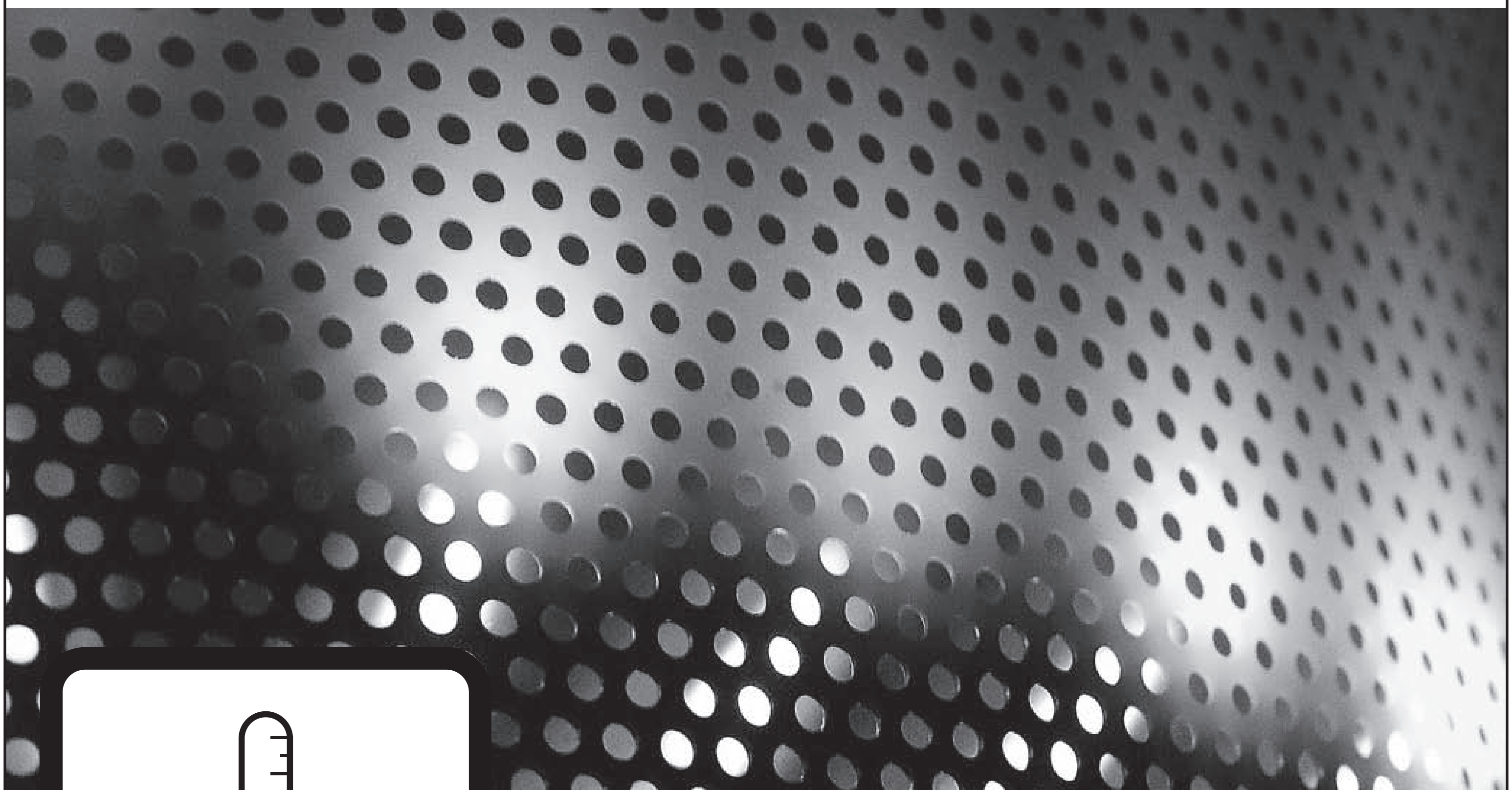
2012-2013

Thermodynamics

Student Guide



National Energy Education Development Project

SECONDARY

Introduction

Your teacher has introduced you to thermal energy and how it affects matter. Now you will investigate and experiment with these concepts in the laboratory. You will explore six lab stations—one each day for the next six days. For you to complete the labs correctly and safely, follow these instructions:

1. Familiarize yourself with the lab guide. It includes:
    - The *Scientific Concepts* covered in the labs
    - *Metric Measurements and Conversions*
    - *Lab Safety Rules*
    - Six lab station guides
2. Read the *Learn About It* sections from all of the labs before you begin.
3. Study and follow the *Lab Safety Rules*.
4. Study the *Metric Measurements and Conversions*.
5. Answer the *Think About It* questions in each lab before performing the experiments.
6. Follow the instructions in the lab guides and record your data carefully. Use a calculator to help with the calculations.
7. Read your instructions and make sure you receive all of the materials you need from your teacher before you begin.
8. Share the load. Make sure that everyone in the group contributes to the process. Assign different tasks for each lab so that everyone has a chance to perform the experiments, record the data, and perform the calculations.
9. Make sure you clean up after the lab, dispose of the waste, and store the equipment properly. Remember, six groups of students must use the materials.
10. If a piece of equipment breaks, tell your teacher. He/she will give you instructions for cleaning the breakage and replacing the equipment.
11. Ask your teacher before proceeding if you have questions about the lab procedures.
12. Check off the concepts on the *Scientific Concepts* page as you investigate them. Ask your teacher if you aren't sure you understand a concept.

Scientific Concepts

- All matter is made of very small particles called atoms.
- Atoms consist of even smaller particles called protons, neutrons, and electrons.
- Protons and neutrons make up the nucleus, which contains almost all the mass of the atom.
- Two atoms that have the same number of protons but a different number of neutrons are called isotopes.
- Protons have a positive electrical charge. Electrons have a negative electrical charge equal in magnitude to the positive charge of one proton. Neutrons have no electrical charge.
- Electrons have very little mass and move in energy levels around the nucleus.
- The distances from the electrons to the nucleus are great, relative to the size of the nucleus—matter is basically empty space.
- Atoms combine to make substances either by transferring electrons and forming ions, or by sharing electrons and forming covalent bonds. Atoms or molecules within a substance are held together because of attractive forces between them.
- Atoms are constantly in motion.
- Of the three states of matter, solids have the lowest amount of thermal energy, and are therefore most affected by the attractive forces between their particles. Gases have a very high amount of thermal energy, and are least affected by attractive forces. A liquid of a given substance falls between solid and gas in terms of their thermal energy, and thus is less affected by attractive forces compared to solids, and more affected compared to gases.
- Increasing the amount of thermal energy in a substance increases its likelihood of overcoming the attractive forces between the atoms or molecules.
- Adding enough energy to a substance allows it to change state.
- Mass is the measure of the amount of matter in a substance. During phase changes and chemical reactions, the amount of mass before and after remains the same.
- Volume is the measure of the amount of space occupied by a substance.
- Adding thermal energy to a substance causes the particles to move faster; as a result, most substances expand when energy is added or contract when energy is removed.
- Water is a unique substance. At 4°C, it is most dense; adding energy to water at this temperature causes it to expand, but cooling water at this temperature also causes it to expand.
- Temperature is a measure of the average kinetic energy of the atoms in a substance.
- Thermal energy can be transferred by conduction—direct contact between two substances with different amounts of thermal energy.
- Substances that transfer thermal energy easily are called conductors. Substances that do not conduct thermal energy well are called insulators.
- Thermal energy can be transferred by convection—currents of fluids caused by differences in temperature and density.
- Energy can be transferred by radiation—energy traveling in electromagnetic waves. Unlike conduction and convection, radiation can occur in a vacuum.
- The specific heat of a substance is the amount of energy required to raise the temperature of one gram of that substance by one Celsius degree.
- The heat of fusion of a substance is the amount of energy needed to change one gram of a substance at its melting point into a liquid without an increase in temperature.
- The heat of vaporization of a substance is the energy needed to change one gram of a liquid substance at its boiling point into a gas without an increase in temperature, at standard pressure.

Metric Measurements and Conversions

Length—*THE DISTANCE FROM ONE POINT TO ANOTHER*

Measuring tool		Metric ruler
Meter (m)		Standard unit of measurement—one yard is 0.9144 m
Centimeter (cm)		one hundredth of a meter—one inch is 2.54 cm
Millimeter (mm)		one thousandth of a meter—one inch is 25.4 mm
Kilometer (km)		one thousand meters—one mile is 1.609 km
1 m	=	100 cm
1 m	=	1,000 mm
1 m	=	1/1,000 km
1 km	=	1,000 m

Volume—*THE AMOUNT OF SPACE A SUBSTANCE OCCUPIES*

Measuring tool		Graduated cylinder
Liter (L)		Standard unit of measurement—one gallon is 3.875 L and one quart is 0.946 L
Milliliter (mL)		one-thousandth of a liter—one teaspoon is 5 mL
1 L	=	1,000 mL
1 L	=	1,000 cm^3
1 mL	=	1 cc or cm^3 (cubic centimeter)

Mass—*THE AMOUNT OF MATTER IN A SUBSTANCE*

Measuring tool		Balance
Gram (g)		Standard unit of measurement— 1cm^3 H_20 at 4°C, a nickel is about 5 g
Kilogram (kg)		1,000 grams—about 2.2 pounds
1 g	=	1,000 mg
1 kg	=	1,000 g

Density—*THE AMOUNT OF MASS IN A GIVEN VOLUME OF A SUBSTANCE*

Density	=	Mass/volume = g/cm^3 (solids) or g/mL (liquids)

Temperature—*THE MEASURE OF HOW HOT OR COLD A SUBSTANCE IS*

Measuring tool		Celsius thermometer
Freezing point of water	=	0°C
Boiling point of water	=	100°C at standard pressure at sea level
Normal body temperature	=	37°C
Room temperature	appr.	21°C

Thermal Energy—*THE AMOUNT OF INTERNAL ENERGY IN A SUBSTANCE*

Measured in Joules

4.184 Joules (one calorie) is the amount of thermal energy needed to raise the temperature of 1g of water 1 Celsius degree.

Fahrenheit/Celsius Conversion

On the Fahrenheit scale, the freezing point of water is 32° and the boiling point of water is 212°—a range of 180°.

On the Celsius scale, the freezing point of water is 0° and the boiling point of water is 100°—a range of 100°.

To convert from Celsius to Fahrenheit, multiply the C number by $\frac{180}{100}$ or $\frac{9}{5}$, then add 32, as shown in the formula below.

$$F = (C \times \frac{9}{5}) + 32$$

If C = 5

$$F = (5 \times \frac{9}{5}) + 32$$

$$F = 9 + 32$$

$$F = 41$$

To convert from Fahrenheit to Celsius, subtract 32 from the F number, then multiply by $\frac{100}{180}$ or $\frac{5}{9}$ as shown in the formula below.

$$C = (F - 32) \times \frac{5}{9}$$

If F = 50

$$C = (50 - 32) \times \frac{5}{9}$$

$$C = 18 \times \frac{5}{9}$$

$$C = 10$$

Lab Safety Rules

Thermodynamics means the "heat movement." All of these labs involve the use of heat. Most of the labs also involve the use of chemicals, fire, and glassware. All of these can cause injury if mishandled.

Eye Safety

- Always wear safety glasses when performing experiments.

Fire Safety

- Do not heat any substance or piece of equipment unless specifically instructed to do so.
- Be careful of loose clothing. Do not reach across or over a flame.
- Keep long hair pulled back and secured.
- Do not heat any substance in a closed container.
- Always use the tongs or protective gloves when handling hot objects. Do not touch hot objects with your hands. Do not use paper towels as 'hot pads.'
- Keep all lab equipment, chemicals, papers, and personal effects away from the flame.
- Extinguish the flame as soon as you are finished with the experiment and move it away from the immediate work area.
- Remember hot objects look the same as cool objects.

Heat Safety

- Always use tongs or protective gloves when handling hot objects and substances. Do not use paper towels as 'hot pads.'
- Keep hot objects away from the edge of the lab table, in a place where no one will accidentally come into contact with them.
- Do not use the steam generator without the assistance of your teacher.
- Remember that many objects will remain hot for a long time after the heat source is removed or turned off.

Glass Safety

- Never use a piece of glass equipment that appears cracked or broken. Report all cracks or breaks to your teacher.
- Handle glass equipment carefully. If a piece of glassware breaks, do not attempt to clean it up yourself. Inform your teacher.
- Glass equipment can become very hot. Use tongs if glass has been heated.
- Clean glass equipment carefully before packing it away.

Chemical Safety

- Do not smell, touch, or taste chemicals unless instructed to do so.
- Keep chemical containers closed except when using them.
- Do not mix chemicals without specific instructions.
- Do not shake or heat chemicals without specific instructions.
- Dispose of used chemicals as instructed. Do not pour chemicals back into a container without specific instructions to do so.
- If a chemical accidentally touches your skin, immediately wash the area with water and inform your teacher.
- Avoid inhaling vapors from chemicals.

Lab One: Matter and Forces of Attraction

Learn About It

Atoms are the building blocks of the universe—all matter is made of atoms. Atoms make up **elements**, the purest form of matter. Atoms are made of even smaller particles. In the center of the atom is the **nucleus**, which is made of **protons** and **neutrons**. Moving around the nucleus in orbitals are **electrons**.

The protons and neutrons in the nucleus contain almost all of the **mass** of an atom. In a carbon atom, for example, the nucleus contains 99.96% of the atom's mass. The electrons orbiting the nucleus are very small and at a great distance from the nucleus.

The protons of atoms carry a positive electrical charge that is equal to the negative electrical charge carried by the electrons. The neutrons carry no charge. The opposing charges of protons and electrons produce a force of attraction between them, as well as between adjoining atoms. The number of valence electrons (electrons in the outer shell) each atom has determines how the atom will form a compound. Different types of compounds are attracted to each other in varying strengths. Adding or removing energy can alter the attraction between molecules. Intermolecular forces are stronger with some types of bonds than others.

One gram of most substances in **solid** form has less thermal energy than one gram of the same substance in **liquid** form; therefore, the attractive forces hold the molecules in the solid in a fixed position. The particles of a liquid have more energy, so they are able to move around each other. However, they do not have enough energy to completely break free of the intermolecular attractions. The energy of **gas** particles is so high that they are able to completely escape the attractive forces that might otherwise hold them together. Generally speaking, the larger the molecule, the greater the intermolecular attraction it has for another molecule. This phenomenon is due to several factors, including the attraction between the nucleus and electrons of a large atom, the degree of polarity of a covalent chemical bond, and overall molecular mass. Substances which are gases at room temperature are usually very small molecules.

Think About It

Using the chart from the previous page, answer the following questions:

1. Will the mass of a sample of water change if it is heated? Will the volume?

2. Will a cylinder of hot water (90°C) have more space between its molecules than a cylinder of cold water (10°C)? Design an experiment to prove your hypothesis.

Experiment With It

Let's investigate the emptiness of matter and the effect that adding thermal energy has on the force of attraction in liquids.

✪ Objectives

- Model the empty spaces within and between molecules and atoms.
- To investigate the force of attraction and the effect of adding thermal energy on liquids.

Hypothesis

Using an "if…then…because…" format, record a hypothesis that describes the effects of adding thermal energy to a liquid.

Supplies and Preparation

Part I - Teacher Demonstration

- Beads
- Marbles
- 2 250 mL Graduated cylinders
- Balance

Part II - Water and Ethyl Alcohol

- 2 250 mL Graduated cylinders
- Ethyl alcohol
- Warm water (about 50°C)
- Triple beam balance or electronic balance

Part III - Water and Salt

- 1 100 mL Graduated cylinder
- 1 25 mL Graduated cylinder
- Container of salt
- Warm water (about 50°C)
- Triple beam balance

Part IV - Forces of Attraction in Liquids

- Waxed paper
- A few drops of water
- 2 Sealed test tubes of corn syrup
- Small container of ice
- Warm water
- Eye dropper

☑ Procedure

Part I - Teacher Demonstration

The beads and marbles represent atoms in the demonstration.

1. Record the data given to you by your teacher on the recording form.

Part II - Water and Ethyl Alcohol

1. Record the mass of one empty cylinder. Add 140 mL of ethyl alcohol and record the exact volume of the liquid and mass of the container with liquid mass.
2. Record the mass of the second empty cylinder. Fill the cylinder with 60 mL of warm water (50°C) and record the mass.
3. Pour the water into the cylinder of ethyl alcohol. Record the resulting volume of the liquid mixture and mass of the cylinder with liquids.
4. Pour the mixture into the lab sink and rinse the cylinders.

Part III - Water and Salt

1. Record the mass of the 25 mL cylinder. Add salt to about the 10 mL mark and record the mass of both.
2. Record the mass of the 100 mL cylinder. Add 50 mL of warm water and record the mass.
3. Pour the salt into the cylinder of water. Gently swirl the mixture. Record the volume of the liquid mixture and mass of the cylinder with liquids.
4. Pour the mixture into the sink and rinse and dry the cylinders.

Part IV - Forces of Attraction in Liquids

1. Place one sealed test tube of corn syrup into the container of ice. Place the other into warm water for about ten minutes.
2. Sprinkle drops of water on the sheet of waxed paper as it is held as flat as possible by two partners. Gently tilt the paper. Observe the force of attraction between the drops of water.
3. Remove the test tube of corn syrup from the container of ice. Turn the test tube upside down.
4. Observe the time it takes for the cold corn syrup to flow from one end of the test tube to the other.
5. Remove the test tube of corn syrup from the warm water. Turn the test tube upside down. Observe the time it takes for the warm corn syrup to flow from one end of the test tube to the other.

Record Data

Part I—Teacher Demonstration	Empty Mass (g)	Full Mass (g)	Mass of Substance (g)	Volume of Substance (mL)
Cylinder with Marbles				
Cylinder with Beads				
Cylinder with Marbles and Beads				

Part II—Water and Ethyl Alcohol	Empty Mass (g)	Full Mass (g)	Mass of Substance (g)	Volume of Substance (mL)
Cylinder with Ethyl Alcohol				
Cylinder with Water				
Cylinder with Ethyl Alcohol				

Part III—Water and Salt	Empty Mass (g)	Full Mass (g)	Mass of Substance (g)	Volume of Substance (mL)
Small Cylinder with Salt				
Large Cylinder with Water				
Large Cylinder with Salt and Water				

Part IV—Forces of Attraction in Liquids	Time to Flow to Bottom of Test Tubes (s)
Cold Corn Syrup	
Warm Corn Syrup	

Calculating and Interpreting

Part I—Teacher Demonstration

1. Calculate the mass of the marbles, beads, and marble/bead mixture.
2. Was the mass of the marble/bead mixture the same as the sum of their masses separately? Explain how this can be.
3. Compare the volume of the mixture to the volumes of the marbles and beads separately.
4. Was the volume of the marble/bead mixture the same as the sum of their volumes separately? Explain how this can be.

Part II—Water and Ethyl Alcohol

1. Calculate the mass of the water, ethyl alcohol, and water/ethyl alcohol mixture.
2. Was the mass of the mixture the same as the sum of the masses of water and ethyl alcohol separately? explain how this can be.
3. Compare the volume of the mixture to the volumes of the water and ethyl alcohol separately.
4. Was the volume of the ethyl alcohol/water mixture equal to the sum of the volumes separately? Explain how this can be.

Part III—Water and Salt

1. Calculate the mass of the water, salt, and water/salt mixture.
2. Was the mass of the mixture the same as the sum of the masses of water and salt separately? Explain how this can be.
3. Compare the volume of the mixture to the volumes of the water and salt crystals separately.
4. Was the volume of the water/salt mixture equal to the sum of the volumes separately? What factor(s) account for this?
5. How would this part of the experiment be different if sand was used instead of salt?

Part IV—Forces of Attraction in Liquids

1. What did the movement of the droplets of water on the waxed paper show you about intermolecular attraction in water?

2. What did the flow rates of the cold and warm corn syrup show you about intermolecular attraction in warm or cold liquids?

3. On a separate sheet of paper, write the objective, materials, and procedure for testing the effect of adding thermal energy to liquids on the intermolecular attraction in those liquids.

Make Sure You Understand It

1. A cylinder with 200 mL of water was poured into a cylinder with 200 mL of ethyl alcohol. Five percent of the water filled in the empty spaces between the alcohol molecules. What was the volume of the mixture?

Extension

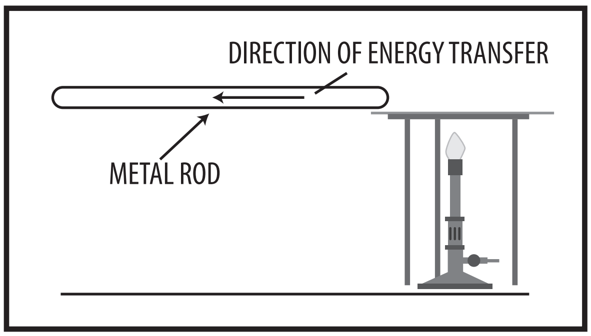
1. How does the bottling company successfully make soda with such a high sugar content? Design an experiment to figure it out.

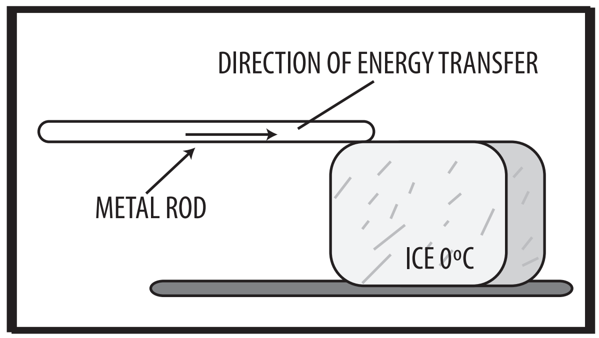
Lab Two: Thermal Energy Transfer by Conduction

Learn About It

Thermal energy can be transferred three ways—by conduction, convection, and radiation. **Conduction** is thermal energy transfer by direct contact between objects. Conduction occurs when two substances with different temperatures come in contact with each other. The warmer, faster moving molecules collide with the cooler, slower moving molecules and give up some of their energy. Conduction occurs in solids, liquids, and gases.

Energy transfer via conduction always occurs from substances or areas of high temperature to substances or areas of lower temperature. If you place a metal rod against a heated object, the vibrating atoms in the heated object will transfer some of their energy to the atoms in the rod and cause them to vibrate more. Those atoms in turn will pass some of their energy to the next atoms in the rod, and so on down the line until the entire metal rod has reached the same temperature as the object with which it first came into contact.





If the same rod is placed on an ice cube, the low-energy molecules of the ice will absorb some energy from the rod and slow down the vibration of the atoms that are touching the ice. The atoms in the warmer section of the rod will then supply some of their energy to the less energetic atoms at the end touching the ice. The atoms supplying the energy to the colder end will become less energetic. This process of supplying energy from the warmer section of the rod to the cooler section will continue until the entire rod cools to the same temperature as the ice.

The atomic mass and structure of a substance affect its ability to transfer energy between adjoining atoms. Substances that transfer thermal energy quickly are called conductors. Substances that transfer thermal energy slowly are called **insulators**. Table 2.1 lists the conductivities of some common substances. The higher the number, the better the substance conducts thermal energy. Good conductors typically are structured with their atoms or molecules arranged very close together. Materials that we commonly use as thermal insulators, such as fiberglass, textile fibers, and feathers, have air pockets within them that do not allow thermal energy to be transferred easily.

Table 2.1: Conductivities of Some Common Substances

Substance	Conductivity *W/(m • K)*
Diamond	42,907
Copper	6,832
Aluminum	4,205
Brass	1,925
Lead	615
Steel	238
Nickel	138
Concrete	21
Glass	12.55
Water	12.55
Wood (Oak)	2.51
Wool	0.84
Goose Down	0.42
Air and Most Gases	0.42
Foam Polystyrene	0.17

Thermal energy is measured in Joules, and thermal conductivity is measured in J/(s•m•K). One Joule per second (J/s) is equal to one watt of power (W). Therefore, conductivity is listed in Watts/(m•K).

Think About It

Using the chart above, answer the following questions:

1. If your teacher placed a thermometer on the bar during their demo, what do you think the temperature would be?

2. Which substance is the best at transferring thermal energy?

3. Which substance is the poorest at transferring thermal energy?

4. Why does a concrete wall feel cooler than a wood wall in the same room?

5. Of the metals listed on the top of the chart, which would be the best to use when making a frying pan?

6. How are substances with low conductivity rates used in our daily lives?

Experiment With It

Let's investigate conduction and insulation.

✪ Objectives

- To investigate the insulating qualities of different materials.
- To investigate the conductivity of different metals.

Hypothesis

Using an "if…then…because…" format, record a hypothesis that describes how different materials transfer thermal energy.

Supplies and Preparation

Part I - Insulation

- Foam polystyrene cup with lid
- Steel crucible with lid
- Porcelain crucible with lid
- 3 Thermometers
- Shallow pan with 1/2 inch of ice water
- Warm water (about 50°C) in beaker
- Beaker tongs

Part II - Conductivity

- Conductometer
- Birthday candles
- Candle (or alcohol lamp)
- Safety pin
- Matches
- Ring stand with clamp
- Watch with second hand

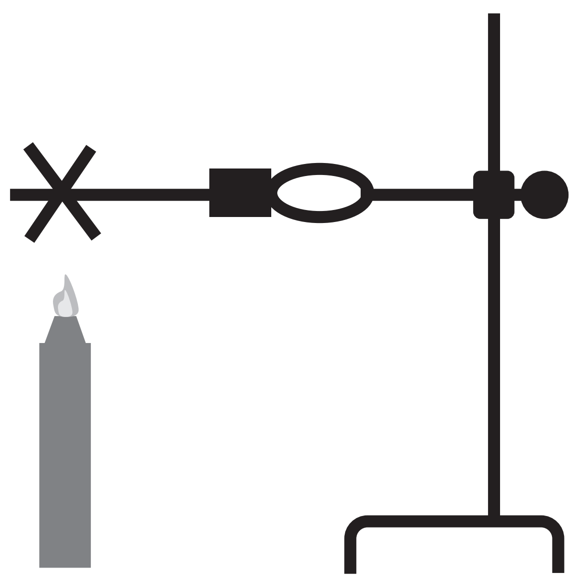
☑ Procedure

Part I - Insulation

1. Measure the temperature of some hot water and record in the data table.
2. Fill the foam cup and crucibles each with 30 mL of hot water, and cover with their respective lids. Place containers in the pan of ice water.
3. Wait five minutes, then record the temperature of the water in each container. Proceed to Part II of the lab while you are waiting.

Part II - Conductivity

1. Examine the conductometer. Note the notched sides of the rod ends. Note that the five rods are labeled with letters.
2. Make a sketch of the conductometer below with the letters so that you will know which metal from which each rod is made while you are performing the experiment.
3. Barely push the point of the safety pin about 2 mm into one birthday candle piece.
4. Carefully light the candle.
5. Gently heat the candle piece about 3 cm above the flame for two or three seconds to soften one side.
6. Quickly press the melted portion of the ball to the notched end of one of the rods on the conductometer.
7. Allow to cool for a few seconds, then carefully remove the safety pin by placing your finger on the candle piece.
8. Attach the other candle pieces to the ends of the rods in the same way. Turn the conductometer over to make sure that all the balls are securely attached.
9. Clamp the handle of the conductometer to the ring stand with the balls located underneath the rods.
10. Place the candle directly under the center disc of the conductometer with the top of the flame 1–2 cm below the disc. It is important to heat the exact center the disc.
11. Start the timer as soon as the candle is placed under the disc.
12. Record the time and the metal rod as each ball falls. Don't stop timing until all the balls have fallen.
13. Extinguish the candle.



Record Data

Part I	Initial Temperature	Final Temperature	Difference in Temperature
Foam Cup			
Iron Crucible			
Porcelain Crucible			

Part II	Aluminum	Brass	Steel	Nickel	Copper
Time					

Calculating and Interpreting

Part I—Insulation
Use Table 2.1 and your lab data to determine which of the three materials (steel, porcelain, and foam polystyrene) fit the description:

(Table) Best Conductor:		(Table) Worst Conductor:	
(Lab) Data Best Conductor:		(Lab) Data Worst Conductor:	

What evidence from your lab indicated the best and worst conductors? How do your results compare to the standard values given in the table?

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Part II—Conductivity

The conductivity values given in Table 2.1 have units of Watts/meter•Kelvin degree. If the procedure is followed precisely, the flame is centered on the conductometer and the paraffin balls are equidistant from the center. Assuming the same amount of thermal energy is being transferred to each piece of metal, each piece of metal is the same length, and each began at the same temperature (room temperature), the only difference will be the amount of time required to transfer energy along each piece of metal. The relative conductivity (how conductive one metal is compared to another) can then be calculated as follows:

$$\textit{Relative Conductivity} = \frac{\textit{Conductivity of metal}}{\textit{Conductivity of standard}}$$

$$\textit{Relative Conductivity} = \frac{\frac{Watts}{meter \times K} \textit{ of metal}}{\frac{Watts}{meter \times K} \textit{ of standard}}$$

$$\textit{Relative Conductivity} = \frac{\frac{\cancel{Joules}}{second \times \cancel{meter} \times \cancel{^{o}C}} \textit{ of metal}}{\frac{\cancel{Joules}}{second \times \cancel{meter} \times \cancel{^{o}C}} \textit{ of standard}}$$

$$\textit{Relative Conductivity} = \frac{\frac{1}{second} \textit{ of metal}}{\frac{1}{second} \textit{ of standard}}$$

$$\textit{Relative Conductivity} = \frac{\textit{time of standard}}{\textit{time of metal}}$$

Look at your lab data. The metal whose paraffin ball fell last was the least conductive metal. This metal will be the standard we use for comparing the relative conductivities of the other four metals. Thus, the relative conductivity for this metal will be 1.0. Using the above relationship, calculate the relative conductivity of the other four metals as compared to the least conductive:

Metal	Time to Melt Wax (es)	Relative Conductivity
Nickel		
Brass		
Copper		
Aluminum		
Steel		

How do your relative conductivities compare to the standards listed in Table 2.1? Is your least conductive metal the least conductive of the five used in the experiment? What about your most conductive metal? Give some plausible explanations for any significant differences.

Make Sure You Understand It

Using Table 2.1 and the formulas on the previous page, answer the following problems:

1. How many times more conductive is aluminum than brass?

2. Two identical sized rods—one glass and one nickel—were heated. The glass object transferred thermal energy at a rate of 1,000 Joules per second through the rod. How many Joules per second were transferred by the nickel rod?

3. Using a conductometer in a perfect environment, the paraffin ball on the nickel rod fell in 8 minutes 15 seconds. Calculate the time that it took for the other paraffin balls to fall.

Lab Three: Convection and Radiation

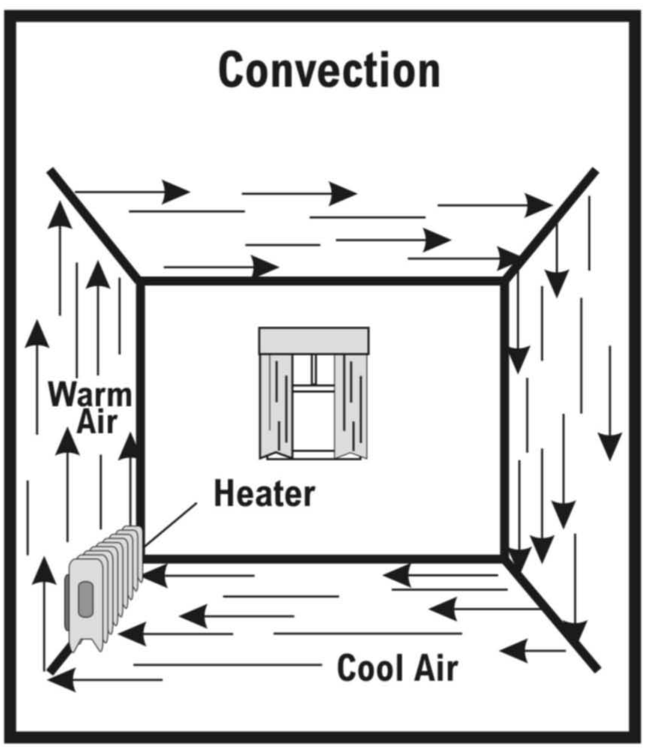
Learn About It

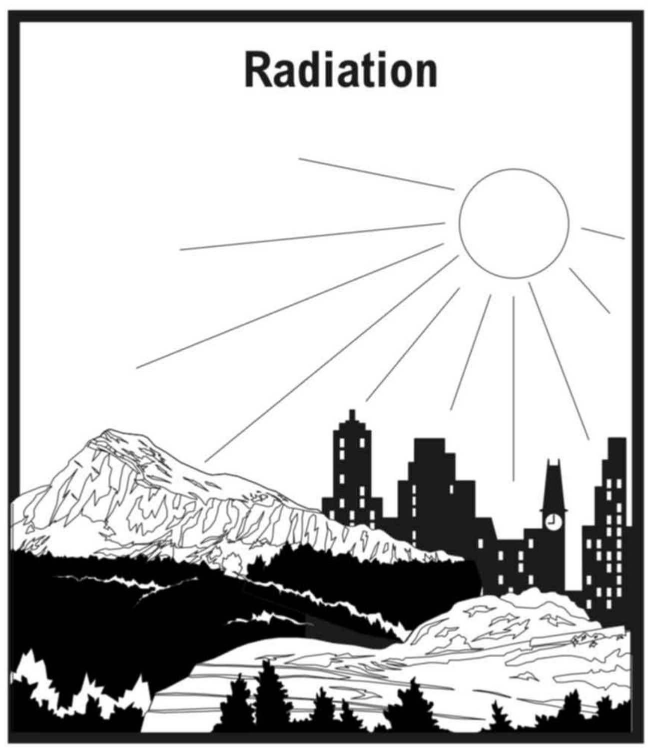
Thermal energy can be transferred three ways—by conduction, convection, and radiation. Conduction, as seen in Lab Two, is thermal energy transfer by direct contact between objects.

Convection is thermal energy transfer by the motion of molecules in currents. Convection can only occur in liquids and gases—in substances in which the molecules are free to move. In solids, the molecules are held in fixed positions, so there can be no flow of molecules.

In liquids and gases, the molecules can move freely. When one part of a substance is heated, the motion of the molecules in that part increases. As the motion increases, the molecules begin to spread out and the gas or liquid becomes less dense. The molecules in the cooler part of the substance are more closely packed together. Since the warmer part of the fluid is less dense, it rises, and cooler gas or liquid moves in to take its place. This motion produces currents that carry thermal energy.

Radiation is the transfer of energy through space in the form of rays. Energy from the sun reaches the earth through radiation. When the rays are absorbed by an object, the energy is changed to thermal energy, the atoms or molecules vibrate faster, and the object feels warmer. Objects with high amounts of energy can emit radiation in the form of light or infrared radiation.





Think About It

1. Why do you feel warmer in the sun than in the shade, when the temperature of the air is the same?

2. If you could only heat one floor of a two-story house in the winter, which floor would you heat? Why?

3. If you burn toast in the kitchen, why can you smell it all over the house?

4. Why do workers in the tropics wear long-sleeved, light-colored clothing?

5. Why does the wind tend to blow inland from the ocean during daylight hours?

Experiment With It

Let's investigate radiation and convection.

Objectives

- To investigate the transfer of energy by radiation.
- To investigate the transfer of thermal energy by convection in liquids and gases.

Hypothesis

Using an "if…then…because…" format, construct a hypothesis describing how thermal energy is transferred in gases and liquids.

Supplies and Preparation

Part I - Radiation

- Black and silver containers
- Thermoconductivity strip
- 2 Thermometers
- Radiation source (infrared lamp)
- Index card
- Tape
- Timer or clock
- Ruler

Part II - Convection in Gases

- Thermoconductivity strip
- Candle or alcohol lamp
- Matches

Part III - Convection in Liquids

- U-tube
- Water
- Food coloring
- Ring stand with clamp
- Candle or alcohol lamp
- Matches

☑ Procedure

Part I - Radiation

1. Observe the color of the crystals in the conductivity strip. Fold an index card tightly around the middle of the crystal strip. Fasten the card with tape.
2. Turn on your source of radiation, following your teacher's instructions.
3. Hold the conductivity strip about 20-25 cm away from the radiation source until the color begins to change. Remove the index card and observe the color of the crystals in the conductivity strip under the card. (Holding the strip too close to the lamp may ruin your strip).
4. Place the thermometers in the silver and black containers and record the temperature of the air inside the containers. Be sure your thermometers are not touching the sides or bottom of the cans.
5. Direct the radiation source onto the containers from a distance of 45-55 cm. Record the temperature of the containers every five minutes for 20 minutes. Proceed to Parts II and III of the lab during the timing.

Part II - Convection in Gases

1. Carefully light the candle or alcohol lamp.
2. Make a note of the color of the crystals.
3. Using the handles, hold the conductivity strip to the side of the candle, about 30 cm away for 15 seconds. Observe and make a note of any color change.
4. Now move the conductivity strip directly over the candle about 30 cm above the top of the flame for 15 seconds. Observe and make note of any color change.
5. Extinguish the candle and set aside for the next demonstration.

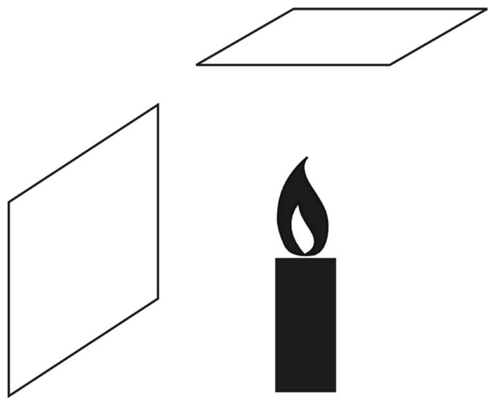
Part III - Convection in Liquids

1. Fill the U-tube with water to the base of the opening.
2. Carefully attach the U-tube to the lab stand with a clamp as shown in the diagram on page 22. The bottom of the U-tube should be about five centimeters above the wick of the candle. Allow the apparatus to stand undisturbed for a minute before proceeding, so that any motion of the water will cease.
3. Place the candle under one corner of the U-tube and carefully light it. Allow the candle to sit under the tube for one minute. Place one drop of food coloring into the water through the opening at the top of the U-tube and observe its movement. Draw arrows to show the movement of the food coloring in the diagram in the recording section on page 22.

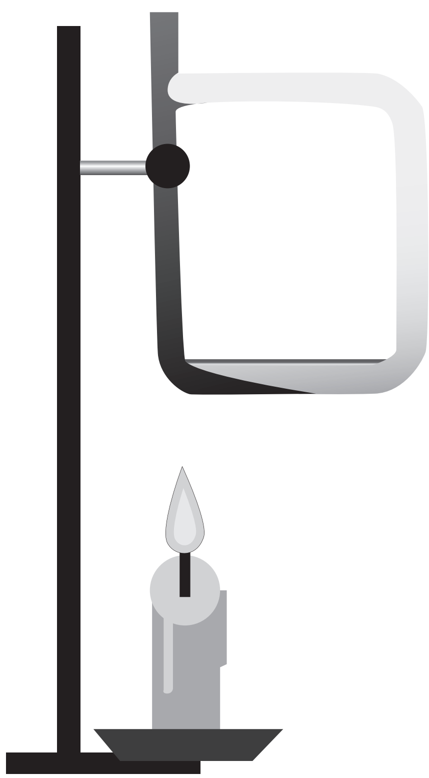
Record Data

Part 1	Initial Temperature	Temperature at 5 Minutes	Temperature at 10 Minutes	Temperature at 15 Minutes	Temperature at 20 Minutes
Black Can					
Silver Can					

Part II	Use colored pencils to record the crystal's color in the different positions.



Part III	Use arrows to show the movement of the food coloring.



Calculating and Interpreting

Part I—Radiation

1. From low thermal energy to high thermal energy, list the colors displayed by the thermoconductivity strip.

2. How did the index card affect the color of the crystals inside the thermoconductivity strip?

3. Explain what you think caused the crystals in the thermoconductivity strip to change color.

4. Describe ways the thermoconductivity strip could be used outside of a science lab.

5. What happened to the air temperature inside the black and silver cans as they were exposed to radiation?

6. Explain any differences you saw in the temperature data between the black and silver cans.

7. Based on your evidence, explain how the finishes of objects should be considered in different climates.

Part II—Convection in Gases

1. Using the evidence from your experiment with the thermoconductivity strip, sketch the pattern or direction of current caused by the heat of the candle.

2. On a separate piece of paper, design an experiment, with objective, materials, and procedure, to demonstrate convection currents.

Part III—Convection in Liquids

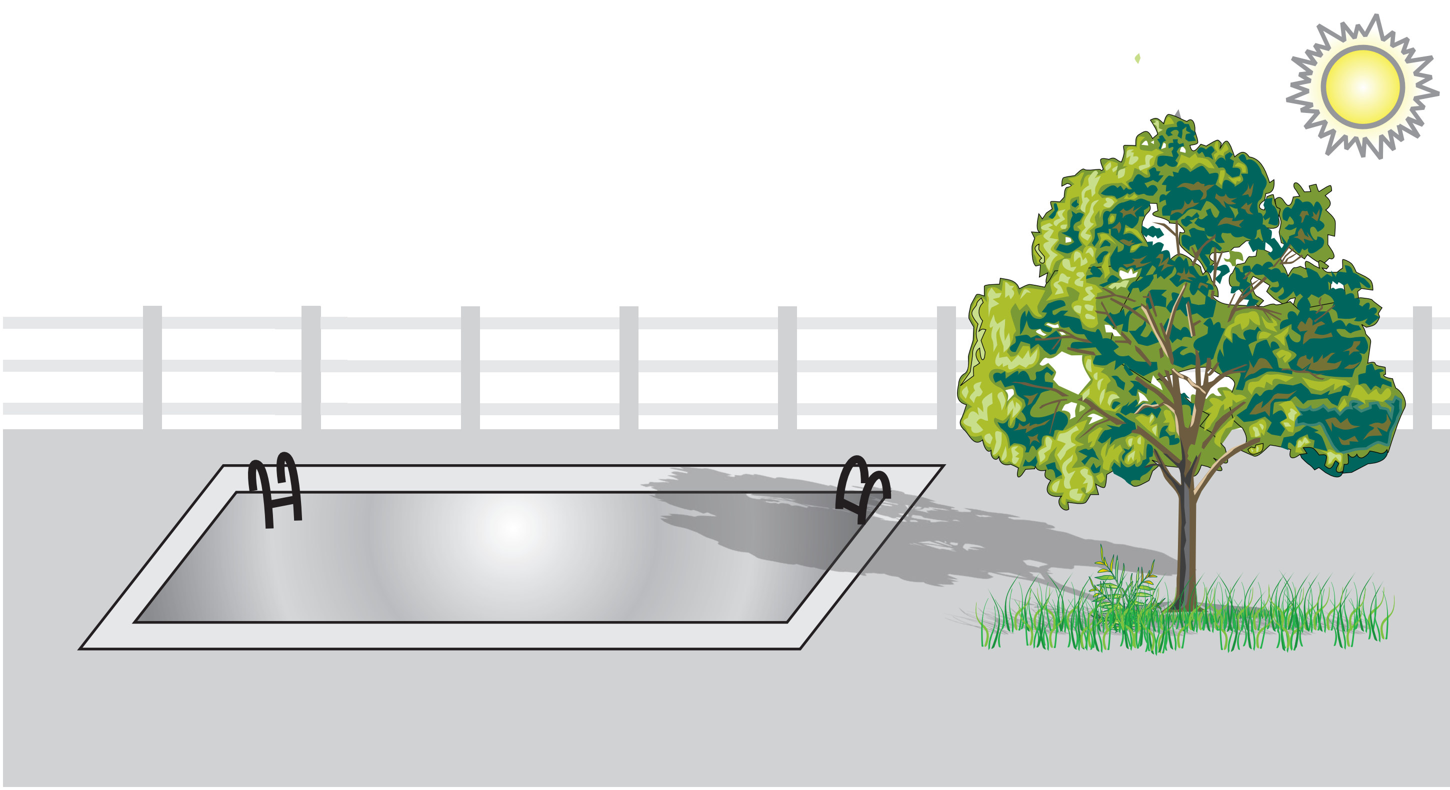
1. Using the evidence from your experiment with the glass tube, describe the convection current caused by the heat of the candle.

2. Based on what you observed, predict how moving the candle to the center of the tube would change the pattern.

3. Using an ocean current map, explain how your experimental result supports the direction of ocean currents.

Make Sure You Understand It

On a hot, summer day, the sun shone down on the concrete at the end of the pool, which was shaded by a big oak tree. All day the temperature in the pool increased, though it remained in the shade. Explain how radiation, convection, and conduction contributed to the heating of the pool. Where in the pool was the water warmest? Describe the energy flows.



Lab Four: Expansion and Contraction

Learn About It

The amount of thermal energy in a substance will determine its state. Whether a substance is a solid, liquid, or gas at a comfortable room temperature—about 22°C—depends on how strongly the atoms and molecules are attracted to each other. **Ionic compounds**, such as everyday table salt, or sodium chloride, have the strongest attraction between the particles. That is why ionic compounds have very high melting points. **Covalent compounds**, like paraffin or methane (CH_4), have much less attraction for each other. Paraffin is a wax and melts fairly easily. Methane, better known as natural gas, is a gas at room temperature. The physical state of a substance depends on the attraction between the particles of the substance as well as the amount of thermal energy in the substance.

Another physical property of matter that is affected by thermal energy is a substance's **density**. When a substance is heated, the particles acquire more thermal energy and vibrate faster. Faster vibrations result in greater space between the particles.

If the substance is a **gas**, the particles will collide more and increase the **pressure** of the gas if the container is rigid, or will cause the gas to expand and become less dense. This is where the saying "hot air rises" comes from. Because it is less dense, warmer air will rise.

If the substance being heated is a **liquid**, the particles will move around each other much more easily than if the substance was colder. If it gets hot enough, the liquid will expand somewhat, but not nearly as much as a gas.

At this point, you may be thinking that **solids** are so rigid they never expand when heated. Not so! Even the hardest substances will expand a little when heated. Different substances will expand at different rates, but they all expand somewhat. That is why bridges are built with "expansion joints" spaced periodically. The expansion joints keep the bridge from buckling or cracking as the temperatures change throughout the year.

Table 4.1 lists some common substances and the rates that their **volumes** change when heated. This is called cubic expansion; as the number increases, the more expansion occurs when heated or the more substance contracts when cooled.

Think About It

Using Table 4.1, answer the following questions.

1. Which substance expands or contracts the most with temperature change? Which substance expands or contracts the least with temperature change?

2. Why are steel rods rather than aluminum rods used to reinforce concrete?

3. Why is Pyrex cookware better than regular glass for cooking? Why does regular glass break more easily when it is used for cooking?

4. If you were building a bridge in an area where the summers are hot and the winters are cold, what would you need to take into account in your construction plans?

5. If you were making a thermometer to measure very slight changes in temperature, which liquid from the chart would you use? Why?

Table 4.1: Cubic Expansion

Substance	State at Room Temperature	Cubic Expansion per °C	
Aluminum	Solid	69×10^{-6}	(0.000069)
Steel	Solid	36×10^{-6}	(0.000036)
Copper	Solid	51×10^{-6}	(0.000051)
Lead	Solid	87×10^{-6}	(0.000087)
Nickel	Solid	39×10^{-6}	(0.000039)
Concrete	Solid	36×10^{-6}	(0.000036)
Glass	Solid	26×10^{-6}	(0.000026)
Pyrex Glass	Solid	10×10^{-6}	(0.000010)
Ethyl Alcohol	Liquid	1120×10^{-6}	(0.001120)
Glycerin	Liquid	505×10^{-6}	(0.000505)
Water	Liquid	207×10^{-6}	(0.00207)
Mercury	Liquid	182×10^{-6}	(0.000182)
Air and Most Gases	Gas	3400×10^{-6}	(0.003400)

Experiment With It

Let's investigate expansion and contraction in the laboratory.

✪ Objectives

- To determine the rates of expansion of water and glycerin.
- To confirm that water at 4°C expands when heat energy is added or removed.
- To observe expansion in metals and gases.

Supplies and Preparation

Part I—Rates of Expansion of Water and Glycerin

- 1,000 mL Beaker of room temperature water (about 22°C)
- 1,000 mL Beaker of hot water (about 70°C)
- Thermometer
- 50 mL Sealed graduated cylinder with 40 mL water
- 50 mL Sealed graduated cylinder with 40 mL glycerin

Part II—Expansion of Water at 4°C

- 2—250 mL Graduated cylinders filled with water
- 2—250 mL Graduated cylinders—sealed with stoppers—with isopropyl alcohol
- Thermometer
- 1,000 mL Beaker of ice water
- 1,000 mL Beaker of warm water (50°C)

Part III—Expansion and Contraction of Solids and Gases

- Compound bi-metal bar
- Gas expansion tube
- Balloon
- Candle or alcohol lamp
- Beaker of ice water
- Beaker of hot water (almost 70°C)
- Matches

☑ Procedure

Part I - Rates of Expansion of Water and Glycerin

1. Place the graduated cylinders in the beaker of room temperature water and carefully remove the stoppers. Set the stoppers aside to replace when you are finished. Allow the cylinders to remain in the beaker for five minutes, then record the temperature of the water in the beaker, and the exact volumes of the water and glycerin in the cylinders. The temperature of the water will be the temperature of the water and glycerin. Record your data in the columns "Initial Temperature" and "Initial Volume."

2. Place the cylinders in the beaker of hot water. Wait five minutes, then record the temperature of the water in the beaker. The temperature of the liquids in the cylinders will be the same temperature as that of the water in the beaker. Record your data in the column marked "Final Temperature."

3. Do not remove the cylinders from the beaker. Hold each upright and record the exact volumes of the liquids in the cylinders. Record the volumes in the column marked "Final Volume."

4. Carefully remove the cylinders from the beaker and allow them to cool to room temperature.

5. Replace the stoppers.

Part II - Expansion of Water at 4°C

1. Warm one cylinder of isopropyl alcohol and one cylinder of water by placing them in a beaker of water at approximately 50°C before beginning the experiment. Place the other cylinders of isopropyl alcohol and water in a beaker filled with ice water.

2. In five minutes, remove the stoppers from the cylinders. With the cylinders remaining in the beakers, record the temperature of the liquids for all four containers at the bottom, then top, of the cylinders, taking care to not stir up the liquids.

Part III - Expansion and Contraction of Solids and Gases

1. Examine the compound bi-metal bar. The bar is made of thin strips of two metals—nickel and steel—bonded together. Look up the expansion rates of these two metals in Table 4.1.

2. Carefully light the candle. Place the middle of the bar into the candle flame for 10 seconds, then into the cold water. Make a diagram of what happened to the bar. In your diagram, be sure to label which side of the bar is nickel and which side is steel.

3. Make sure the balloon is securely over the top of the gas expansion tube. Place the bulb of the tube into a beaker of hot water for one minute, and then place it into a beaker of cold water. Record your observations.

Record Data

Part I	Initial Temperature	Final Temperature	Change in Temperature	Initial Volume	Final Volume	Change in Volume
Cylinder of Water- Room Temperature Beaker						
Cylinder of Glycerin-Room Temperature Beaker						
Cylinder of Water- Hot Water Beaker						
Cylinder of Glycerin-Hot Water Beaker						

Part II	Bottom Temperature	Top Temperature	Difference in Temperature
Cylinder One Warm Water			
Cylinder Two Cold Water			
Cylinder Three Warm Alcohol			
Cylinder Four Cold Alcohol			

Part III		
Diagram of Bi-metal Bar	**Diagrams of Gas Expansion Tube**	
In candle	**Hot Water**	**Cold Water**
In water		

Calculating and Interpreting

Part I - Rates of Expansion of Water and Glycerin

The expansion rate of a liquid can be calculated using the following relationship:

$$\textit{Expansion rate} = \frac{\textit{change in volume}}{\textit{change in temperature x original volume}}$$

$$\textit{Expansion rate} = \frac{\Delta V}{\Delta T \times V_0}$$

Example: A container with 100 mL of ethyl alcohol at 0°C was heated to 100°C. The ethyl alcohol expanded to 111 mL. What is the expansion rate of ethyl alcohol?

$$\textit{Expansion rate} = \frac{\Delta V}{\Delta T \times V_0}$$

$$\textit{Expansion rate} = \frac{(111 - 100)}{(100 - 0) \times (100)}$$

$$\textit{Expansion rate} = \frac{11}{10{,}000} = 1.1 \times 10^{-3}/°C$$

Now calculate the expansion rates of water and glycerin with your own experimental data:

Expansion Rate of Water=	Expansion Rate of Gllycerin=

Compare your results to the actual expansion rates listed in Table 4.1. How do your calculations compare with these values? What might have affected your results?

Part II - Expansion of Water at 4˚C

Most liquids contract when cooled and expand when heated, regardless of the original temperature. As the liquids contract, they become more dense. Cooler liquids, being more dense, will tend to sink to the bottom of a container, and warmer liquids, being less dense, will tend to rise to the top.

If a tall cylinder of liquid is left undisturbed, a two to four degree temperature difference is possible, with the higher temperature at the top and the lower temperature at the bottom. However, water does not always adhere to this principle. Water is most dense at 4°C. If a warm cylinder of water is cooling, the warmest temperature will be at the top and the coolest at the bottom. However, a cylinder of ice water that is being warmed by its surroundings will be around 0°C at the top, and as warm as 4°C at the bottom. This contributes to why lakes and large ponds have a layer of ice at the top and do not freeze solid, and why ice cubes float in a glass of water.

Look at your data for Part II. Do your numbers follow this principle? Describe any factors that might have interfered with the accuracy of your experimental data.

Part III - Expansion and Contraction of Solids and Gases

1. Describe how the change in the bi-metal bar illustrates that different metals have different rates of expansion and contraction.
2. How could a bi-metal strip be used in everyday applications?
3. How would you characterize the speed with which gases expand and contract as compared to liquids and metals?
4. Using your experimental results, describe what would happen to a railroad tank car if cleaned with steam before being closed tightly. Do an image search online to see if you were correct!

Make Sure You Understand It

Using the formulas for calculating change in volume and resulting volume, as shown in the example, solve the problems that follow.

Cubic expansion rate = (change in volume) / (change in temperature)x(original volume)

Change in volume = (original volume)x(temperature change)x(cubic expansion rate)

Resulting volume = original volume + change in volume

Example: A container with two liters of mercury experiences a 50°C increase in temperature. How much will the mercury expand because of this temperature change? What is the resulting volume?

Step One: Calculate the change in volume of the mercury using the formula. Remember the following conversions:

1 liter = 1,000 milliliters $1 \times 10^{-6} = 0.000001$

Change in volume = (original volume) x (temperature change) x (cubic expansion rate)

Change in volume = $(2\ L) \times (50^\circ C) \times (182 \times 10^{-6})$

Change in volume = $(2) \times (50) \times (182 \times 10^{-6}) = 18{,}200 \times 10^{-6} = 0.0182\ L$

Change in volume = 0.0182 L

Step Two: Calculate the resulting volume using the formula.

Resulting volume = original volume + change in volume

Resulting volume = (2 L) + (0.0182 L)

Resulting volume = 2.0182 L

1. What would be the volume of a 10 cubic meter block of aluminum if its temperature increased from 10˚C to 510˚C?
2. A beaker with 500 mL of a mystery liquid at 20˚C was heated to 120˚C. Its volume at that temperature had increased to 510 mL. What is the rate of expansion of the mystery liquid?
3. A container of 100 L of ethyl alcohol was heated until the volume increased to 101.1 L. How many degrees Celsius did the temperature of the ethyl alcohol increase?

Lab Five: Specific Heat

Learn About It

The **specific heat** of a substance is the amount of thermal energy needed to raise the temperature of one gram of that substance 1°C.

Table 5.1 shows the specific heat of several substances. The specific heat of water is 4.184 J/g•°C. It requires 4.184 Joules of energy to raise the temperature of one gram of water 1°C, this is also known as a **calorie.**

Liquid mercury is 13.6 times denser than water. It requires only 0.138 Joules of heat to raise the temperature of one gram of mercury 1°C.

Adding 4.184 Joules of thermal energy to one gram of water will increase its temperature 1°C. That same amount of thermal energy—4.184 **Joules**—will raise the temperature of one gram of mercury 30°C.

The atomic mass and structure of some substances are such that only a small amount of energy is needed to increase the motion of their atoms. They have low specific heats. Substances with high specific heats require more energy to produce the same increase in molecular motion.

The converse is also true. A small loss of energy in some substances causes a large decrease in molecular motion—a large decrease in temperature. Water—with its high specific heat—retains its thermal energy, its molecular motion. A small drop in energy causes a small decrease in temperature.

Think About It

Using Table 5.1 on page 31, answer the following questions.

1. Generally, what effect does density have on the specific heat of a substance?

2. What is the relationship between the atomic mass of a substance and its specific heat? Why do you think this is?

3. Which substance requires the most amount of thermal energy to cause a one degree increase in its temperature?

4. Which substance requires the least amount of thermal energy to cause a one degree increase in its temperature?

5. Why do you think the human body has such a high specific heat? How do you think a detective could utilize this information in a murder case?

Table 5.1: Specific Heat

Substances	Specific Heat (J/g°C)	Density (g/cm^3)	Molecular or Atomic Mass (amu)
Aluminum	0.900	2.7	27
Tin	0.226	7.3	119
Copper	0.385	8.9	63
Ice at -15°C	1.999	0.9	18
Gold	0.126	19.3	197
Iron	0.448	7.8	56
Lead	0.126	11.3	207
Silver	0.234	10.5	108
Stainless steel	.50	7.4-8.0	—
Human Body at 37°C	3.473	—	—
Zinc	0.383	7.14	65
Ethyl Alcohol	2.452	0.81	46
Mercury	0.138	13.6	201
Water	4.148	1.0	18
Air (Average 79% Nitrogen and 20% Oxygen)	1.004	0.0013	14.4

Experiment With It

Let's investigate specific heat in the laboratory.

✪ Objective

- To determine the specific heats of several metals.

Supplies and Preparation

- Metal samples from Specific Heat Demo
- Foam cups with lids—1 per metal sample
- 5 Thermometers
- Tongs and safety gloves
- Triple beam balance or electronic balance
- Beaker
- Hot plate or Bunsen burner and ring stand or tripod
- Distilled water
- Marker
- 100 mL Graduated cylinder

☑ Procedure

1. Determine and record the mass of each of the metal samples.
2. Place the metal samples in a beaker of boiling water for two minutes. Keep the water boiling. The metals will increase in temperature until all are the same temperature as the water. Record the temperature of the water.
3. Mark each foam cup with the name of one of the metal samples. Fill the cups with an amount of room temperature distilled water equal to the mass of the metal samples and record the temperature. (For example, if the samples weigh 50 g, place 50 mL of water in the foam cups.)
4. Using the safety gloves and tongs, remove each sample from the boiling water and place into a marked foam cup. Cover each foam cup with a lid. After 30 seconds, gently shake the foam cups for three seconds.
5. Wait an additional 30 seconds. Record the final temperature of the water in each cup.

Record Data

Metal	Mass of Metal (g)	Temperature of Boiling Water (ºC)	Temperature of Water in Cup (ºC)	Final Temperature of Water and Metal in Cup (˚C)	Specific Heat of Metal (J/g ºC)
1					
2					
3					
4					
5					

Calculating and Interpreting

The amount of thermal energy transferred can be calculated using the following relationship:

$q = m\Delta Ts$

Where

$q =$ amount of thermal energy transferred (J)

$m =$ mass of substance (g)

$\Delta T =$ change in temperature (°C)

and $s =$ specific heat of the substance (J/g•°C).

Example: A sample of iron with a mass of 55 g is taken from a container of boiling water (100°C) and placed into a foam container with 55 mL (55 g) of water at 20°C. After two minutes, the temperature of the water and iron sample combined is 27.7°C. Calculate the specific heat of this sample of iron.

Step 1: Calculate the thermal energy gained by the **water**.

$$q = m\Delta Ts$$
$$q = 55g \times (27.7\ ℃ - 20\ ℃) \times 4.184\ J/g \bullet ℃)$$
$$q = 1{,}771.9\ J$$

Step 2: Calculate the specific heat of the **iron**. The thermal energy lost by the iron is the same as the thermal energy gained by the water.

$$q = m\Delta Ts$$
$$1{,}771.9\ J = 55g \times (100°C - 27.7°C) \times s$$
$$\frac{1{,}771.9\ J}{55g \times (100°C - 27.7°C)} = s$$
$$s = 0.4456\ J/g \bullet °C$$

Table 5.1 lists the specific heat of iron as 0.448 J/g • °C, which is essentially the same as the value calculated from the experimental data.

Following the example above, calculate the specific heat values for each metal sample.

How do the values you calculated compare to the standard values given in Table 5.1? List sources of error that may have affected your results.

Make Sure You Understand It

Using Table 5.1 and the formulas for calculating heat lost/gained and specific heat, solve the problems below.

1. A 10 g copper sample at 150°C is placed in a foam cup of water containing 50 mL of water at 30°C. After two minutes, the copper has given off some of its heat energy and the copper and water are at the same temperature. What is that temperature?

2. An unknown metal at 125°C with a mass of 10 g is placed in a foam cup with 40 mL of water at 20°C. After two minutes, the metal and water reach 22.75°C. What is the metal?

3. Three 20 g samples of iron, lead, and aluminum are placed on a concrete sidewalk on a sunny 38°C day. Before being placed in the sun, the samples were stored in a container at 20°C. Which metal absorbs the most thermal energy when heated? If the samples are then put in a freezer at 0°C for 1 hour, which metal loses the smallest amount of thermal energy?

Lab Six: Heat of Fusion and Vaporization

Learn About It

Energy is required to overcome the forces of attraction between atoms and molecules. To change a substance from a solid into a liquid requires energy to break the forces of attraction that keep the atoms or molecules in fixed positions.

The **heat of fusion** is the energy needed to change one gram of a substance, at its melting point, from a solid into a liquid without an increase in temperature.

To change from a liquid into a gas requires even more energy per gram to overcome the forces of attraction.

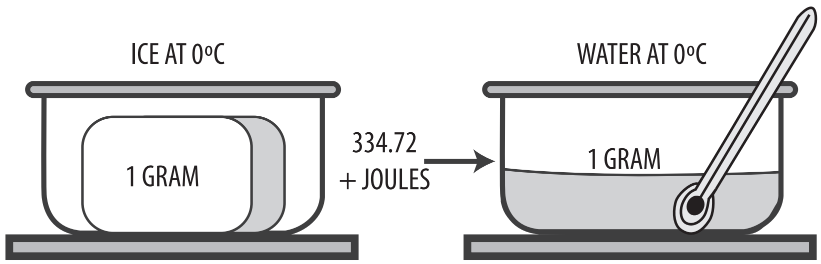


Figure 6.1. The heat of fusion of water is 334.72 J/g.

The **heat of vaporization** is the energy needed to change one gram of a substance, at its standard boiling point, from a liquid into a gas without an increase in temperature. Table 6.1 lists the melting and boiling points of several substances, and their heats of fusion and vaporization. Notice how much greater the heat of vaporization is than the heat of fusion for the substances listed.

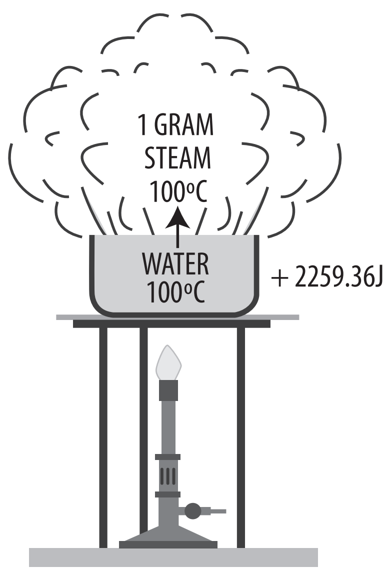


Figure 6.2. The heat of vaporization of water is 2259.36J/g.

Table 6.1: Melting and Boiling Points, Heat of Fusion and Vaporization

Substance	Melting Point (°C)	Heat of Fusion (J/g)	Boiling Point (°C)	Heat of Vaporization (J/g)
Copper	1,083	207.11	2,566	4,727.92
Ethyl Alcohol	-114	107.95	78	853.54
Gold	1,063	62.76	2,808	1719.62
Lead	327	23.01	1,750	857.72
Mercury	-39	11.30	357	297.06
Oxygen	-218	13.81	-183	213.38
Water	0	334.72	100	2259.36

Let's look at one familiar substance from the chart—a gram of water. What happens as thermal energy is added to ice at -20°C? Look at the graph below. As energy is added to the ice, its temperature rises from -20°C to 0°C.

At 0°C, H_20 molecules can exist as either a solid or a liquid—the only difference is the amount of thermal energy of each molecule. In a solid state, the molecules are in fixed positions—they do not have enough energy to break the bonds that hold them together. The forces of attraction between the molecules are great—you can stand on a frozen lake.

In a liquid state, however, the H_2O molecules are free to move about. As energy is added to the ice at 0°C, it changes into a liquid—water—at 0°C. All the added energy is used to change the water from a solid into a liquid—to break the bonds holding the molecules in a fixed position—not to increase the temperature. Notice how the line on the graph in Figure 6.3 remains horizontal during this phase. The energy added is the heat of fusion.

If we continue to add energy to the water, its temperature increases to 100°C. At this temperature, the H_2O molecules can exist either as a liquid or as a gas. As additional thermal energy is added to the sample, the energy is used to break free of the forces of attraction between the water molecules. Eventually, all the H_2O molecules have the energy to escape from the bonds of the other molecules. The additional energy changes the liquid into gas—vapor or steam—without producing an increase in temperature. This amount of energy is called the heat of vaporization.

Again, notice that the line on the graph is horizontal during this phase. Once all the water has been changed into steam, additional energy will increase the temperature of the steam.

The converse is true when going from a gas to a liquid, or liquid to a solid. The same amount of energy must be removed from a substance to change a gas to a liquid as was added to change the liquid to a gas.

The heat of fusion for water is 334.72 Joules per gram (J/g). It takes 334.72 Joules to change one gram of ice at 0°C into water at 0°C. Conversely, 334.72 Joules must be removed to change one gram of water at 0°C to ice at 0°C. The heat of vaporization for water is 2,259.36 J/g. This amount of energy must be added or removed from one gram of water or steam to change its state.

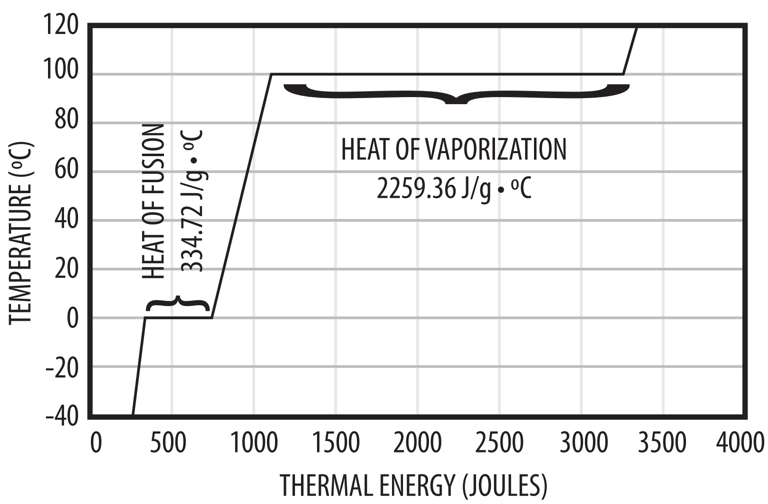


Figure 6.3. Temperature vs. Thermal Energy Added for Water

Think About It

Using Table 6.1, answer the following questions.

1. Is the melting point of a substance related to its heat of fusion?

2. Does it take more energy to change a solid into a liquid or a liquid into a gas? Why?

3. Which substance requires the most amount of thermal energy to change from a liquid into a gas?

4. Which substance requires the least amount of thermal energy to change from a solid into a liquid?

5. Do the heats of fusion and vaporization for substances seem to be related in any way?

Experiment With It

Let's investigate changes of state in the laboratory.

✪ Objectives

▪To determine the heat of vaporization of water.
▪To determine the heat of fusion of water.

Supplies and Preparation

Part I - Heat of Vaporization

▪Steam generator with hose to be operated by teacher
▪Foam cup with 100 mL (g) of distilled water at room temperature (about 22°C) and lid
▪Thermometer
▪Graduated cylinder
▪Bunsen burner or hot plate with ring stand to operate steam generator

Part II - Heat of Fusion

▪Ice cube made with distilled water
▪Foam cup with 100 mL (g) of warm distilled water (about 50°C) and lid
▪Thermometer
▪Graduated cylinder

☑ Procedure

Part I - Heat of Vaporization

1. Record the temperature of the 100 mL of water in the cup.
2. Cut a slit in the lid of the cup if one is not already present.
3. With your teacher's supervision, insert the hose from the steam generator into the foam cup through the slit in the lid.
4. Wait approximately eight minutes. (While you are waiting, proceed to Part II).
5. Remove hose and immediately record the temperature of the water in the cup.
6. Carefully pour the water from the cup into the graduated cylinder and record the volume of the water.

Part II - Heat of Fusion

1. Record the temperature of the 100 mL of warm water.
2. Place the ice cube in the cup of water and cover the cup with the lid.
3. After 30 seconds gently swirl the cup.
4. Within one minute, the ice cube should be completely melted. If not, wait until the ice is melted and record the final temperature of the water.
5. Carefully pour the water from the cup into the graduated cylinder and record the final volume of the water.

Record Data

	Initial Temperature	Final Temperature	Change in Temperature	Initial Volume	Final Volume	Change in Volume
Part I						
Part II						

Calculating and Interpreting

Part I - Heat of Vaporization

The heats of fusion and vaporization of water can be calculated as shown below:

Thermal energy gained or lost = (mass) x (change in temperature) x (specific heat)

$$q = m\Delta Ts$$

$$\text{Heat of vaporization} = \frac{\textit{thermal energy gained or lost}}{\textit{mass of steam}}$$

$$L_v = \frac{q}{m}$$

$$\text{Heat of fusion} = \frac{\textit{thermal energy gained or lost}}{\textit{mass of ice}}$$

$$L_f = \frac{q}{m}$$

Example: A cup with 100 mL water at 22°C was attached to a steam generator. After five minutes, the volume of water in the cup had increased to 106 mL. The temperature of the water had increased to 52°C. Calculate the heat of vaporization of the water.

1. Calculate the mass of the steam changed into water. Subtract 100 mL—the initial volume in the cup—from the final volume of 106 mL. Then, using the density of water as 1.00 g/mL, we have 6 g of water that started as steam and 100 g that started as water.
2. Calculate the amount of thermal energy gained by the water using its specific heat.

$$q = m\Delta Ts$$
$$q = 100g \text{ x } (52\ ℃ - 22\ ℃) \text{ x } 4.184\ J/g \cdot ℃)$$
$$q = 12{,}552\ J$$

3. Using the same relationship, calculate the thermal energy lost as the steam condenses.

$$q = m\Delta Ts$$
$$q = 6g \text{ x } (100\ ℃ - 52\ ℃) \text{ x } 4.184\ J/g \cdot ℃)$$
$$q = 1204.99\ J$$

The amount of thermal energy gained by the water is much more than the energy lost by the steam. This difference is the energy released as steam at 100°C is changed to water at 100°C.

4. Calculate the heat of vaporization:

$$L_v = \frac{q}{m}$$

$$L_v = \frac{(12{,}552 - 1{,}204.99)\ J}{6g} = \frac{11{,}347.01\ J}{6g} = 1{,}891.17\ J/g$$

The heat of vaporization in this experiment was calculated to be 1,811.17 J/g; according to Table 6.1 the accepted value is 2,259.36 J/g.

5. Using your data, calculate the heat of vaporization of water.

6. How does your calculated value compare to the accepted value in Table 6.1? Identify areas where some thermal energy may have been lost during the experiment that haven't been accounted for in your measurements. Are there any other sources of error?

Part II: Heat of Fusion

Example: An ice cube at 0°C was placed into a cup with 100 mL of water at 50°C. In one minute, the ice cube had melted. The volume of water in the cup was 120 mL and the temperature of the water was 30°C. Calculate the heat of fusion of water in this experiment.

1. Calculate the mass of the ice that changed into water. Subtract 100 mL—the initial volume in the cup—from the final volume of 120 mL. Then, using the density of water as 1.0 g/mL, we have 20 g of water that started as steam, and 100 g of water.

2. Calculate the amount of thermal energy lost by the water using its specific heat.

$$q = m\Delta Ts$$
$$q = 100g \times (50\ °C - 30\ °C) \times 4.184\ J/g \cdot °C)$$
$$q = 8{,}368\ J$$

3. Using the same relationship, calculate the thermal energy gained by the ice from the water.

$$q = m\Delta Ts$$
$$q = 20.00\ g \times (30.0\ °C - 0\ °C) \times 4.184\ J/g \cdot °C)$$
$$q = 2{,}510.4\ J$$

   The amount of thermal energy lost by the water is much more than the energy gained by the ice. This difference is the energy released as ice at 0°C is changed into water at 0°C.

4. Calculate the heat of vaporization:

$$L_f = \frac{q}{m}$$

$$L_f = \frac{(8{,}368 - 2{,}510.4)\ J}{20\ g} = \frac{5{,}857.6\ J}{20\ g} = 292.88\ J/g$$

The heat of fusion in this experiment was calculated to be 292.88 J/g; according to Table 6.1 the accepted value is 334.72 J/g.

5. Using your data, calculate the heat of fusion of water.

6. How does your calculated value compare to the accepted value in Table 6.1? Identify areas where some thermal energy may have been gained during the experiment that haven't been accounted for in your measurements. Are there any other sources of error?

Make Sure You Understand It

Using the formulas for calculating heat lost/gained and the heats of vaporization and fusion, solve the problems below.

1. A 20 g piece of ice at 0°C is placed in a foam container with 50 g of water at 100°C. After two minutes the ice has completely melted. What is the temperature of the water?

2. Two grams of steam at 100°C are put into a foam container with 50 g of water at 20°C. What is the final temperature of the water?

3. An unknown amount of steam is added to 90 g of water at 20°C. The final temperature of the water is 82°C. What is the mass of the steam added to the water?

4. A 100 g piece of ice at 0°C is placed into a foam container. A total of 20 g of steam at 100°C is introduced into the container. In three minutes, the ice has melted and all the steam has condensed into water. What is the temperature of the water in the container?

Glossary

atom	a tiny unit of matter made up of protons and neutrons in a small dense core, or nucleus, with a cloud of electrons surrounding the core
calorie	measure of thermal energy; one calorie is the amount of energy required to raise the temperature of one gram of water one Celsius degree
conduction	transfer of thermal energy by direct contact
convection	transfer of thermal energy by moving or flowing fluids (gases or liquids)
covalent compound	substance made of two or more elements that are held in place by atoms sharing one or more pairs of electrons between them
density	the amount of mass per unit of volume; the amount of space a certain mass occupies
electron	very tiny, negatively charged, subatomic particle that moves around the nucleus of the atom
element	most pure of all matter; all other matter is made of various combinations of elements
gas	state of matter with the highest thermal energy; gases have no definite shape or volume
heat of fusion	the amount of thermal energy required to change one gram of a solid substance into a liquid without increasing the temperature of the substance
heat of vaporization	the amount of thermal energy required to change one gram of a liquid substance into a gas without increasing the temperature of the substance
ionic compound	substance made of two or more elements in which electrons have been transferred from at least one atom to at least one other atom; the particles in an ionic compound are held in place by electrostatic force
insulator	transfers thermal energy slowly
Joule	SI unit for measuring energy; one Joule is equal to one kilogram•meter^2/second^2; one calorie is equal to 4.184 Joules
liquid	state of matter that has a definite volume but no definite shape; liquids are able to flow but do not necessarily fill their containers
mass	measure of the amount of matter in a substance; the SI unit for measuring mass is the kilogram (kg)
neutron	subatomic particle with no electric charge found in the nucleus of an atom
nucleus	the center of an atom composed of protons and neutrons where the vast majority of the atom's mass is found
pressure	amount of force per unit area; the amount of pressure exerted by a gas is a direct consequence of the number of collisions between gas particles and the walls of their container
proton	subatomic particle with a positive electric charge found in the nucleus of an atom
radiation	energy in the form of electromagnetic waves; visible light, infrared, and radio waves are all examples of radiant energy
solid	state of matter with the least amount of thermal energy; the particles of a solid are held in place, giving the solid both a definite shape as well as a definite volume
specific heat	amount of thermal energy required to raise the temperature of one gram of a substance by one Celsius degree
thermal energy	internal energy of a substance; the greater a substance's thermal energy, the more the particles in that substance are moving
volume	amount of space occupied by a certain number of particles of a substance

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