3. Precise experiments have given ~273.15°C as the value of this important constant, but we shall usually round the value to ~273°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.

### 1.4 Ideal Gas Equation

From Charles' law, we know that volume varies proportionally with absolute temperature. From Boyle's law, we know that volume varies inversely proportional with pressure. Combining the two laws, we obtain:

$$V = k \left(\frac{T}{P}\right), \text{ or } PV = kT$$
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*, where \( R = 8.31 \text{ dm}^3 \cdot \text{kPa} / \text{mol} \cdot \text{K} \). 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:

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

### 1.5 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 differing numbers of molecules present for each gas. Since each gas is 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.

### 2.0 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.

#### 2.1 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 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.

#### 2.2 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 into 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 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.

2.3 Effect 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 take off at an airport 5,000 ft. 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 for every 5 km increase in altitude.

2.4 Effect 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 a clearing of the weather.

2.5 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 known as wind. The atmosphere tries to maintain an equilibrium, but with the sun's heat, the oceans need 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.

3.0 Gas Law Experiments

3.1 Background

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

3.2 Safety Precautions

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

- DO NOT eat, drink, chew, or smoke in the lab.
- DO NOT wear open-toed sandals, thongs, or high heels in the lab.
- You MUST wear eye protection at all times during this experiment!
3.3 Materials Required

Experiment 6-1
Equipment:
- 400-ml beaker
- Hot plate or Bunsen burner
- 125-ml Erlenmeyer flask with rubber stopper
- Battery Jar or large beaker
- 100-ml graduated cylinder
- Crucible tongs
- Ice
- Thermistor and thermometer

Experiment 6-2
Equipment:
- Battery Jar or large beaker
- Copper wire
- Gas collection tube (buret)
- Thermometer
- Barometer

Reagents:
- 6 M HCl in plastic dispenser bottle
- Magnesium ribbon

Experiments 6-3 through 6-4
Equipment:
- Meter stick with two 3-foot sections of 5 mm glass tubing bound to the stick with rubber bands. The tubing should be connected at the bottom with a piece of tygon tubing.
- 15 to 20 cm capillary tube closed at one end, with a liquid plug (like glycerine) in it.
- Thermistor
- Hot plate
- Beakers
- Ice
- Small Cork

3.4 Experimental Procedures

The following experiments are provided:

Experiment 6-1  **Charles' Law:** Demonstrate that the ratio of volume to the absolute temperature is constant.

Experiment 6-2  **Molar Volume of Hydrogen:** Using partial pressures and the ideal gas equation, determine the volume of gas produced by a reaction of magnesium with HCl.

Experiment 6-3  **Boyle's Law:** Measure the relationship of pressure and volume at a constant temperature.

Experiment 6-4  **Charles' Law:** Investigate the relationship of gas volume and temperature.

Experiment 6-5  **Adiabatic Gas Expansion:** Determine the amount of heat energy required for a gas under pressure to expand in a system at constant temperature.

Experiment 6-6  **Boiling Point Elevation Due to Pressure:** Determine the effect of increased pressure on the boiling point of water and compare with a phase diagram.
Pressure Sensor Calibration

The SCI Technologies, Inc. Pressure Sensor is a device which is able to detect small changes in pressure. The Pressure Sensor consists of two main components; a pressure sensitive transducer and an amplifier. The sensor is capable of detecting pressures from 0 to 30 psi (0 to 207 kPa), and its output varies linearly with pressure with an accuracy of ±0.25% over the full scale.

The amplification of the pressure transducer provides a range of 2000 mV for the applied pressures from 0 to 207 kPa. Since the interface is able to accurately measure voltage signals to ±2mV, the sensitivity of the sensor is approximately 200 Pa.

Due to its linearity, the Pressure Sensor may be used to monitor pressure in terms of output voltage for qualitative measurements. However, for quantitative measurements, you must calibrate the sensor. By calibrating the sensor, you will be essentially determining the offset and slope for two different pressures and the corresponding output voltages.

To calibrate the Pressure Sensor, you will need to have a way of applying two known pressures. Three possibilities for obtaining known pressures are creating a water column, compressing a closed container of air (syringe), and measuring applied pressure with a pressure guage. Keep in mind that the LabWorks™ program will be calculating a straight line for the two known pressures you apply. Therefore, select a method which supplies pressures below and above the pressure you wish to measure. Otherwise, the program will be extrapolating the pressure data, which decreases the accuracy of your measurement.

PRESSURE SENSOR CONNECTIONS

Step 1 — Connect the Pressure Sensor output to the pH/mV 1 or pH/mV 2 input of the LabWorks™ interface with a dual plug, BNC style cable. See page R6-6 for an illustration of this connection, if needed.

Step 2 — Plug the power cable of the Pressure Sensor into the 15-pin D-sub miniature (accessory) port on the back of the LabWorks™ interface. The "Power" indicator on the front of the Pressure Sensor should be illuminated. If not, press the "Power" switch on the face of the interface.

WATER COLUMN CALIBRATION METHOD

Step 1 — Select the CALIBRATE SENSORS function from the LabWorks™ main menu. You will immediately see a sub-menu asking which sensors you wish to calibrate.

Step 2 — Select the mV1 probe or mV2 probe option.

Step 3 — Select the linear calibration procedure. The Pressure Sensor output is linear with respect to the applied pressure.
Step 4 — Fill a length of rubber tubing with water. The length of tubing should be one foot long or longer, depending on the range of pressures you wish to measure.

Step 5 — Connect one end of the tubing to the "Pressure 1" input of the Pressure Sensor. Try to leave a 2-3 inch air space in the tubing so that the pressure of the water column will not force water into the sensor. This makes the measurement of the length of the water column difficult.

Step 6 — Hold the free end of the tubing so that the water level in each end is equal. This is the first pressure calibration point.

Step 7 — Enter a value of pressure for the first calibration point. Remember that the pressure is not zero unless you wish to measure pressure relative to the atmospheric pressure. NOTE: Determine the pressure in the system of units you wish to use.

Step 8 — Raise the free end of tubing above the Pressure Sensor and measure the vertical length of the water column. This is the second calibration point. NOTE: The length of the tubing is not necessarily the same as the length of the water column unless you are holding the tubing so that it is straight and directly over the Pressure Sensor.

Step 9 — Enter a value of pressure for the second calibration point. The pressure for this point is the pressure you determined for the first calibration point plus the pressure due to the length of the water column. NOTE: Use the same system of units that you used for the first calibration point.

Step 10 — The calibration is complete. When you write or edit your LabWorks™ programs, be sure to select the mV1probe or mV2probe inputs. The mV1 and mV2 inputs are simply the value of millivolts applied to the respective inputs. The probe calibration does not affect the mV1 or mV2 inputs readings.

SYRINGE CALIBRATION METHOD

Step 1 — Select the CALIBRATE SENSORS function from the LabWorks™ main menu. You will immediately see a sub-menu asking which sensors you wish to calibrate.

Step 2 — Select the mV1probe or mV2probe option.

Step 3 — Select the linear calibration procedure. The Pressure Sensor output is linear with respect to the applied pressure.

Step 4 — Pull the plunger of a 50 cc or larger syringe back to an easily readable mark (50 cc, for instance).

Step 5 — Note the value of output voltage (displayed on the screen during calibration) and connect a short piece of plastic tubing to both the syringe and the Pressure Sensor. NOTE: Do not use latex rubber tubing, as this type of tubing will expand if the applied pressure is great enough.

Step 6 — Carefully adjust the syringe plunger to attain the output voltage which you noted before attaching the rubber tubing. This will account for any compression of air which occurred when rubber tubing was attached.

Step 7 — Enter a value of pressure for the first calibration point. Remember that the pressure is not zero. It is actually equal to the atmospheric pressure. NOTE: You may use units of
Step 8 — Push the plunger of the syringe to a volume equal to a volume of one half the original volume. This is the second calibration point. NOTE: The first volume is the volume after equalizing the pressure inside the syringe. In addition, the tubing you use to connect the syringe to the Pressure Sensor and the tubing inside the sensor housing (approximately 2 inches) should be factored into the overall volume (Volume cylinder = 2πr²h).

Step 9 — Enter a value of pressure for the second calibration point. The pressure for this point should be double the pressure you determined for the first calibration point. NOTE: Use the same system of units that you used for the first calibration point.

Step 10 — The calibration is complete. When you write or edit your LabWorks™ programs, be sure to select the mV1probe or mV2probe inputs. The mV1 and mV2 inputs are simply the value of millivolts applied to the respective inputs. The probe calibration does not affect the mV1 or mV2 inputs readings.

**PRESSURE GAUGE CALIBRATION METHOD**

Step 1 — Select the CALIBRATE SENSORS function from the LabWorks™ main menu. You will immediately see a sub-menu asking which sensors you wish to calibrate.

Step 2 — Select the mV1probe or mV2probe option.

Step 3 — Select the linear calibration procedure. The Pressure Sensor output is linear with respect to the applied pressure.

Step 4 — Connect the Pressure Sensor and pressure gauge to your source of air pressure (or vacuum) to the Pressure 1 input on the front of the sensor housing (use Pressure 2 for vacuum applications). NOTE: Your source of air pressure could be a syringe, air compressor, regulated oxygen tank, etc.

Step 5 — Depending on your type of pressure gauge, you may have to apply a small amount of pressure in order to obtain an accurate reading. Apply as small a pressure as possible.

Step 6 — Enter a value of pressure for the first calibration point. NOTE: You may use units of your pressure gauge or convert to another system of units.

Step 7 — Increase the pressure (or vacuum) applied to the Pressure Sensor. Be sure not to exceed pressures of 30 psi.

Step 9 — Enter a value of pressure for the second calibration point. NOTE: Use the same system of units that you used for the first calibration point.

Step 10 — The calibration is complete. When you write or edit your LabWorks™ programs, be sure to select the mV1probe or mV2probe inputs. The mV1 and mV2 inputs are simply the value of millivolts applied to the respective inputs. The probe calibration does not affect the mV1 or mV2 inputs readings.
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:

\[
\frac{V}{T} = k
\]

or

\[
\frac{V_1}{T_1} = \frac{V_2}{T_2}
\]

**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 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 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 decreased 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 6-1

**Data Summary Sheet**

**Charles' Law Experiment**

<p>| | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>Name</td>
<td></td>
</tr>
<tr>
<td>Date</td>
<td></td>
</tr>
<tr>
<td>Lab Section</td>
<td></td>
</tr>
</tbody>
</table>

Temperature of the hot water bath

_________ °C

Volume of water in 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 cold air

_________ ml/°K

Volume / temperature ratio for hot air

_________ ml/°K

1. What is the mathematical expression of Charles' Law?

2. A sample of gas had a volume of 275 ml when its temperature was 298K. If the pressure remains constant, what will be its volume when the temperature is decreased to 235K?

3. A sample of argon had a volume of 19.7 liters at 25°C. At what Celsius temperature would its volume be 25.0 liters? (Assume constant pressure).
In this experiment, you will use demonstrate Boyle's Law using the LabWorks\textsuperscript{TM} interface and a pressure sensor as well as a visual measurement.

Boyle's Law can be used to show a relationship between pressure and volume for gases. The relationship states that the pressure and volume are inversely proportional:

\begin{equation}
V \propto \frac{1}{P}
\end{equation}

For this experiment, you will use a setup similar to that shown in Figure 6-5. Two sections of 5-mm glass tubing bound to a meter stick with rubber bands are held vertically with a ring stand or other device. The glass tubes are connected at the bottom with rubber tubing, and the pressure sensor is connected to the top end of one of the glass tubes, leaving the other glass tube open. With this setup, a small amount of water can be added to create both a trapped air space as well as a column of water. As water is added to the open end, the column of water created will compress the trapped air.

The volume of water in the water column is not important, as the column will generate the same amount of pressure regardless of the diameter of the column. Think about this concept. The pressure developed by water at the bottom of a lake is the same as the pressure developed by a small tube of water with the same vertical height as the depth of the lake. For the trapped air, however, the volume is important and must be calculated for this experiment. The volume for a cylinder of air is:

\begin{equation}
V_{cylinder} = 2\pi r^2 h
\end{equation}

where \( r \) is the radius of the column and \( h \) is the height. Since the volume of air trapped in the rubber tubing between the pressure sensor and the top of the glass tube is constant, it can be calculated once and used later. The height of the column of air in the glass tubing will change and must be measured and recorded.

The SCI Tech pressure sensor is used in this experiment to show its capability to measure pressure. Its output varies linearly with pressure and can be used with or without calibration. For this experiment, simply use the uncalibrated sensor by measuring and recording the millivolt output (mV1 or mV2, depending on which input is used).

**Step 1** — Assemble your pressure-volume apparatus as shown in Figure 6-5. Be sure that the glass tubing is secured in a vertical position.

**Step 2** — Remove the rubber tubing which connects the pressure sensor to the glass tube and add enough water to the glass tubing to form a U-shaped volume approximately 2 cm tall.

**Step 3** — Reconnect the rubber tubing from the pressure sensor to the glass tube and measure the length
of the rubber tubing from the top of the glass tube to the base of the tubing connector of the pressure sensor. Record this length. NOTE: The water level in the tubing may change, but it is not a problem.

Step 4 — Connect the pressure sensor to the LabWorks™ interface.

Step 5 — (Optional) Write a LabWorks™ program which contains the following steps: (a) Request input for the water column height from the keyboard (kbd) and save to Col-A of the spreadsheet; (b) request input for the air column height from the keyboard (kbd) and save to Col-B; (c) get the input from the pressure sensor (mV1 or mV2) and save to Col-C; (d) determine if the user is ready to quit; and (e) if not ready to quit, repeat steps a-d.

Step 6 — Run the LabWorks™ program BOYLE.EXP or the program which you wrote in step 5 in spreadsheet mode.

Step 7 — Without adding any water to the open glass tube, measure and enter the initial height of the water column and air column. The height of the water column is the difference in the height of water in the two tubes. The height of the air column is the difference between the top of the glass tube and the level of water in the closed glass tube. NOTE: To take accurate level readings, it may be helpful to hold an index card between the glass tube and the meter stick such that the card is at the height of the bottom of the meniscus. The height of the water column may then be read where the card crosses the graduations of the meter stick.

Step 8 — Add enough water to raise the level in the open glass tube about 10 cm.

Step 9 — Measure and enter the height of the water column and air column.

Step 10 — Repeat Steps 8-9 until the open glass tube is filled or the water level reaches the top of the meter stick.

Step 11 — Stop your program and open the data file you just created with the Spreadsheet.

Calculations

Step 12 — Use the Spreadsheet to determine the pressure for each step. Remember, the pressure is the atmospheric pressure plus the pressure created by the column of water (i.e. $P_{\text{total}} = P_{\text{atm}} + P_{\text{water}}$). NOTE: You may wish to convert the atmospheric pressure to cm of H$_2$O so that you can use the height of the water column directly.

Step 13 — Use the Spreadsheet to determine the volume for each step. The volume is the air volume in the pressure sensor hose as well as the air volume in the glass tube (i.e. $V_{\text{total}} = 2\pi[(r_t^2 \times h_t) + (r_g^2 \times h_g)]$, where $r_t$ and $h_t$ are the radius and height of the rubber pressure sensor tubing and $r_g$ and $h_g$ are the radius and height of the air in the glass tube).

Step 14 — Plot pressure vs. volume and obtain a printout. Does this plot form a straight line? Should it?

Step 15 — Calculate the inverse of the pressure column ($1/P$) and make a plot of $1/P$ vs. volume. Does this plot form a straight line? Should it?

Step 16 — Plot a linear regression of $1/P$ vs. volume and obtain a printout. According to the linear regression equation, at what value for $1/P$ does the volume equal zero?

Step 17 — Repeat Steps 14-16 with the measurements taken with the pressure sensor. How do these results compare to the results with the pressure taken from the height of the water column?

Step 18 — Plot a linear regression of the pressure output in Step 17 against the volume taken from the pressure sensor. Is the plot linear? Why or why not?

Step 19 — If you ran the experiment again, could you use the linear regression equation and the output from the pressure sensor to obtain accurate pressure readings?
Level 3

Experiments 6-5 and 6-6

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 6-5  
Adiabatic Gas Expansion. When compressed gases expand, they must perform work to overcome the attractive forces between the molecules. Design and test an experimental setup to determine the amount of work (as heat) developed by such an expansion.

Experiment 6-6  
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 8½ x 11 inch paper. Typing the text portions is recommended. You and your lab partner will submit separate reports, although your data will be identical. The due date is up to your instructor.

As part of this project you will be required to give a five to ten minute presentation to the class. You and your partner may divide this up any way you wish, but you both must participate in the presentation.

The following sections should be included in your report. Grammar, spelling and neatness do count!

A. Purpose — This should be in your own words, and brief; no more than two or three sentences.

B. Experimental Measurements and Data — Include all data that you used in calculations, discussion, and conclusions. Present your data as tables wherever you can. Spreadsheet graphs and your raw notes of observations during the experiments should also be included.

C. Data Analysis/Calculations — Any formula calculations you used to analyze data must be shown here. If a calculation was repeated several times the multiple results should be shown as a table, but the calculation only needs to be shown once. In many cases a composite table with the raw data and the results all together is necessary, and your lab instructor may require you to do this.

D. Discussion/Conclusions — Briefly 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 need to explain why and support that explanation with your data.

E. References — Document any outside references you used in preparing your report. You should use books, journals, and people as references. The people in your lab and your instructor do not have to be documented as references.