Section Preview of the

Laboratory Investigations Manual and **Student Book**

for

A Natural Approach to Chemistry, 2nd Edition

Chapter 3

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Name:________________________________Section:_________Date:_______

3A: Heat and Temperature

What is heat? Are heat and temperature the same thing?

Temperature tells you whether something is hot or cold. You know that adding heat can raise the temperature. In this investigation, we will learn precisely how heat and temperature are related. Heat and temperature are not the same thing, but they are closely related.

Materials

- Lab-Master system with temperature probe
- Three 8 oz foam cups
- One 16 oz foam cup
- Hot and cold water

Part 1: Measuring temperature

1. Set up the Lab-Master system with the temperature probe.
2. Fill a foam cup about half full of ice cold water, but be sure there is no solid ice in the cup.
3. Fill a second foam cup half full of hot water.
4. Measure the temperature in each cup. Record these values in Table 1 along with the time of day to the minute.

Table 1: Water temperatures

Clock time	Cold water temperature (°C)	Hot water temperature (°C)

5. Let the cups stand while you answer the questions in Part 2. Measure the temperatures again after about 5 minutes.

Part 2: What happens?

a. Consider that temperature describes a type of energy. Do you think hot or cold water has more of this kind of energy? Explain in one sentence why you think so.

b. How do you expect the temperature of the water in each of the two cups to change over time?

c. Describe the flow of energy that would cause the changes you predicted in question 2b.

d. Measure the temperatures in each cup and see whether the actual temperatures changed as you expected.

Part 3: Making heat flow

1. Prepare a small foam cup with 100 g of hot water.
2. Prepare a second small foam cup with 100 g of cold water.
3. Measure the temperatures in each cup just before you mix them in Step 4.
4. Mix the hot and cold water into the larger (empty) foam cup.
5. Stir the mixture with the temperature probe and quickly record the temperature of the mixture.

Measure the temperature of the hot and cold water just before mixing

Mix the hot and cold water and measure the final mixture temperature

Table 2: Water temperatures

Cold water temperature before mixing (°C)	Hot water temperature before mixing (°C)	Mixture temperature (°C)

Part 4: Stop and think

a. Which cup of water had more energy: the one with hot water or the one with cold water? Why?

b. What did you think the temperature of the mixture would be? Why?

c. If the system includes both cold and hot water, compare the energy of the system before mixing to the energy after mixing. You may ignore any energy going into the air or lost from friction.

Part 5: Analyzing the data

	Mass (g)		Temperature (°C)		gram degrees (g • °C)
Before mixing					
Cold Water		x	0 °C	=	
	+		+		+
Hot Water		x		=	
After mixing	=				=
Mixture		x	unknown	=	

a. Fill in the measurements for the hot and cold water, and calculate the gram-degrees of each.

b. For the mixture, fill in the total mass and total gram-degrees. Do not fill in the mixture temperatures!

c. What does the "unknown" box represent?

d. Calculate the unknown temperature from the mass and total gram-degrees.

e. How close did this value come to your actual measured temperature?

Part 6: A more complex experiment

1. Prepare three foam cups containing different amounts of hot and cold water. This time, measure the mass of water in each cup. Use at least 100 g of water of any temperature.
2. Measure and record the temperatures before mixing.
3. Mix the water in a large foam cup, stir well, and measure the final temperature.

What happens when you add three cups together at different temperatures?

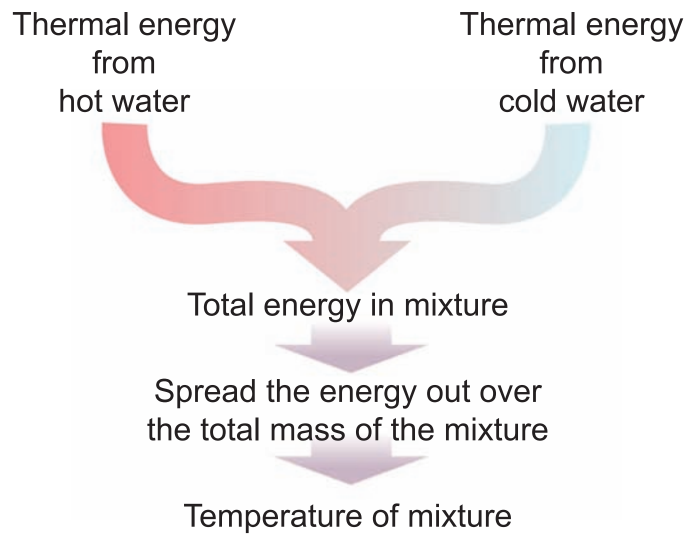
Table 3: Data for mixing unequal masses of water

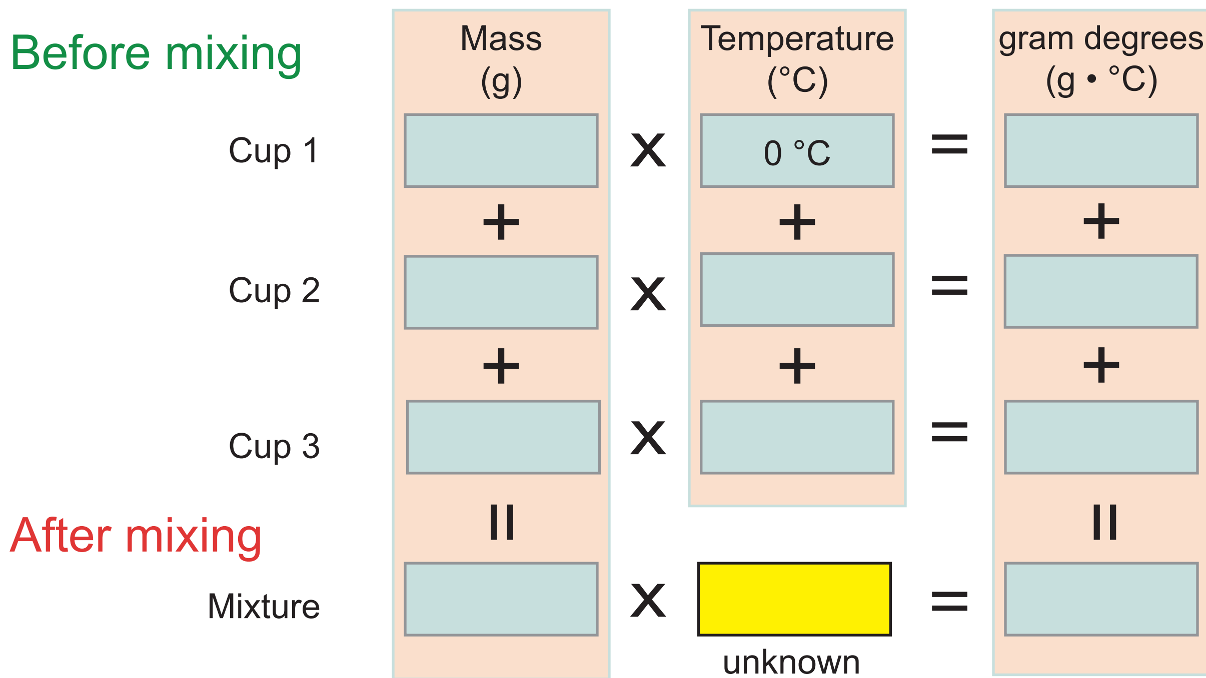
Sample	Mass (g)	Temperature (°C)
Cup #1		
Cup #2		
Cup #3		
Mixture		

Part 7: Doing the math

The thermal energy associated with a certain amount of mass is related to its temperature. The thermal energy in the water is proportional to the mass of water multiplied by the temperature. The energy is only proportional because different materials store different amounts of thermal energy, even at the same temperature.

For now, assume the "energy" is in units of gram-degrees, or g·°C. Here's how to think about the experiment in terms of energy:





a. Fill in the light blue boxes in the "Before mixing" section. Calculate the gram-degrees for both hot and cold water.

b. Add up the masses and the gram-degrees to get the total mass and gram-degrees for the mixture.

c. Solve the "After mixing" section to get the mixture temperature.

Part 8: Why did the calculation work?

a. Did the result of the experiment agree with your prediction? Discuss the meaning of "agree" in terms of the accuracy and precision of your experiment.

b. Assume you have 10 cups of water with different masses and temperatures. Describe a way to predict the temperature of the mixture if you know the masses and temperatures of the water in the cups.

c. Describe a situation where two objects have the same temperature but different amounts of energy.

d. Describe a situation where two objects have the same energy but different temperatures.

e. As you can see, "temperature" and "energy" have different meanings in everyday life and chemistry. Discuss these two meanings verbally with your lab group and write down your definitions.

3B: Specific Heat

If you know the temperature of something, how much energy does it have?

Two objects of the same mass and the same temperature can have different amounts of energy. This may seem odd, but it is true. Do the experiment and see for yourself.

Materials

- Ten 1/2” steel washers tied with a string
- Ice
- Water
- Mass balance
- Lab-Master with temperature probe
- Two 8 oz foam cups

Part 1: The experiment

1

Tare the balance to 0.0 with an empty cup. Then measure the mass of at least 100 g of steel washers in the cup.

2

Cover the steel washers with crushed ice and a little water. Measure the cold water temperature.

3

Measure out 100 g of hot water in a second foam cup.

4

Measure the hot water temperature just before mixing.

5

Remove the cold steel washers quickly and drop them in the hot water. Stir for a minute or so, then record the final temperature. *Be sure you do not get any ice in the water!*

Table 1: Temperature data for combining water and steel washers

Washer mass (g)	Washer temp. before mixing (°C)	Hot water mass (g)	Hot water temp. before mixing (°C)	Mixture temp. (°C)

Part 2: Analyzing the data

?	x			0°C	=		A
specific heat of steel (?)		mass of steel		initial temperature of steel		joules	
						+	
4.18 J/(g°C)	x				=		B
specific heat of water		mass of water		initial temperature of water		joules	
						=	
Before mixing					total energy		A+B

After mixing

4.18 J/(g°C)	x				=		C
specific heat of water		mass of water		mixture temperature		Subtract the energy left in the water from the total energy (A+B – C)	
	x				=		A+B-C
specific heat of steel		mass of steel		mixture temperature		energy in steel washers	

Part 3: Thinking and communicating about what you observed

a. Propose an explanation for why the temperature of the steel and water mixture did not come out halfway between the two original temperatures, even though you mixed equal masses of steel and water.

b. Now that you have a measurement of the specific heat, assume 0°C represents zero relative energy. (This means that we are measuring the energy relative to 0°C, not that the actual energy is zero.) How many joules of energy did the steel contribute to the mixture?

c. How many joules of energy did the water contribute to the mixture?

d. How good was the approximation we started with, that the steel contributed no energy to the mixture?

e. Go back and recalculate the total energy using the actual energy for the steel. Use the actual temperature you measured for the steel just before mixing.

f. Now calculate a new (more accurate) value for the specific heat of steel. How different is this new value from the one you had?

Part 4: The specific heat of steel

Specific heat is a property of a material that describes how temperature and thermal energy are related. For example, the specific heat of water is 4.18 J/g·°C. That means it takes 4.18 J of energy to raise the temperature of 1 g of water by 1°C.

The total amount of thermal energy stored in a material depends on three things:

- specific heat,
- mass, and
- temperature.

The relationship among these is:

Formula | **Thermal energy**

Mass (g)

Thermal energy (J) $$E = mc_p \overbrace{(T_2 - T_1)}^{\text{Temperature change (°C)}}$$

Specific heat [J/(g•°C)]

Example: How much energy is needed to raise the temperature of 10 g of water by 1°C?

$$E = (10\text{ g}) \times [4.18\text{ J/(g•°C)}] \times (1\text{°C})$$
$$= 41.8\text{ J}$$

Part 5: Thinking and communicating about what you observed

a. Suppose you add 100 J of energy to 50 g of water. By how much will the temperature of the water increase?

b. Describe a situation where two objects have the same mass and the same temperature but different amounts of thermal energy.

c. Describe a situation where two objects have the same mass and the same amount of thermal energy but different temperatures.

d. The specific heat of gold is 0.13 J/g·°C. Suppose you add 100 g of gold at 100°C to 100 g of water at 0°C. Is the mixture temperature likely to be

i. closer to 0°C than to 50°C,

ii. closer to 100°C than to 50°C, or

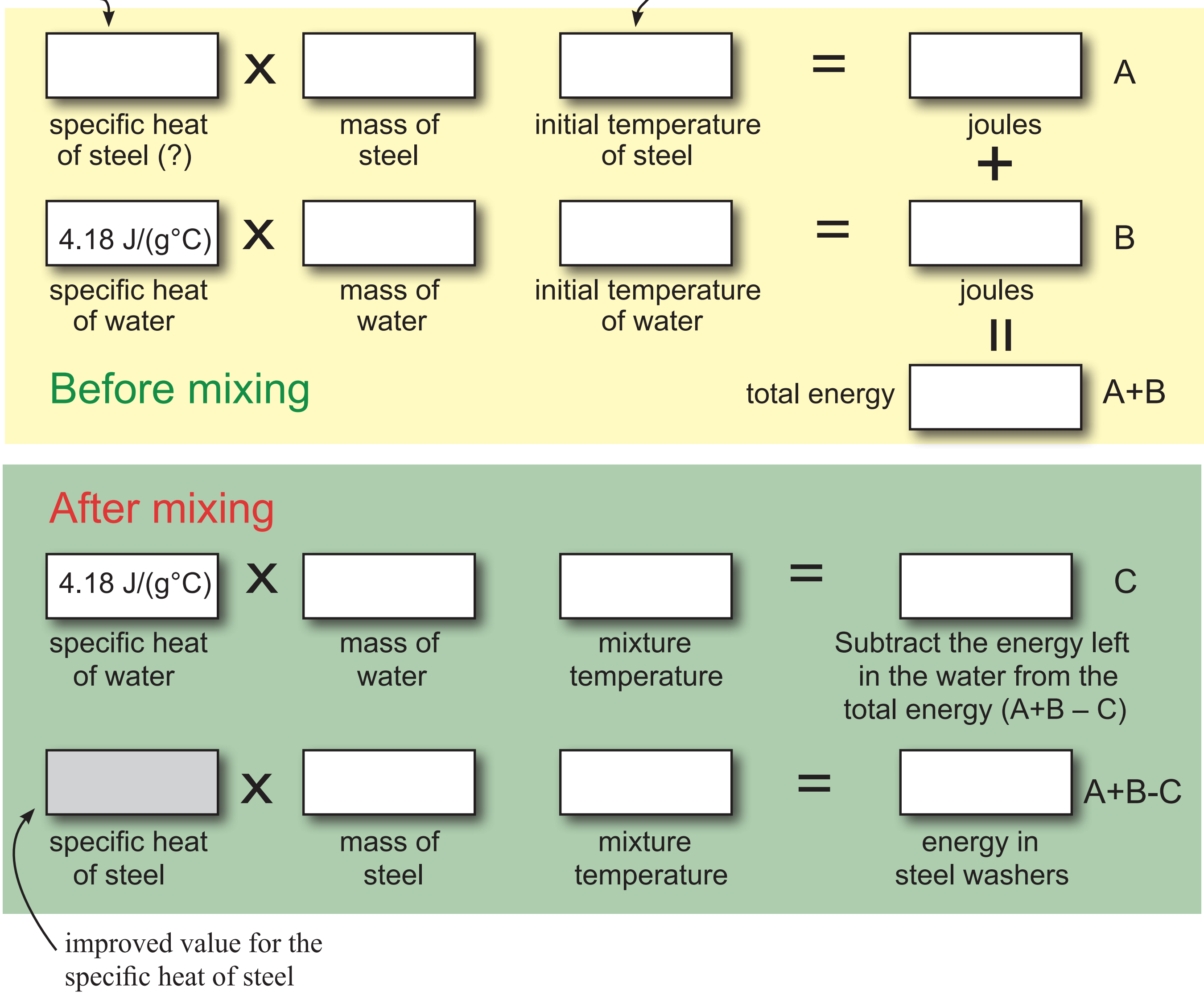
iii. around 50°C?

e. Explain why the mixture will be at this temperature.

Part 6: Were we sloppy with our math?

We ignored the energy contained in the steel when it was cold. We assumed that the temperature was 0°C: therefore it did not matter what the specific heat was. With our old assumption, the contribution would still be zero. Of course, this is not exactly true! You likely found that the actual temperature of the steel was higher than zero.

a. The diagram below is similar to the one used on page 28. Here we will enter the value of specific heat you calculated in Part 2. Here we will enter the actual temperature of the steel.



b. Calculate the energy contribution from the steel.

c. Is the energy from the steel large or small compared to the energy in the water?

d. Complete the calculation using your value for the specific heat of steel. By the time you get to the bottom of the chart, your new answer will be a better estimate for the specific heat of steel.

e. Is the new estimate close to the first one? Why do you think that is?

The technique you just learned is called successive approximation. At first we made an assumption that allowed us to get to an answer. We then used our answer to check our assumption and arrived at an even better answer.

3C: Heat Flow and Thermal Equilibrium

Why and how does heat flow?

Everybody knows that an ice-cold drink warms up if you leave it in a warm room. The drink gets warmer, so heat energy must be flowing into it. When the drink gets to the same temperature as the room, however, it stops warming up. What slows down and eventually stops the flow of heat energy from the room to the drink?

Materials

- Lab-Master with temperature probe and heater
- One-hole rubber stopper with insulation ring
- Cold tap water
- 25 mm test tube

Part 1: Temperature and heat

a. Set up the Lab-Master with the heater and a temperature probe.

b. Fill the test tube with room-temperature water until the level is just above the lip of the test tube heater. The temperature probe should be below the surface of the water.

c. Set the heater to 50°C.

d. Record the temperature every 30 s for 5 minutes in the appropriate column of Table 1 (see p. 32). Make sure that you continually stir the water with the temperature probe while taking measurements.

e. Repeat Steps b–d, but with the heater temperature set to 100°C.

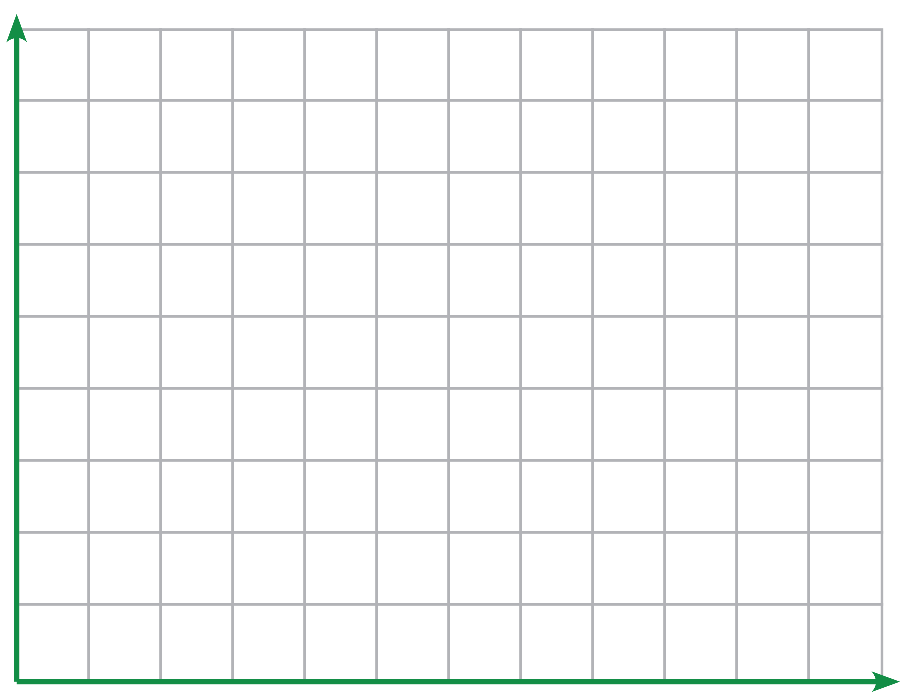
f. Plot the data from Table 1 on the graph paper below the table. Use a different color for each set of data. A good graph must have labeled axes, and each data point must be presented clearly. [Here for example you may label the horizontal axis "Time (s)" and the vertical axis "Water temperature (°C)."]

Heater

Set the heater to 50°C for the first part of the experiment

TABLE 1: Water temperature (°C)

Time (min)	Heater set to 50°C	Heater set to 100°C
0.0		
0.5		
1.0		
1.5		
2.0		
2.5		
3.0		
3.5		
4.0		
4.5		
5.0		



Identify three differences and three similarities between the two plots, and write them below.

Similarities:

1.

2.

3.

Differences:

1.

2.

3.

Part 2: Thinking about it

a. Why is it important not to have too much water in the test tube?

b. Why do you have to stir the water while heating it?

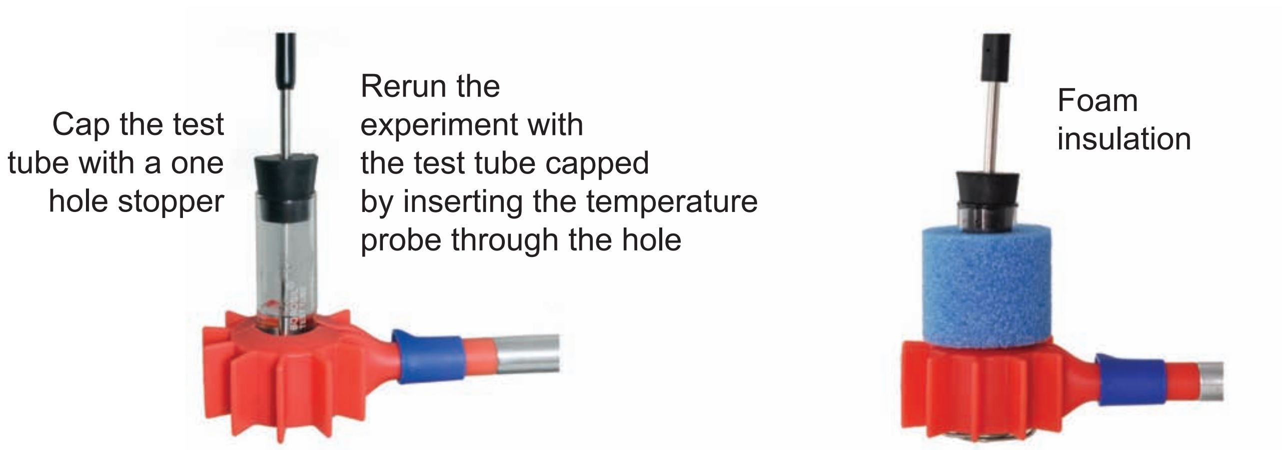
c. What was the highest reading you measured on the temperature probe?

d. Describe the temperature versus time graph. What are the differences between the two plots?

e. Was heat transferred from the heater to the water at the same rate the entire time? Was the energy transfer reduced or even stopped at some point? What evidence do you have to support your claim? (Hint: Look at the power display on the lower right of the Lab-Master screen.)

f. Was heat transferred from the heater to the water at the same rate the entire time?

Part 3: Heat flow



1. Add a foam insulation ring to the test tube.
2. Put the temperature probe through a one-hole stopper so it sits below the surface of the water.
3. Observe the temperature for a few minutes while the heater is set to 50°C. Does the final temperature get higher than before, or does it stay about the same?

Part 4: Thinking about what you observed

a. What was the purpose of insulating the test tube? Think about heat as energy and where the energy goes.

b. A covered pot boils much faster than an open pot. Discuss why that is and how it relates to why putting the cap on the test tube changed the maximum temperature of the water.

c. Explain why the water became warmer, even though the temperature of the heater stayed the same.

d. Explain why the *power* of the heater starts high, but drops to a very low value shortly after.

e. *Thermal equilibrium* is the situation when all temperatures have become equal. No heat flows in thermal equilibrium. Is your test tube in thermal equilibrium or not? Why do you think so? This is a hard question! Discuss it with your class and your lab group, then write up a short answer.

Name:________________________________Section:_________Date:_______

3D: Heat of Fusion

Why doesn't the temperature change as ice melts?

When you add heat to a sample of ice and water, the temperature doesn't change. You can see ice melt as the mixture becomes more liquid. However, as long as there is still some solid ice, the temperature stays constant. Why?

Materials

- Lab-Master with temperature probe
- Ice
- Hot water (60–80°C)
- Mass balance
- Two 8 oz foam cups

Part 1: The experiment

Attach the temperature probe.

1

2

Measure the mass of about 100 g of ice into a foam cup.

In a second cup measure about 130 g of hot water.

3

Measure and record the hot water temperature *just* before adding it to the cold ice.

4

5

Mix the hot water with the ice. Measure and record the mixture temperature after about a minute of *gentle* stirring and *after* the ice has melted.

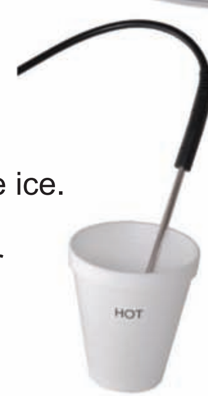


Table 1: Temperature data for combining water and ice

Ice mass (g)	Ice temp. before mixing (°C)	Hot water mass (g)	Hot water temp. before mixing (°C)	Mixture temp. (°C)
	0			

Part 2: Analyzing the data

[ ]	x	[ ]		0°C	=	[ ]	A
specific heat of ice?		mass of ice		initial temperature of ice		joules	
						+	
4.18 J/(g°C)	x	[ ]		[ ]	=	[ ]	B
specific heat of water		mass of water		initial temperature of water		joules	
						=	
Before mixing				total energy		[ ]	A+B

After mixing							
4.18 J/(g°C)	x	[ ]		[ ]	=	[ ]	C
specific heat of water		total mass of water		mixture temperature		Subtract the thermal energy in the water from the total energy (A+B – C)	
		[ ]	x	[ ]	=	[ ]	A+B-C
		heat of fusion of ice		mass of ice		energy needed to melt ice	

Part 3: Thinking about what you observed

a. Suggest an explanation for why the temperature of the water did *not* end up halfway between cold and hot, even though you mixed equal masses.

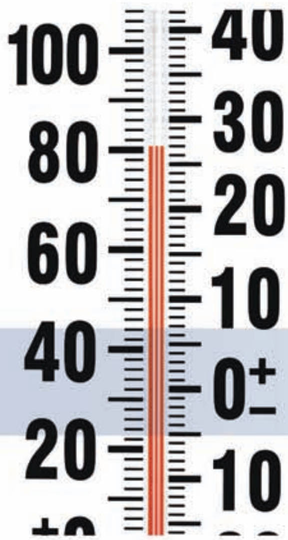
b. We have assumed that a temperature of 0°C represents zero thermal energy by measuring relative to the reference point of 0°C. How many joules of thermal energy did the solid ice contribute to the mixture?

c. How many joules of thermal energy did the water contribute to the mixture?

d. How does your value for the heat of fusion of ice compare to the accepted value?

e. All substances that undergo phase changes have a heat of fusion. How do other substances compare to water? Research the heat of fusion for at least four other substances.

f. Suggest an explanation for why the heat of fusion of ice is similar to or different from the other substances you chose.



CHAPTER 3
Temperature, Energy, and Heat

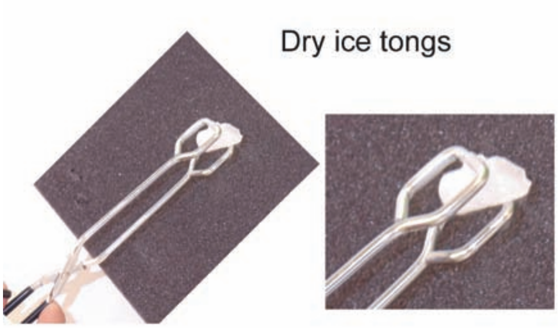
What is temperature?
Why is temperature important in chemistry?
How is energy related to temperature?
Is there a difference between heat and temperature?

The coldest place on Earth is in Antarctica, where geologists recorded a temperature of –89°C (–129°F) on July 21, 1983. This is so bitterly cold that an unprotected human would perish in minutes. The Antarctic penguins and sea birds spend most of their lives on the water. The largest purely land-dwelling creature living in Antarctica is an insect no bigger than your fingernail. This insect produces a chemical called *glycerol* in its body, which is a natural antifreeze!

The hottest temperature recorded on Earth was 58°C (136°F) in Libya on September 13, 1922. This narrowly beat the previous record of 57°C set in Death Valley, California on July 10, 1913. Fortunately, the average temperature on the surface of our planet is 15°C, or 59°F. This is perfect for living things, as it is comfortably within the range for which water is a liquid. Above 100°C, water boils and important chemicals break down, or react quickly, as they do in cooking. Below 0°C, pure water is a solid and the chemistry necessary for life cannot take place in a solid.

When an object is *hot*, what is different about its matter compared to the same object when it is *cold*? This chapter explains the concept of temperature and what it really means.

Temperature and state of matter

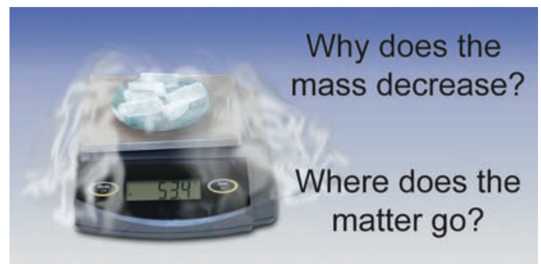


Dry ice is an amazing substance. Use some heavy gloves or tongs to hold a small piece of it. *Never* touch dry ice with bare skin. *Always* wear proper protection when working with dry ice.

- *How would you describe this material? Is it solid, liquid, or gas?*
- *What is the "smoke" you see? Is it hot or cold?*
- *Does the "smoke" go up or down?*

Place the chunk of dry ice on a balance. Record the mass every 30 s for 2 min.

- *Is dry ice matter? How do you know?*
- *Is the "smoke" matter? How do you know?*
- *What happens to the "lost" mass as the dry ice chunk gets smaller? Where did the mass go?*



Place a small piece of dry ice in a beaker of water. Observe what happens. You can touch the vapor with your finger; however, keep your skin away from any solid dry ice that may be floating on the water.

- *What is the vapor that you see?*
- *Is the vapor hot or cold?*
- *Does the vapor flow up or down? Why?*

Take about 100 mL of warm water and add a couple squirts of dish detergent. Stir the mixture to thoroughly mix the soap into the warm water. Pour the mixture into a 250 mL graduated cylinder or other tall, thin container.

Add a small piece of dry ice to the mixture and stand back!

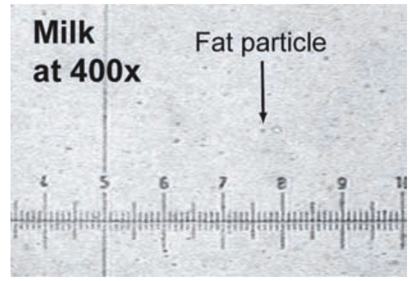
- *Explain what caused the explosion of soap suds.*
- *Do the bubbles that form float up or down? Why?*

At room temperature, or 22°C (72°F), air is a gas and water is a liquid. Heat water up to 100°C and it becomes a gas called steam. The water molecules still exist as H_2O. However, the molecules have so much energy that they fly apart from each other to make steam. Carbon dioxide (CO_2) is a gas at room temperature, but it becomes a solid at –79°C. Dry ice is frozen CO_2 gas. It does not melt into a liquid but *sublimates* directly back into gas as the temperature increases.

3.1 Temperature

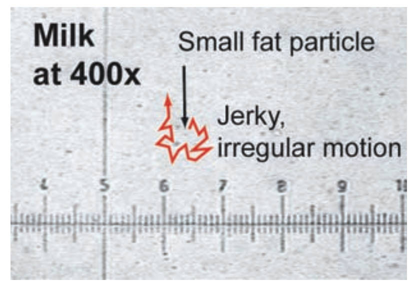
Particles of matter are in constant motion

Milk looks like a uniform liquid, but it really isn't. Under the microscope, you can see tiny globules of fat suspended in water. These fat particles are in constant motion! The particles jitter around in a very agitated way and *never slow down or stop*. This is strange when you think about it. Motion requires energy. What possible source of energy keeps the fat particles dancing around?



Brownian motion

If you look more carefully at the very smallest particles, you see they don't move smoothly as they would if they were floating. Instead, they move in a jerky, irregular way. The jerky movement of a very small particle in water is called **Brownian motion** and is a direct consequence of atoms and temperature. In 1905, Albert Einstein proved that matter was made of atoms by explaining Brownian motion.



Why Brownian motion occurs

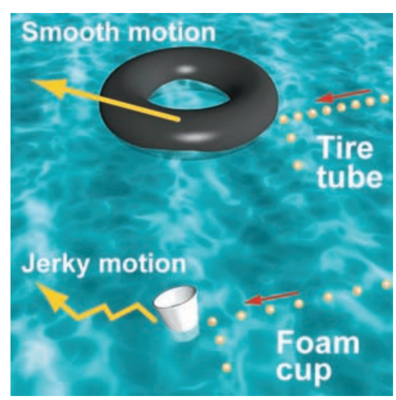
Brownian motion occurs for two important reasons:

1. Matter (including water) is made of atoms.
2. Atoms and molecules are in constant, agitated motion.

If the fat particle is *very* small, collisions with single molecules of water are visible because the mass of a water molecule is not that much smaller. The constant motion of individual water molecules causes Brownian motion.

A human-sized example

Imagine throwing marbles at both a tire tube and a foam cup floating in a pool. The motion of the tube is smooth because each marble has a lot less mass than the tube. The foam cup jerks under the impact of each marble, like the fat particle in Brownian motion. This is because the mass of the cup is not much greater than the mass of a single marble. Brownian motion proves that matter exists in discrete atoms and molecules. It also proves that, *at room temperature*, atoms and molecules are in constant, agitated motion.



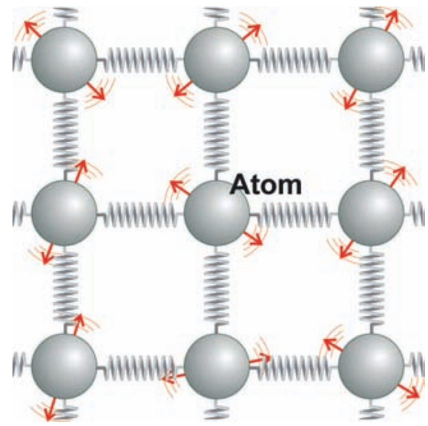
Chemistry terms

Brownian motion: the erratic, jerky movement of tiny particles suspended in a fluid caused by the random impacts of individual molecules in thermal motion.

The explanation of temperature

Temperature and energy

Brownian motion demonstrates that atoms are constantly jiggling around. Jiggling implies motion, and motion always involves energy. Ordinary matter, even sitting still, contains energy in the microscopic motion of its constituent atoms. This embedded energy is what **temperature** measures. Temperature is a measure of the average **kinetic energy** of individual atoms (or molecules). When the temperature goes up, the energy of motion increases and the atoms move more vigorously.

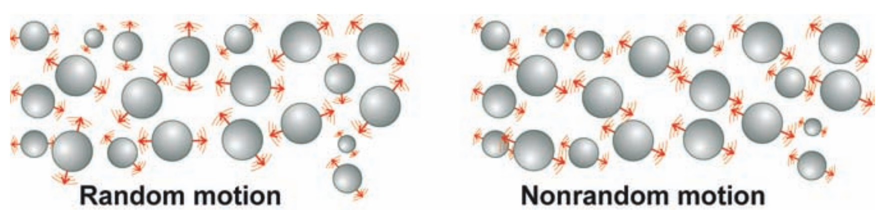


Average motion

If atoms are always moving, how can a grain of sand just stand still? The answer to this question has three parts:

1. Atoms are tiny and there are trillions of them in a grain of sand.
2. The motion resulting from temperature is fast, but very short, typically the width of a single atom in liquids and solids (but more in gases).
3. Constant collisions cause the motion of individual atoms to be *random*.

Atoms are typically so close together that they constantly bang into each other and change direction. At any given moment, there are as many atoms bouncing one way as there are the other way. The *average* speed of the whole group is zero. However, no individual atom is ever standing still.



Temperature and random motion

Random motion is motion that is scattered equally in all directions. Temperature only measures the energy in the *random* motion of atoms and molecules. Temperature is not affected by energy resulting from motion of the whole group. That is why throwing a rock does not make it hotter. When you throw a rock, you give each atom in the rock the same average motion because all the atoms move together.

Chemistry terms

kinetic energy: the energy of motion.

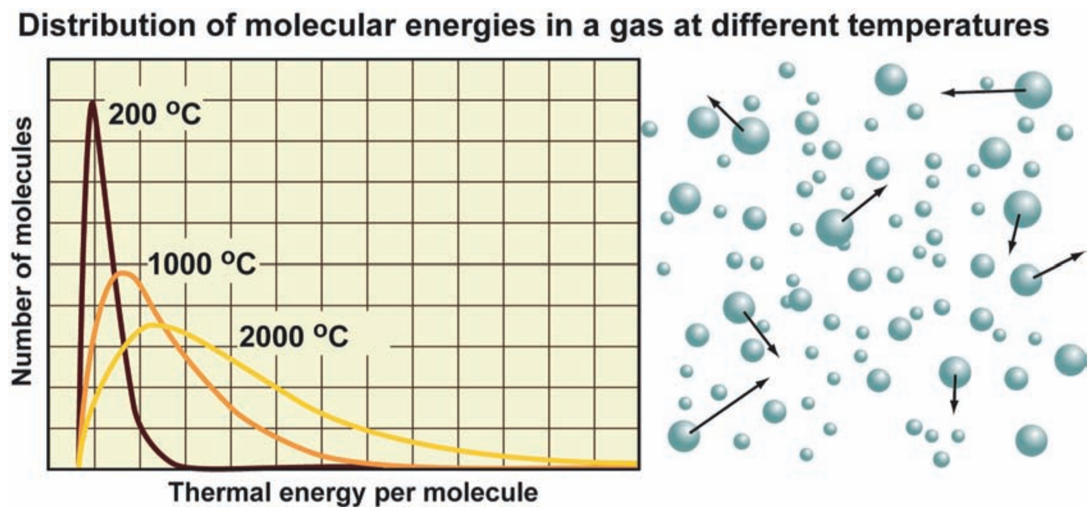
temperature: a measure of the average kinetic energy of atoms or molecules.

random: scattered equally among all possible choices with no organized pattern.

Temperature is an average

Molecules at a given temperature have a range of energies

In a given quantity of matter there are trillions of atoms. Some of them have more energy than the average and some have less energy than the average. This is easiest to see in a gas where molecules are not bound together. The graph below shows thermal energy per molecule on the horizontal axis and the number of molecules on the vertical axis. The graph shows that molecules have a wide range of energy. A few have a lot more energy than average and a few have a lot less.



The energy changes with temperature

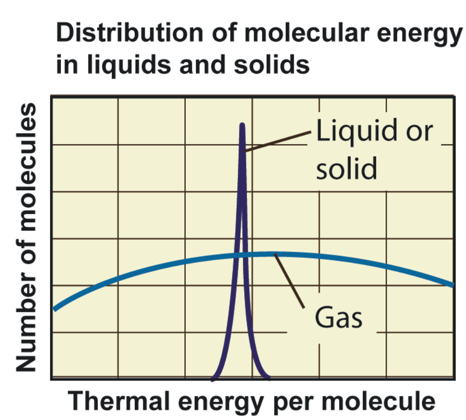
As the temperature increases, the graph changes in two ways:

1. The average energy of the molecules increases, so the peak of the graph shifts to the right.
2. The graph gets wider, so the range of energy of the molecules increases.

Molecular energy in liquids and solids

In a gas, the average distance between molecules is thousands of times the size of a single molecule. Molecules in a gas interact when they happen to hit each other. This weak interaction is why the graph of molecular energy of a gas spreads out so much.

In a liquid or a solid, the molecules are very close together, separated by less than their own size. Molecules are in continuous motion, making them bump into each other constantly. The constant bumping creates a strong interaction between molecules. In turn, the strong interaction means the energy curve is much narrower. Any molecule with significantly *more* energy quickly bounces into its neighbors and loses some. Any molecule with significantly *less* energy is bumped and jostled by its neighbors and speeds up a little, gaining energy.



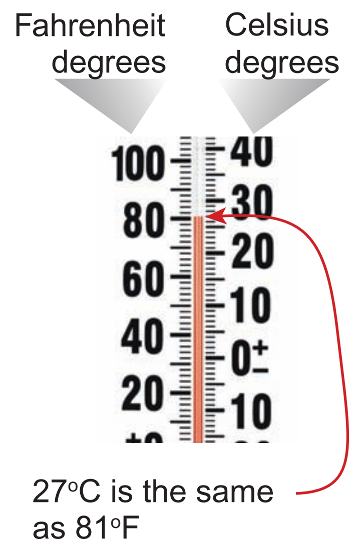
Temperature scales

Why is temperature important?

Humans can sense relative temperature through nerves in the skin. Ice feels cold because its temperature is *lower* than the temperature of your skin. A cup of coffee feels hot because its temperature is *higher* than the temperature of your skin. Temperature is important to life because most of the internal processes in living organisms only work in a narrow range of temperatures.

The Fahrenheit scale

There are two commonly used temperature scales: the Fahrenheit scale and the Celsius scale. On the **Fahrenheit scale**, water freezes at 32 degrees and boils at 212 degrees. There are 180 Fahrenheit degrees between the freezing point and the boiling point of water. Temperature in the United States is commonly measured in Fahrenheit. For example, 81°F is the temperature of a warm summer day.



The Celsius scale

The **Celsius scale** divides the difference between the freezing and boiling points of water into 100 degrees (instead of 180). Water freezes at 0°C and boils at 100°C. Most science and engineering temperature measurement is in Celsius. Most countries use the Celsius scale for all descriptions of temperature, including daily weather reports. The same 81°F warm summer day is 27°C.

Other countries use Celsius degrees for everyday temperature

If you travel to other countries, you will want to know the difference between the two temperature scales. A temperature of 21°C in Paris, France, is a pleasant spring day, suitable for shorts and a T-shirt. A temperature of 21°F would be like a cold January day in Minneapolis, Minnesota! You would need a heavy winter coat, gloves, and a hat to be comfortable outdoors in 21°F. The United States is one of the few countries still using the Fahrenheit scale.

Chemistry terms

Fahrenheit scale: a temperature scale with 180 degrees between the freezing and boiling points of water; water freezes at 32°F and boils at 212°F.

Celsius scale: a temperature scale with 100 degrees between the freezing and boiling points of water; water freezes at 0°C and boils at 100°C.

Converting Fahrenheit to Celsius

Differences between the two temperature scales

Fahrenheit degrees are smaller than Celsius degrees. For example a temperature change of 9°F is the same as a temperature difference of 5°C. Besides the different size of a degree between the two scales, there is also a 32°F *offset* (0°C is 32°F) that must be considered when converting from one scale to the other. Therefore, when we convert from °C to °F, we must first scale the °C by 9/5 and then add 32 degrees. The formulas for the conversions are as follows:

Converting between Celsius and Fahrenheit

$$T_{\text{Fahrenheit}} = \frac{9}{5} T_{\text{Celsius}} + 32 \qquad T_{\text{Celsius}} = \frac{5}{9}\left(T_{\text{Fahrenheit}} - 32\right)$$

Fahrenheit to Celsius

To convert from Fahrenheit to Celsius, subtract 32, then multiply by 5/9. Subtracting 32 is necessary because water freezes at 32°F and 0°C. The factor of 5/9 is applied because the Celsius degree is larger than the Fahrenheit degree.

Solved problem

What temperature in Celsius is the same as 100°F?

Asked: *Temperature in °C*

Given: *100°F*

Relationships: $T_C = \frac{5}{9}(T_F - 32)$

Solve: $T_C = \frac{5}{9}(100 - 32) = \frac{5}{9}(68) = 37.8$

Answer: *100°F is the same temperature as 37.8°C.*

Celsius to Fahrenheit

To convert from Celsius to Fahrenheit, multiply by 9/5 then add 32. The factor of 9/5 accounts for the different size degrees and the 32 corrects for the zero offset.

Solved problem

What is the Fahrenheit equivalent of 15°C?

Asked: *Temperature in °F*

Given: *15°C*

Relationships: $T_F = \frac{9}{5} T_C + 32$

Solve: $T_F = \frac{9}{5}(15) + 32 = 27 + (32) = 59°F$

Answer: *15°C is the same temperature as 59°F.*

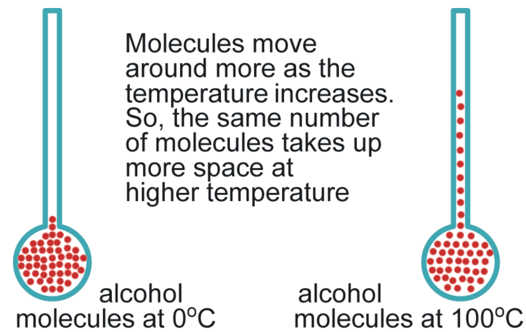
Measuring temperature

The human

Scientists use instruments to measure temperature because the human sense of temperature is not very accurate. You can feel when something is warm or cold, but you cannot determine its exact temperature. For example, if you walk into a 65°F room from being outside on a winter day, the room feels warm. The same room will feel cool if you come in from outside on a hot summer day.

Thermometers

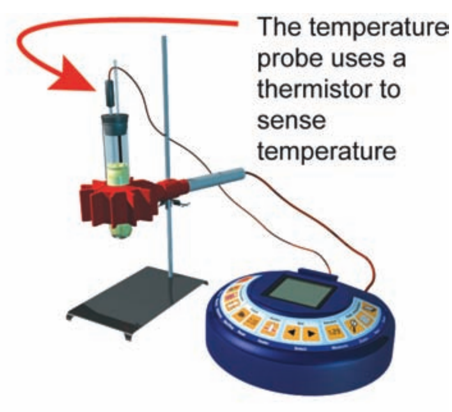
Our sense of temperature is thus very relative. We can easily compare the temperature of things that we touch, but we can't easily say what the actual temperature is. In science, we try to perform absolute measurements. This means that we need to know the actual value of the temperature. A **thermometer** is an instrument that measures temperature. The common alcohol thermometer uses the expansion of liquid alcohol. As the temperature increases, the alcohol expands and rises up a long, thin tube. The temperature is measured by the height the alcohol rises. The thermometer can indicate small changes in temperature because the bulb at the bottom has a much larger volume than the tube.



How thermometers work

All thermometers are based on a physical property that changes with temperature. A **thermistor** is a temperature sensor that changes its electrical properties as the temperature changes. A **thermocouple** is another electrical sensor that measures temperature. The temperature probe you use in your laboratory experiments uses one of these two sensor types.

Some chemicals change color at different temperatures. One type of aquarium thermometer has a stripe that changes color when the water is too hot or too cold.



Chemistry terms

thermometer: an instrument that measures temperature.

thermistor: an electronic sensor that measures temperature by detecting changes in electrical resistance.

thermocouple: an electronic sensor made of two different metals that detects a temperature-dependent voltage across them.

Absolute zero

Absolute zero

As the temperature gets lower, atoms have less and less thermal energy. This implies that a temperature can be reached at which the thermal energy becomes *zero*. This lowest possible temperature, called **absolute zero**, is where atoms have essentially zero thermal energy. Absolute zero is – 273°C or – 459°F. Think of absolute zero as the temperature where even atoms are completely frozen, like ice, with no motion. *It is not possible to have a temperature lower than absolute zero, or – 273.15°C.*

There are no temperatures below absolute zero

Technically, we believe atoms can never stop moving *completely*. Even at absolute zero some tiny amount of thermal energy is left. For our purposes, however, this "zero point" energy might as well be truly zero because the rules of quantum physics prevent the energy from ever going any lower. Figuring out what happens when atoms are cooled to absolute zero is an active area of research.

The Kelvin scale

The **Kelvin** temperature scale is useful for many scientific calculations because it starts at absolute zero. For example, the pressure in a gas depends on how fast the atoms are moving. The Kelvin scale is used because it measures the actual thermal energy of atoms. A temperature in Celsius measures only the *relative* energy, relative to zero Celsius.

Converting to kelvins

The kelvin (K) unit of temperature is the same size as the Celsius unit. Add 273 to the temperature in Celsius to get the temperature in kelvins. For example, a temperature of 21°C is equal to 294 K (21 + 273). Note that the word "degree" is not used with the kelvin scale. A temperature of 300 K is read as "300 kelvin" and is abbreviated without the degree symbol.

$$T_{\text{Kelvin}} = T_{\text{Celsius}} + 273$$

Solved problem

Convert 27°C into kelvins.

Asked: *Temperature in kelvin*

Given: *27°C*

Relationships: $T_{\text{Kelvin}} = T_{\text{Celsius}} + 273$

Solve: $T_K = 27 + 273 = 300\ K$

Answer: *300 K is the same temperature as 27°C.*

Chemistry terms

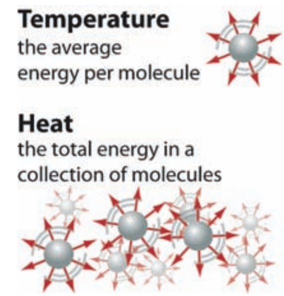
absolute zero: the lowest possible temperature, at which the energy of molecular motion is essentially zero, or as close to zero as allowed by quantum theory.

Kelvin scale: a temperature scale that starts at absolute zero and has the same unit intervals as the Celsius scale: $T_{\text{Kelvin}} = T_{\text{Celsius}} + 273$.

3.2 Heat and Thermal Energy

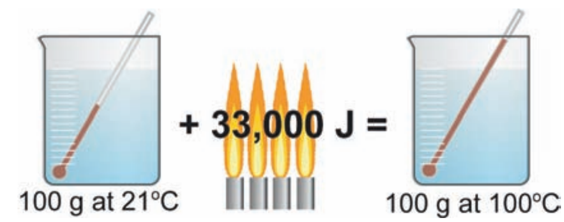
Thermal energy is also called heat

Heat is another word for thermal energy. On the molecular level, thermal energy is the random kinetic energy of a collection of atoms and/or molecules. On a macroscopic level, thermal energy is the energy stored in matter that is *proportional to temperature*. To change the temperature of matter, you need to add or subtract heat. You add heat to warm your house in the winter. If you want to cool your house in summer, you remove heat from it.



Joules

The **joule** (J) is the fundamental SI unit of energy and heat. A joule is a fairly small unit of energy. Heating 100 mL (100 g) of water from room temperature to boiling requires about 33,000 J of heat.

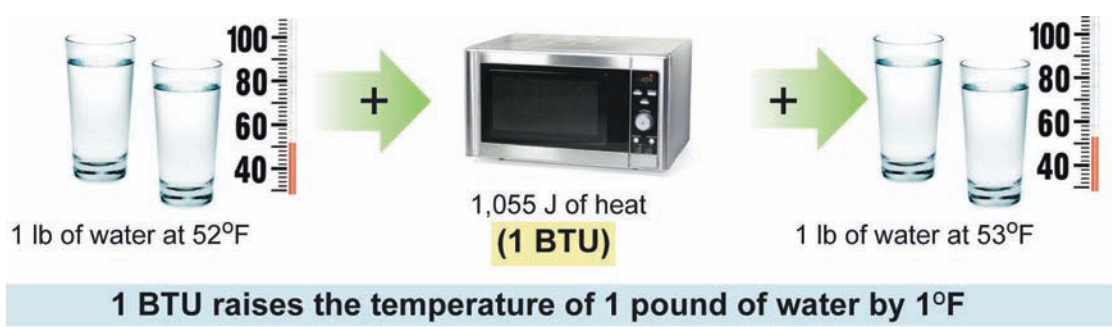


Calories

The **calorie** is an older unit of heat used in chemistry. One calorie is the amount of heat required to raise the temperature of one gram of water by one degree Celsius. There are 4.184 J in one calorie so a calorie represents more energy than a joule. To make things more confusing, the Calories listed in foods (with capital "C") are really *kilocalories*. One Calorie (Cal) = 1,000 calories (cal) = 4,184 joules (J).

British thermal units (BTU)

The air conditioner or furnace in your house is rated in **British thermal units (BTU)**. One BTU is the amount of heat required to raise the temperature of one pound of water by one degree Fahrenheit. A typical home-heating furnace can produce 10,000 to 100,000 BTU per hour. One BTU equals 1,055 J.



heat: thermal energy, energy resulting from temperature; the total energy in random molecular motion contained in matter.

joule: the fundamental SI unit of energy (and heat).

calorie: an older unit of heat; 1 calorie = 4.184 joules.

British thermal unit (BTU): a large unit of heat. 1 BTU = 1,055 joules.

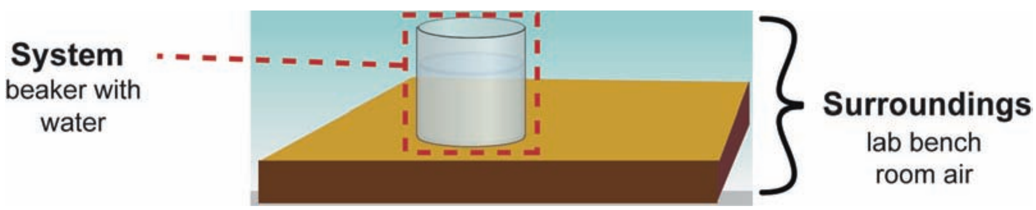
Thermodynamics and systems

Heat flows from hot to cold (2nd law)

Imagine standing outside on a cold winter day without a coat. Your body feels cold because you are losing energy to your surroundings. That's because your body temperature is higher than the air temperature. Heat (energy) moves from high temperature to lower temperature. This common-sense rule is called the **second law of thermodynamics**.

Choosing a system

To see how heat moves and what it does, we must look carefully at what we are studying. To a chemist, the part of the "universe" under study is called the **system**. In the no-coat example, the system is your body; everything else is called the *surroundings*. By carefully defining the system and the surroundings, you can keep track of how much energy there is and where it goes. For example, if the system is defined as a beaker containing 200 mL of water at 60°C, then the surroundings would be the lab bench on which the beaker is resting and the air surrounding the beaker.



Why heat flow is important

The flow of heat energy is important to virtually everything that occurs in both nature and technology. The weather on Earth is a giant heat recycler among the oceans, the land, and the atmosphere. A car's engine converts heat from burning gasoline to energy of motion of the car.

Open, closed, and isolated systems

Systems can be open, closed, or isolated. In an **open system**, matter and energy can be exchanged between the system and the surroundings. In a **closed system**, only energy can be exchanged between the system and surroundings. An example of a closed system would be a covered coffee cup. Heat can still be exchanged or lost between the air and the coffee, but no matter can be exchanged. In the third type of system, which is **isolated**, the coffee would be inside a perfectly insulated container from which no matter or heat could escape.

second law of thermodynamics: law of nature that states that energy (heat) spontaneously flows from higher temperature to lower temperature.
system: a group of interacting objects and effects that are selected for investigation.
open system: a system which can exchange matter and energy with the surroundings.
closed system: a system which can exchange only energy with the surroundings.
isolated system: a system in which neither matter nor energy can be exchanged with the surroundings.

The first law of thermodynamics: energy conservation

The first law of thermodynamics

The **first law of thermodynamics** is the law of energy conservation: All the energy lost by one system must be gained by another system. Consider an experiment where you mix 100 g each of hot and cold water. How can the temperature of the mixture be predicted?

Assume the two cups are an isolated system; no energy or mass is allowed to cross the boundary. That means that the energy after mixing is the same as it was before mixing.

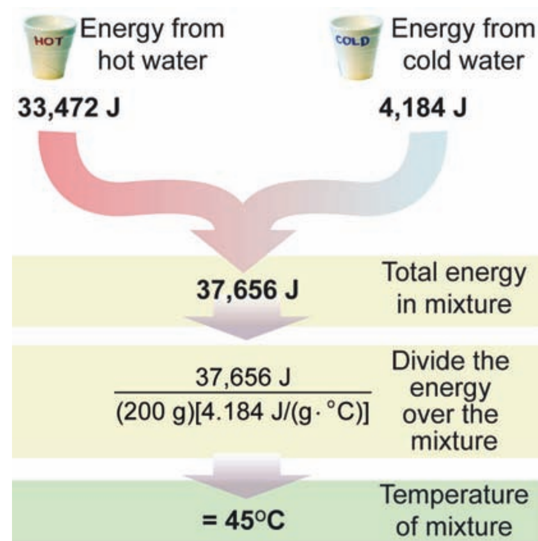
Energy depends on mass and temperature

The amount of heat energy in a quantity of matter, such as a cup of water, is proportional to mass and temperature. More mass means more energy at any temperature, so the higher the temperature, the higher the energy.

Tracing the energy

To find the final mixture temperature, you need to know the total energy of the system. To make things easy, start with 0°C as a reference point. Each gram of water stores 4.184 J of energy for every degree Celsius.

1. The hot water contributes 33,472 J.
2. The cold water contributes 4,184 J.
3. The mixture has a total energy of 37,656 J (hot + cold).
4. The 200 g of water contains 37,656 J of heat at 45°C.



How to solve the problem

How did we solve this problem? We added up all the energy in the system. We then analyzed where the energy went. In this case, all the energy went to changing the temperature of the mixture. The problem was solved by calculating what temperature the mixture *had to be* to contain all the energy in the system.

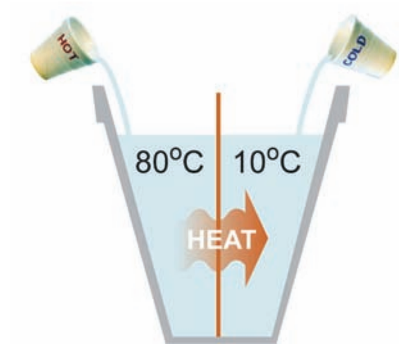
Chemistry terms

first law of thermodynamics: law of nature that states that energy can neither be created nor destroyed, and thus the total energy in an isolated system remains constant; all the energy lost by one system must be gained by the surroundings or another system.

Thermal equilibrium

What stops heat from flowing?

Suppose we pour the hot and cold water into a divided cup so they can't mix, but heat can still flow. Heat flows from the hot water to the cold water. The hot water gets cooler as it loses heat. The cold water gets warmer as it gains the same amount of heat. *When does the heat flow stop?* Why doesn't the hot water keep losing heat until it gets *colder* than the cold water? Obviously, this does not happen; heat flow eventually stops.

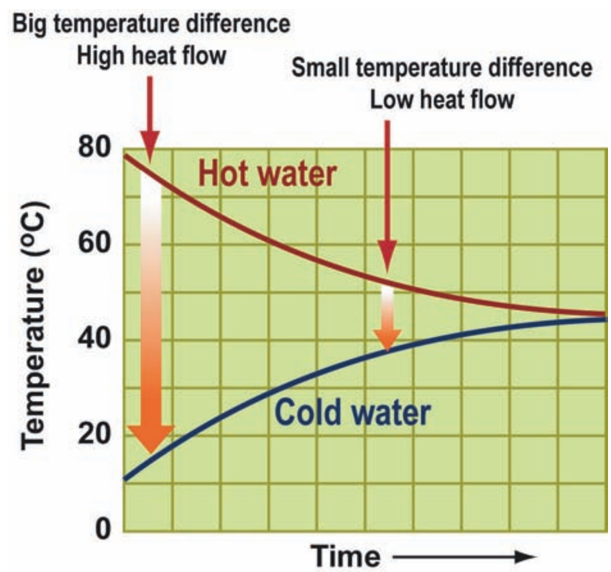


Thermal equilibrium

Two bodies are in **thermal equilibrium** when they have the same temperature. In thermal equilibrium, no heat flows because the temperatures are the same. In nature, heat always flows from hot to cold until thermal equilibrium is reached. Making the divider out of insulating foam only slows the process down. Heat still flows, only slower. The hot and cold water take *longer* to reach thermal equilibrium, but the end result is the same.

The rate of heat flow

The rate at which heat flows drops off gradually as the temperatures get closer together. A lot of heat flows quickly when the temperature difference is large, but as the temperature difference gets smaller, the rate of heat flowing gets proportionally smaller, too. As two objects approach thermal equilibrium, the rate of heat flow between them becomes zero because both reach the same temperature.



Heat transfer in living things

Heat flow is necessary for life because biological processes release energy. Your body regulates its temperature through the constant flow of heat. The inside of your body averages 37°C. Humans are most comfortable when the air is about 25°C because the rate of heat flow out of the body matches the rate at which the body generates heat. If the air is 10°C, you get cold because heat flows out of your body too rapidly. If the air is 40°C, you feel hot because your body cannot get rid of its excess internal heat.

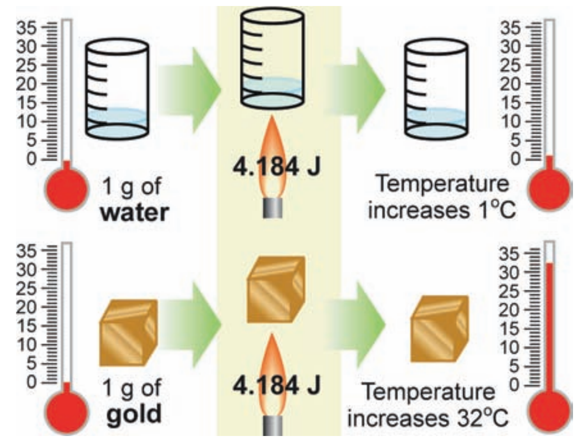
Chemistry terms

thermal equilibrium: a condition in which two systems have the same temperature and no net heat flows from one system to the other.

Specific heat

Differences in materials

The same amount of heat causes a different change in temperature in different materials. For example, it takes 4.184 J of heat energy to raise the temper-ature of one gram of water by one degree Celsius. If you add the same quantity of heat to one gram of gold, the temperature goes up by 32.4°C! The rise in temperature is not the same because different materials have different abilities to store thermal energy.



Specific heat

The **specific heat** in defined as the quantity of energy it takes to raise the temperature of one gram of a material by one degree Celsius. Water is an important example; the specific heat of water is 4.184 J/(g·°C). It takes 4.184 J to raise the temperature of one gram of water by one degree Celsius. The specific heat of gold is 0.129 J/(g·°C). It only takes 0.129 J to raise the temperature of one g of gold one degree Celsius. The specific heat of water is 32 times higher than it is for gold.

TABLE 3.1. Specific heat of some common substances

Material	Specific heat [J/(g·°C)]	Material	Specific heat [J/(g·°C)]
Air at 1 atm	1.006	Oil	1.900
Water	4.184	Concrete	0.880
Aluminum	0.900	Glass	0.800
Steel	0.470	Gold	0.129
Silver	0.235	Wood	2.500

The heat equation

The heat equation is used to calculate how much hear (Q) it takes to make a temperature change ($\Delta T = T_2 - T_1$) in a mass (m) of material with specific heat (c_p).

Heat equation

Heat (J) Mass (g) Temperature change (°C)

$$Q = mc_p \Delta T$$

Specific heat [J/(g·°C)]

Chemistry terms

specific heat: the quantity of heat energy, usually measured in J/(g·°C), it takes per gram of a certain material to raise the temperature by one degree Celsius.

Why specific heat varies

Different substances have different specific heats

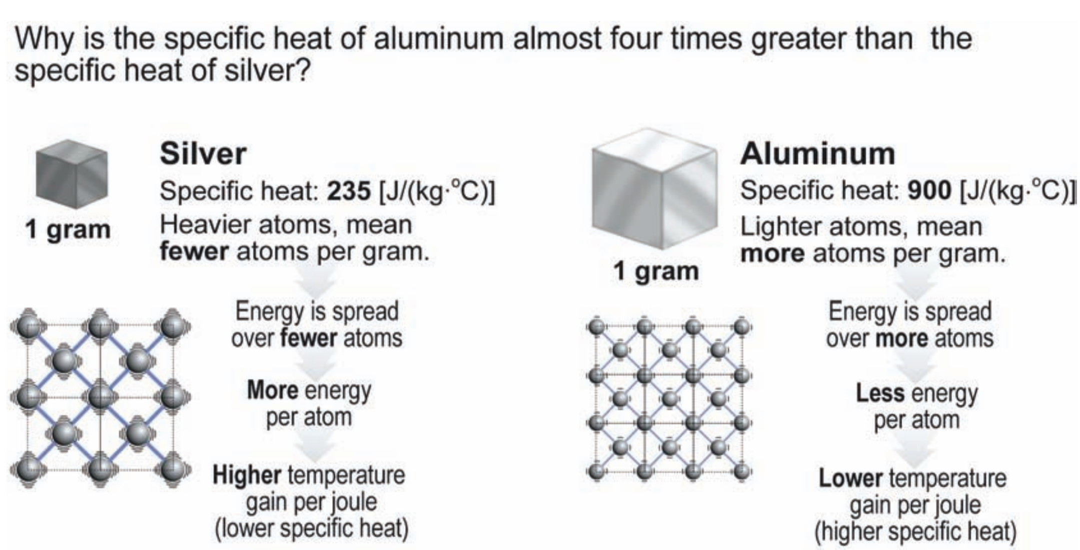
Substances have a wide range of specific heats. Pure metals, such as gold, tend to have low specific heats. Molecular substances, such as water and oil, tend to have higher specific heats. Specific heat varies for many reasons.

Molecular substances can absorb energy in ways that don't increase the temperature, such as internal motion of the atoms within a molecule. This is because bonds between atoms are not rigid rods, like the diagrams show. Bonds behave more like flexible springs that can bend and stretch. Typically, *only the motion of whole molecules affects the temperature*. The motion of atoms within a molecule, however, may not affect temperature. When energy is absorbed in ways other than motion of the whole molecule, temperature goes up less and the specific heat increases.

Stronger forces between molecules mean it takes more energy to cause a single molecule to move a given amount. This makes the specific heat higher. In general, strong bonds between molecules raise the specific heat because they limit the thermal motion of individual molecules (or atoms).

Why specific heat varies

Materials with heavy atoms or large molecules have lower specific heat compared with materials with lighter atoms. This is because temperature represents the energy per atom. Heavy atoms imply fewer atoms per kilogram. Energy that is divided among fewer atoms means more energy per atom, and therefore more temperature change. Silver's specific heat is 235 J/(kg·°C) and aluminum's specific heat is 900 J/(kg·°C). One gram of silver has fewer atoms than a gram of aluminum because silver atoms are heavier than aluminum atoms. When heat is added, each atom of silver gets more energy than each atom of aluminum because there are fewer silver atoms in a gram. Since the energy per atom is greater, the temperature increase in the silver is also greater.



Calculating temperature and heat

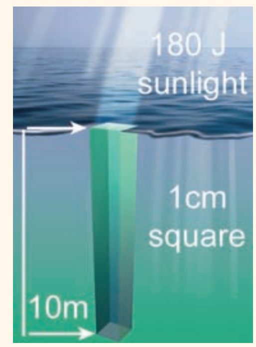
Most specific heat problems include one or both of the following two calculations:

1. Calculate the temperature change from a given heat input.
2. Calculate how much heat is needed to reach a specified temperature.

Solved problem

On a sunny day, each square centimeter of the ocean absorbs 180 J of energy from the Sun each hour. Assume all the heat is absorbed in the first 10 m, which has a mass of 1,000 g. How much does the water temperature increase?

This is a lengthy question! Before going on, discuss with your classmates what is asked and what is given.For example, what does "absorb energy from the sun" means?



Asked: *Temperature increase, ΔT*

Given: *Energy input = 180 J,*
mass of water = 1,000 g

Relationships: $Q = mc_p \Delta T$

Solve: $180\ J = (1{,}000\ g) \times\ 4.184\ [J/(g{\cdot}°C)] \times\ \Delta T$
$\Delta T = 180 \div 4{,}184 = 0.04°C$

Answer: *The temperature rise in the top 10 m of the ocean is 0.04°C per hour of full sunlight.*

Solved problem

A metal–working process needs to heat steel from room temperature (20°C) to 2,000°C. If the mass of steel is 100 g, how much heat is required?

Asked: *Amount of heat required*

Given: *100 g of steel,* $c_p = 0.470\ J/(g{\cdot}°C)$
temperature difference ΔT is
$\Delta T = 2{,}000°C – 20°C = 1{,}980°C$

Relationships: $Q = m\ c_p\ \Delta T$

Solve: $Q = (100\ g) \times\ 0.470\ [J/(g{\cdot}°C)] \times\ 1{,}980°C = 93{,}060\ J$

Answer: *It takes 93,060 joules to raise the temperature of 100 g of steel to by 1,980°C, assuming no heat gets lost during the process (which is not a very good assumption!).*

Applying the first law of thermodynamics

Heat flow out of a system equals heat absorbed by surroundings

When you leave a cup of hot coffee on the table, it cools down. Heat flows from the hot coffee to the cooler air in the room. The thermal energy of the hot coffee is decreased, and the thermal energy of the air is increased by the same amount. Everything balances; the increase in thermal energy of the air is exactly the same as the decrease in thermal energy of the coffee.

Solving problems with the first law of thermodynamics

The first law of thermodynamics is useful because it allows you to solve problems quickly and with minimal work. In real life, a hot coffee cup heats up the nearby air, which rises and makes currents that change how fast the heat flows. It would be difficult to calculate precisely how much heat flows every second. By looking at the energy at the start and then after the coffee has completely cooled down, we can get a pretty good answer without knowing the exact details of exactly how the heat went from cup to air.

TABLE 3.2. Applying the first law

Step 1	Step 2	Step 3
Identify the system and all the sources and uses of energy in the system before the change.	Set the total energy of the system after the change equal to the total energy before the change.	Account for all the uses of energy after the change and solve the problem.

A mass of 300 g of water at 80°C cools down to 20°C. Assume all the heat from the water is absorbed by 100 m^3 of air (a small room) with a mass of 100,000 g. What is the temperature change in the air?

Asked: *Temperature change in °C*

Given: *300 g of water [c_p = 4.184J/(g°C)], change of 60°C (80°C – 20°C), and 100,000 g of air [c_p = 1.006 J/(g·°C)]*

Relationships: $Q = m\,c_p(T_2 - T_1)$

Solve: *Energy lost by the water = (300 g)[4.184 J/(g·°C)](60°C) = 75,312 J*

Energy gained by the air = energy lost by water = 75,312 J

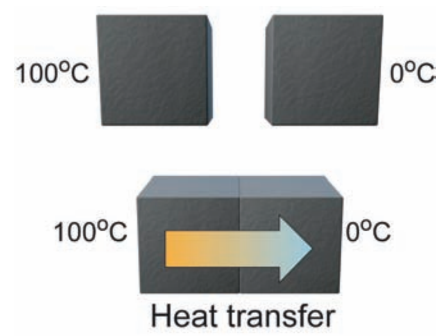
$$\text{Temperature change of air} = \frac{75{,}312\ \text{J}}{100{,}000\ \text{g} \times 1.006\ \text{J/(g·°C)}} = 0.75\text{°C}$$

Answer: *The air in the room gets warmer by about 0.75°C.*

Heat transfer

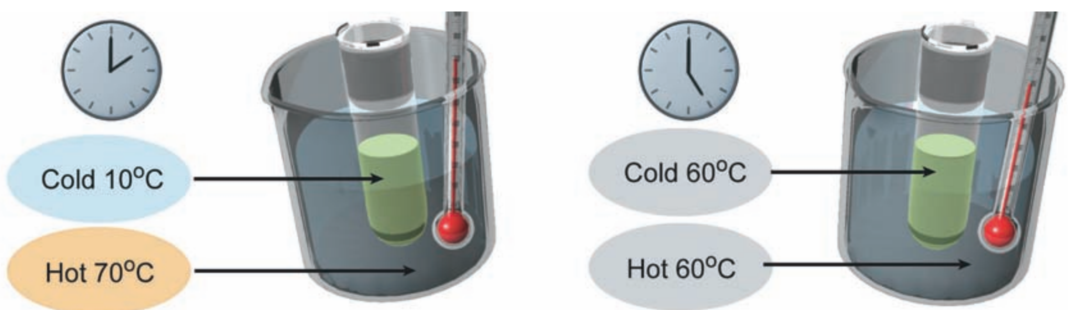
Conduction

Conduction is the transfer of heat *through* materials by the direct contact of matter. Imagine bringing in contact two pieces of material—one hot at 100°C and the other cold at 0°C. When the two pieces come in contact with each other, heat begins to flow. The heat flows from the hot piece to the cold piece. The cold piece warms up and the hot piece cools down. There will be a net heat flow until the two pieces are at the same temperature.



Thermal conductors

Glass and metal are **thermal conductors.** Think about holding a test tube of hot water. Thin glass conducts heat relatively well so heat flows rapidly through the glass from the hot water to your skin. You cannot hold on for long because the heat raises your skin temperature quickly.



Thermal insulators

A **thermal insulator** is a material that conducts heat poorly. Styrofoam is a good example. You can comfortably hold a hot test pipe surrounded by a centimeter of foam insulation. Heat flows very slowly through the foam so that the temperature of your hand does not rise very much. Plastic foam gets its insulating ability by trapping spaces of air. We use thermal insulators to maintain temperature differences without allowing much heat to flow.

The ability to conduct heat depends on many factors

All materials conduct heat at some rate. Solids usually are better heat conductors than liquids, and liquids are better conductors than gases. The ability to conduct heat often depends more on the structure of a material than on the material itself. For example, solid glass is a thermal conductor when it is made into windows. When glass is spun into fine fibers and made into insulation (fiberglass), the combination of glass fibers and trapped air makes a thermal insulator.

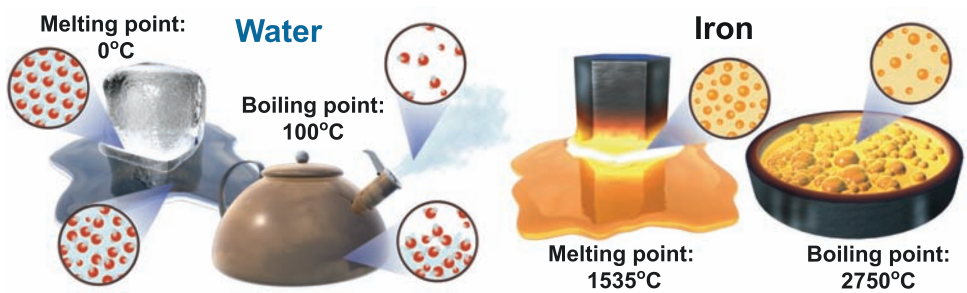
Chemistry terms

conduction: the flow of heat energy through the direct contact of matter.
thermal conductor: a material that conducts heat easily.
thermal insulator: a material that resists the flow of heat.

3.3 Phase Changes

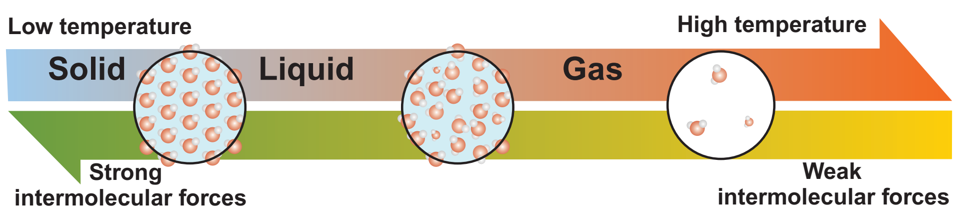
Solid, liquid, and gas

During a **phase change**, a substance rearranges the order of its particles (atoms or molecules). Examples of phase change include melting (solid to liquid) and boiling (liquid to gas). The most familiar example is water, which freezes at 0°C and boils at 100°C. All substances experience phase change, even metals such as iron. Iron melts into liquid at 1,535°C and boils into a gas at 2,750°C.



Why phase changes occur

Phase changes come from the competition between temperature and attractive intermolecular forces. On one side of this competition are intermolecular forces, which tend to attract molecules together into rigid structures. On the opposite side is the effect of temperature. Thermal energy is disruptive. Molecules with lots of thermal energy shake back and forth so much they cannot stay in an orderly structure, like a solid.



Melting point

The **melting point** is the temperature at which a substance changes from solid to liquid. Melting occurs when the thermal energy of individual atoms becomes comparable to the attractive force between atoms.

Different materials have different melting points because their intermolecular forces have different strengths. Water melts at 0°C (32°F). Iron melts at a much higher temperature, about 1,535°C (2,795°F). The difference in melting points tells us that the attractive force between iron atoms is much greater than the attractive force between water molecules.

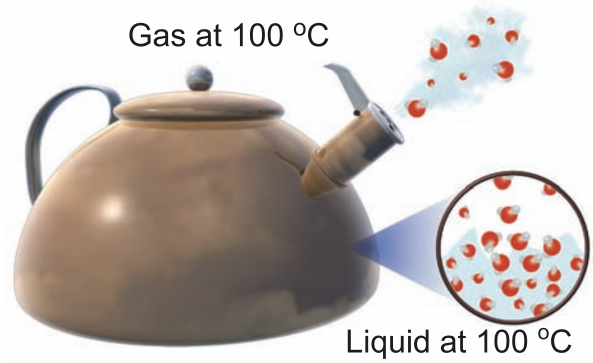
phase change: conversion of the organization of molecules in a substance without changing the individual molecules themselves, such as changing from solid to liquid or liquid to gas.

melting point: the temperature at which a substance changes phase from solid to liquid; for example, the melting point of water is 0°C.

Boiling and the heat of vaporization

Boiling point

The **boiling point** is the temperature at which the phase changes from liquid to gas. In a gas, all the bonds between one atom and its neighbors are completely broken. Water boils at 100°C (212°F) at a pressure of one atmosphere. The steam above the teapot is made of water molecules in the gas phase. Iron boils at 2,750°C (4,982°F).



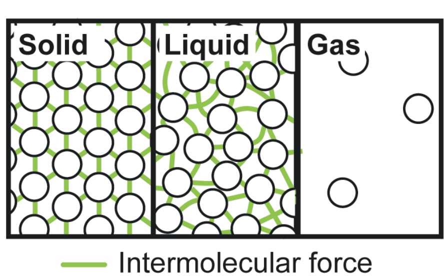
Heat of vaporization

It takes energy for a molecule to completely break its bonds with its neighbors and go from liquid to gas. The **heat of vaporization** is the amount of energy it takes to convert one gram of liquid to one gram of gas at the boiling point. Some representative values are given in Table 3.3. A quick calculation shows it takes 2,256,000 J to turn a kilogram of boiling water into steam! This explains why stoves require so much electric power.

TABLE 3.3. Heat of vaporization for some common substances

Substance	Heat of vaporization, ΔH_v (J/g)
Water	2,256
Alcohol	854
Liquid nitrogen	201
Iron	6,265
Silver	2,336

Comparing heats of vaporization and fusion



The heat of vaporization is much greater than the heat of fusion because breaking bonds between atoms or molecules takes much more energy than exchanging bonds. In a liquid, molecules move around by exchanging bonds with neighboring molecules. The energy needed to break one bond is recovered when the molecule forms a new bond with its neighbor.

Chemistry terms

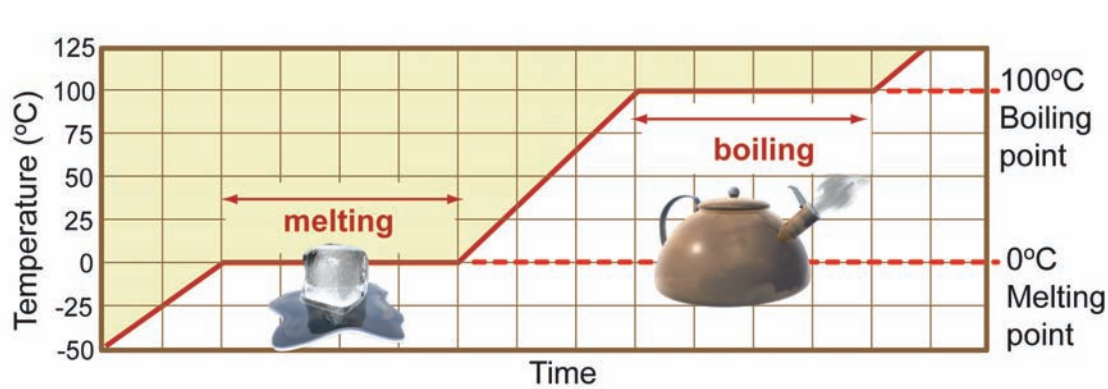
heat of vaporization: the energy required to change the phase of one gram of a material from liquid to gas or gas to liquid at constant temperature and constant pressure at the boiling point.

boiling point: the temperature at which a substance changes phase from liquid to gas which is the temperature at which the vapor pressure is equal to the atmospheric pressure. For example, the boiling point of water at atmospheric pressure is 100°C.

The heat of fusion

Temperature does not always rise when heat is added!

Think about heating a block of ice with an initial temperature of –20°C. As heat energy is added, the ice warms. However, once the ice reaches 0°C, *the temperature stops increasing!* This is because ice is melting to form liquid water. As heat is added, more ice becomes water but the temperature stays the same. This can easily be observed with an ordinary thermometer in a *well-stirred* experiment.



It takes energy to change phase

It takes energy to separate molecules in a solid and make them liquid, *even if the temperature stays the same*. The temperature does not change while the ice is melting because the added energy goes into loosening the bonds between neighboring water molecules. When thermal energy is added or subtracted from a material, either the temperature changes, or the phase changes, but usually not both at the same time. Once all the ice has become liquid, the temperature starts to rise again as more heat is added.

The heat of fusion

The **heat of fusion** is the amount of energy it takes to change one gram of material from solid to liquid or vice versa. Table 3.4 gives some values for the heat of fusion (ΔH_f) for common materials. Note how large the values are. It takes 335,000 J of energy to turn one kilogram of ice into liquid water.

TABLE 3.4. Heat of fusion for some common substances

Substance	Heat of fusion, ΔH_f (J/g)
Water	335
Aluminum	321
Iron	267
Paraffin (wax)	200–220

Chemistry terms

heat of fusion: the energy required to change the phase of one gram of a material from liquid to solid or solid to liquid at constant temperature and constant pressure at the melting point.

Solving phase change problems

Condensation releases the heat of fusion

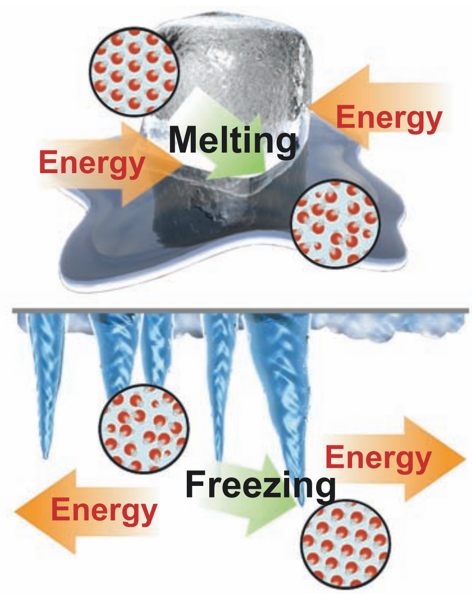
Energy can be released or absorbed in phase changes. For example, condensation occurs when a gas turns back into a liquid again. When water vapor condenses, it gives up the heat of vaporization! Steam at 100°C causes severe burns because each gram of water vapor delivers 2,256 J of energy to your skin from the heat of vaporization. By comparison, a gram of water at 100°C only releases 314 J as it cools down to skin temperature (25°C).

Solving phase change problems

Many problems involving phase changes can be solved by thinking about where the energy is absorbed or released.

1. Energy can be absorbed or released during phase changes.
2. Energy can be absorbed or released by changes in temperature.

The important idea is to realize that energy is conserved. Any energy used to change phase is no longer available to change temperature, and vice versa. For example, if 100 J of energy is applied to a system and 40 J is used to raise the temperature, then only 60 J is available for changing phase.



Solved problem

Ice cubes with a temperature of –25°C are used to cool off a glass of punch. Which absorbs more heat: warming up the ice or melting the ice into water? The specific heat of ice is 2.0 J/(g·°C).

Asked: *Which absorbs more heat, warming ice by 25°C or melting it?*

Given: *The ice starts at –25°C. The specific heat of ice is 2.0 J/(g·°C).* ΔH_f *(water) = 335 J*

Relationships: $Q = mc_p(T_2 - T_1)$ and $Q = m\Delta H_f$

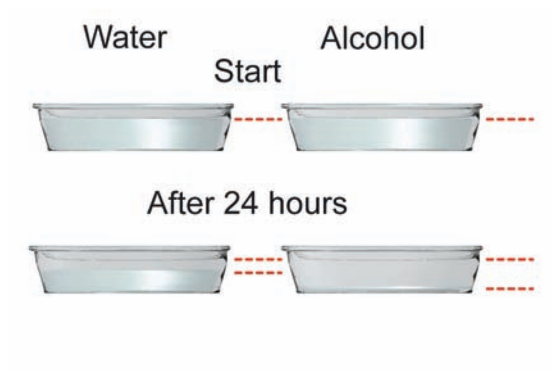
Solve: *First, let's calculate the energy that it takes to warm up a gram of ice from –25°C to 0 °C.*
Energy needed to go from T_1 *to* T_2 *is* $mc_p(T_2 - T_1) = (1\ \text{g})[2.0\ \text{J/(g·°C)}](25°\text{C}) = 50\ \text{J}$
So it takes 50 J to warm up 1 g of ice from –25°C to 0°C.
The same gram of ice takes 335 J to melt into liquid water.

Answer: *Changing phase (melting) absorbs 335 J per gram of ice but warming the ice only absorbs 50 J/g. The phase change is responsible for most of ice's cooling effect on drinks!*

Evaporation

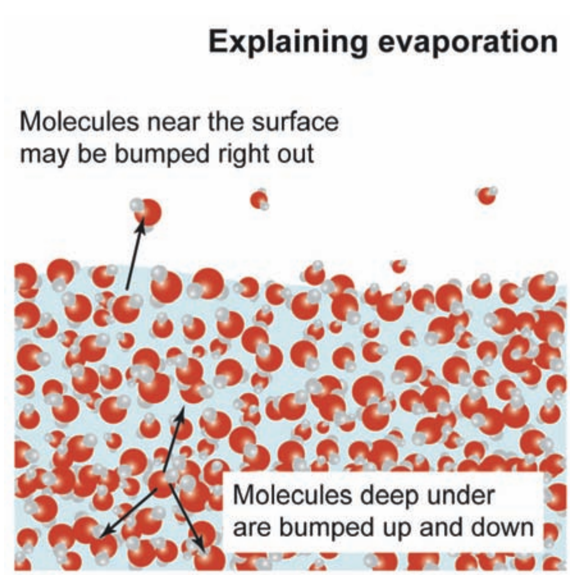
Why does cold water evaporate?

Think about wiping a surface with a wet cloth. Over time the water *evaporates*. In **evaporation**, liquid molecules change phase into gas at a temperature far below the boiling point! Virtually all liquids evaporate, some much faster than others. Dampen a surface with alcohol and it evaporates almost immediately. A surface wet with water stays wet for much longer. Why does evaporation occur?



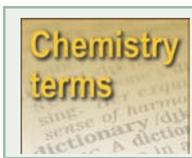
The average energy of molecules

Evaporation happens because there is a surface between liquid and gas. Think about water molecules at the surface compared to molecules deep below the surface. All molecules are in constant jostling motion. The deeper molecules are hit upward by their neighbors as often as they are hit downward. They stay in the water. The molecules on the surface, however, have no molecules on top to send them back down. When a surface molecule is jostled hard enough by a molecule below it, the surface molecule is ejected from the water and becomes a gas. This is the explanation for evaporation.



Evaporation cools liquids

How does an evaporating molecule get enough energy to break its connections with the other molecules and escape? The energy is taken away from the molecules just below that did the actual bumping! It actually takes the same amount of energy to evaporate a gram of water at room temperature as it does to boil a gram of water at 100°C! Evaporation transfers thermal energy from a liquid to a gas, and therefore it *cools* the remaining liquid. That is why your wet skin feels cold when you get out of the shower. The evaporating water from your skin takes heat away from your body. This cooling effect is also why you sweat while exercising. The evaporation of sweat from your skin removes heat generated by your muscles.

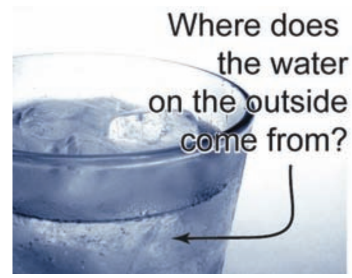


evaporation: a phase change from liquid to gas at a temperature below the boiling point.

Condensation

Condensation

Have you ever noticed that the outside of a glass of cold water becomes wet on a humid summer day? Where did the water come from? Did it come through the glass?

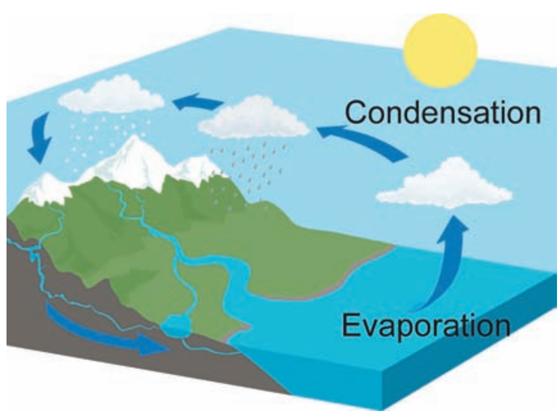


The water on the outside of the glass comes out of the air in a process called *condensation*. **Condensation** is the changing of phase from gas to liquid. Water vapor in the air condenses on a cold glass to become liquid. Condensation occurs below a substance's boiling point.

Water cycle

Although all liquids and gases can condense, the process is particularly important with water because evaporation and condensation create the water cycle of Earth. Condensation may occur when a substance in its gas phase is cooled to below its boiling point. Clouds are an excellent example. Water vapor evaporates from the ocean and rises with warm air. The upper atmosphere is colder, so the water vapor condenses into tiny droplets of liquid to make clouds. Clouds are not water vapor, which is a gas. Clouds are condensed liquid water. Evaporation and condensation are responsible for the water cycle in the atmosphere.

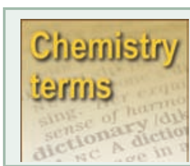
Condensation releases heat



Condensation is the exact opposite process of evaporation. The energy transferred during condensation is exactly the same, but in the opposite direction, as in evaporation. From Table 3.3 we know that the heat of vaporization of water is 2,256 J/g. When a gram of water evaporates, it carries away 2,256 J of energy, *cooling* any matter left behind. When a gram of water vapor condenses, it gives off 2,256 J of energy, *warming* any matter it condenses on.

Latent heat

The heat energy given off during condensation is often called the **latent heat**. Latent heat is thermal energy associated with a phase change. For evaporation and condensation, the latent heat is the heat of vaporization. For melting and freezing, the latent heat is the heat of fusion. The latent heat released during cloud formation transfers heat from the oceans to the upper atmosphere, warming the atmosphere significantly. This is a very significant component of Earth's energy balance.



condensation: a phase change from gas to liquid; a substance in its gas phase may condense at a temperature below its boiling point.

latent heat: thermal energy that is absorbed or released during a phase change.

How pressure affects phase changes

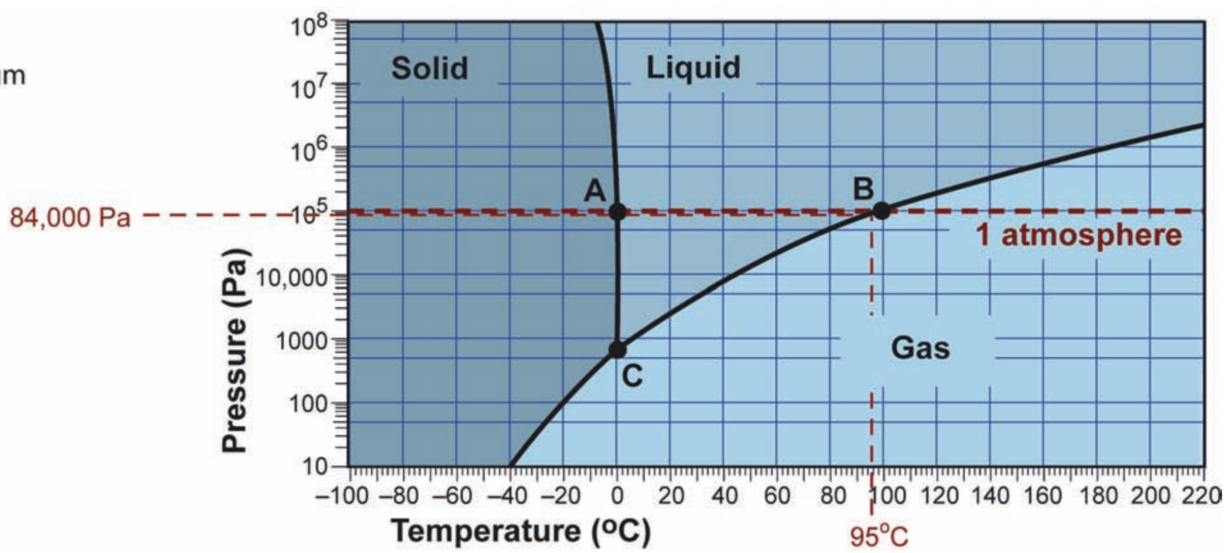
Water boils at a lower temperature at high altitude

Cookbooks often have special instructions to cook food for longer at high altitudes. This is because water boils at different temperatures depending on the air pressure. For example, in Denver, Colorado, water boils at 95°C (203°F) instead of the usual 100°C (212°F). The boiling point is lower because Denver is at an altitude of 1,610 m and the air pressure is lower than it is at sea level.

The phase diagram

The temperature at which a phase change occurs typically depends on pressure. The relationship among the phases of matter, temperature, and pressure is shown on a *phase equilibrium diagram*. Point A is the melting/freezing point which is 0°C at 101,325 Pa (1 atm). Point B is the boiling point, which is 100°C at 1 atm.

Phase equilibrium diagram of water



What the curves mean

Each curve represents a phase change. Crossing the liquid–gas curve means changing from liquid to gas or gas to liquid. This curve represents the boiling point. For Denver, the pressure is 84,000 Pa. Notice that the line for 84,000 Pa crosses the liquid–gas curve at 95°C. Crossing the solid–liquid curve means changing phase from solid to liquid or vice versa. Water freezes or melts at the same temperature from 800 to 200,000 Pa. At higher pressures, water freezes at a temperature below 0°C.

The triple point

Point C on the phase equilibrium diagram is called the *triple point*. The **triple point** of water is at exactly 273.16 K (0.01°C) and a pressure of 611.73 Pa (0.006037 atm). At the triple point, liquid, solid and vapor can all be present at the same time. A sample of water at its triple point can become solid, liquid, or gas with very small changes in pressure or temperature.

Chemistry terms

triple point: the temperature and pressure at which the solid, liquid, and gas phases of a substance can all exist in equilibrium together.

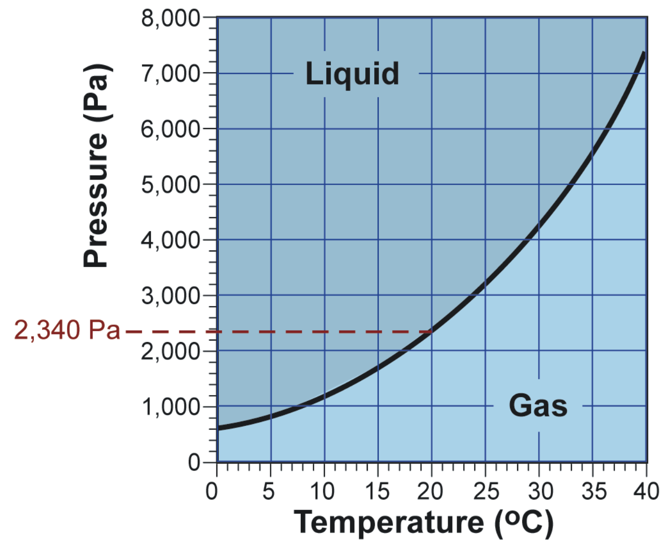
Relative humidity (Rh)

The meaning of humidity

A 90°F day in Houston, Texas, feels much more uncomfortable than a 90°F day in Phoenix, Arizona. The reason is the *relative humidity*. **Relative humidity** describes how much water vapor is in the air compared to how much water vapor the air can hold when it is completely saturated.

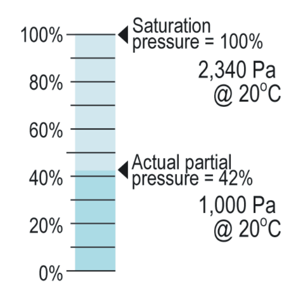
Reading the equilibrium vapor pressure

The diagram shows water's phase equilibrium diagram between 0°C (32°F) and 40°C (104°F). The curve represents the boiling point. The curve tells you that water boils at 20°C at a pressure of 2,340 Pa. The curve also represents the equilibrium partial pressure of water vapor in air. For example, at 20°C water vapor has a partial pressure of 2,340 Pa. This is called the *saturation* vapor pressure. Like any saturated solution, if more water vapor is added to air at 20°C, it condenses into liquid. This interpretation is correct because a liquid boils when its vapor pressure is equal to or greater than the ambient atmospheric pressure.



Calculating the relative humidity

Suppose some 20°C air contains less than 2,340 Pa of water vapor. This is an *unsaturated* solution. It means the air could absorb more water vapor without it condensing out. The relative humidity (Rh) is defined as the actual partial pressure of water vapor in air divided by the saturation vapor pressure at the same temperature. For example, the air in Phoenix is very dry, and even at 20°C it might contain only 1,000 Pa of water vapor. The relative humidity is 1,000 ÷ 2,340 = 0.42 or 42%.



The dew point

Another important concept is the *dew point*. The **dew point** is the temperature at which air is saturated with water. For example, suppose air has a partial pressure of 1,000 Pa of water vapor. On the diagram, this corresponds to a dew point of 8°C. At 8°C the relative humidity is 100%. If the air gets cooler than 8°C, the mixture becomes supersaturated and water contained in air will condense. Above 8°C, the relative humidity is less than 100%.

Chemistry terms

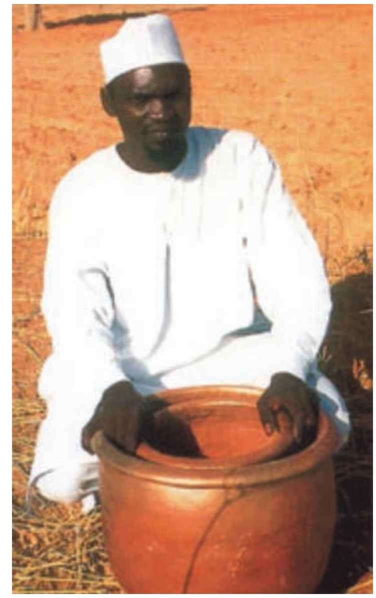
relative humidity: the actual partial pressure of water vapor in air divided by the saturation vapor pressure at the same temperature.

dew point: the temperature at which air is saturated with H_2O vapor (Rh = 100%).

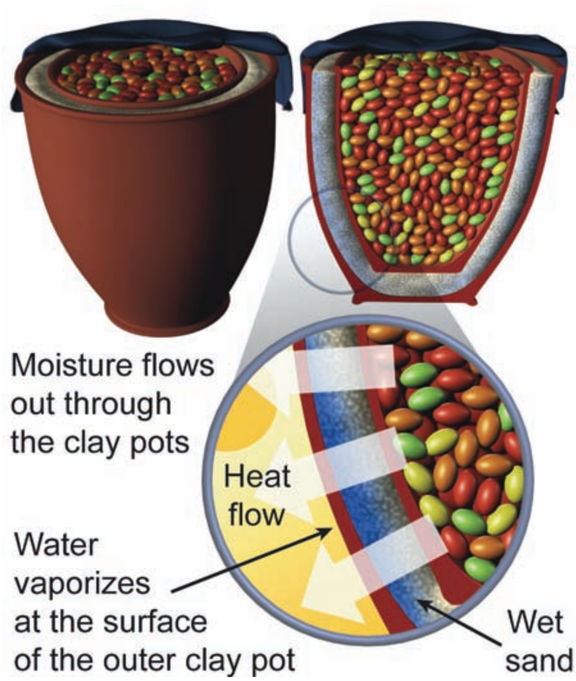
Refrigeration by evaporation

Chemistry Connections

Mohammed Bah Abba is a smart, young African teacher who designed a clever refrigerator that needs no electricity! His invention uses earthenware clay pots and the principles of evaporation. As a young man Abba had a fascination with clay pots and their ability to absorb water. He wondered how the clay could absorb so much water, yet still hold together. He had observed that most things that absorb water become soft and lose their shape. He was also curious as to why the clay pots kept things cool. In college, Mohammed took some classes in biology, chemistry and geology and with this background he developed the "pot-in-pot refrigerator."



The pot-in-pot refrigerator is designed using two clay pots of different sizes, one large and the other small. The small pot is placed inside the larger pot and sand is added to fill the space between the walls of the two pots. The sand is moistened with water and a damp cloth is used to cover the top.



How does the pot-in-pot refrigerator keep things cool? To explain this, we will use the concepts you have learned about the phase changes of water and why they occur. When water molecules evaporate, they absorb energy because they are going from the liquid phase to the gas phase. Do you recall why this requires energy? It is because the intermolecular forces holding the water molecules together in the liquid phase must be *broken* to allow the water molecules to separate and become a vapor:

H_2O (liquid) + energy $\rightarrow$ H_2O (gas)

The hot climate in Africa provides warm, dry air outside the pot, causing evaporation to occur on the walls of the large outer pot. The moisture from the damp sand is pulled toward the outer pot as evaporation takes place. The porous clay allows the water molecules to travel from the moist interior to the outer pot where they are warmed and evaporate. The warmth outside the pots provides the liquid water molecules with the energy they need to evaporate and change into water vapor.

The heat of vaporization (ΔH_v) for water is 2,256 J/g or 40.6 kJ/mole. This large amount of energy causes significant cooling that results in a temperature drop of several degrees! The more water molecules evaporate, the cooler the temperature becomes. This causes the inner pot to become cool. The cooler temperature of the inner pot preserves foods that would otherwise spoil quickly in the heat.

Throughout human history, single clay pots have been used to keep things cool. Usually the clay pots are soaked in water for 20 minutes or so and then removed. Throughout the next few hours the water absorbed by the clay slowly evaporates. This process is used by street vendors to keep their produce and drinks cool for their customers. The single pot design works, but it is less efficient and does not last as long as Mohammed's pot-in-pot design. His invention shows how important it is to use the ideas behind simple technology in an applied way! Mohammed wanted to help the poor people in the rural regions of northern Nigeria, and his inspiration gave way to this elegant, efficient pot-in-pot refrigerator.

To fully appreciate and understand the significance of Mohammed's invention it helps to understand the people of northern Nigeria. The geography of this area is semi-desert scrubland, and the rural people live in the poorest of conditions. There is no electricity and therefore no refrigeration. Even in the towns, the power supply is not dependable. Most people cannot afford to buy electricity even if it is available. The villagers' way of life is based on agriculture, and young girls have the responsibility of rapidly selling the food that is grown each day before it spoils. Mohammed's invention allows these young girls to attend school more often instead of having to sell food.

How well does it work? Mohammed did many tests with different fruits and vegetables and found the following results: Eggplant stayed fresh for 27 days instead of three, tomatoes and peppers lasted for three weeks or more, and African spinach, which usually lasts only a day, remained edible for 12 days.

Mohammed worked to refine his invention for two years before beginning to distribute the pot-in-pot. He hired local pot makers to produce the first 5,000 pot-in-pots, which he distributed for free to five local villages in Jigawa. Later, with some limited financial backing, he distributed 7,000 more to another dozen local villages. Educating the illiterate local villagers on how to use the pot-in-pot was a challenge that Mohammed overcame by using a video-recorded play of local actors who acted out how to use the pot-in-pot and demonstrated how food stayed fresh for a much longer time. As a result of the pot-in-pot, the farmers' income levels have risen because they can now sell their produce to consumers based on demand, where before they had to sell quickly because the food would spoil. The number of girls attending local village schools has risen, because they no longer need to sell produce each day to survive. The pot-in-pot has also given rise to healthier food for all people to consume, therefore decreasing disease and illness in these areas.

Mohammed Bah Abba was only in his twenties when his invention won him the Rolex Award for enterprise in 2000. He won the prestigious award not only for his design of the pot-in-pot refrigerator but for his efforts to make it affordable and available to the people who really needed it. These pots sell for thirty to forty cents a set!

Chapter 3 Review

Vocabulary

Match each word to the sentence where it best fits.

Section 3.1

kinetic energy	thermometer
Brownian motion	thermistor
Fahrenheit	absolute zero
random	Kelvin
temperature	Celsius

1. The erratic movement of a small particle in water due to constant collisions of molecules is called _______________.

2. _______________ is a measure of the average thermal energy of atoms or molecules.

3. The _______________ temperature scale, used in the United States, has a boiling temperature of water at 212.

4. The energy of motion is referred to as __________.

5. A device used to measure temperature is called a/an _______________.

6. The theoretical temperature where all motion stops is called _______________.

7. The SI derived unit of temperature is called the _______________. When using this unit, water freezes at zero degrees.

8. Motion that is scattered equally in all directions is called _______________ motion.

9. A/An _______________ is a temperature sensor that detects changes in electrical properties as the temperature changes.

10. The temperature scale that starts at absolute zero is called the _______________ temperature scale.

Section 3.2

specific heat	system
joule	heat
calorie	first law
thermal equilibrium	conductor
second law	insulator

11. _______________ is a measure of the thermal energy of a substance.

12. A group of interacting objects and effects that are selected for investigation is called a/an _______________.

13. The _______________ of thermodynamics states that thermal energy moves from higher temperature to lower temperature.

14. The law of conservation of energy is also called the _______________ of thermodynamics.

15. The units of _______________ and _______________ are used to measure heat energy. The unit of the _______________ is larger than the unit of the _______________.

16. When the temperature of two systems is the same and heat no longer flows between them, they are in _______________.

17. A material that easily permits the flow of heat across it is called a/an _______________.

18. The amount of heat needed to raise one g of a substance by one degree is called the _______________.

19. A material that does not permit the flow of heat across it is called a/an _______________.

Section 3.3

gas	heat of fusion
solid	melting point
liquid	boiling point
triple point	phase change
vaporization	dew point
evaporation	

20. In the _______________ phase of matter, the atoms or molecules are able to move and change positions, but are still touching each other.
21. The process of _______________ requires a great deal of energy, because it breaks the bonds that hold molecules together in their liquid phase and separates them.
22. In the_____________ phase of matter, the atoms or molecules are not able to move or change positions.
23. The amount of heat required to melt one mole of ice is called the _______________.
24. Liquids kept in an uncovered container slowly lose volume over time through the process of _______________.
25. A substance undergoes a/an _______________ when its molecules are rearranged without changing the individual molecules.
26. The temperature at which a substance changes from a solid to a liquid is called the ___________.
27. The temperature at which the vapor pressure of a liquid is equal to the atmospheric pressure is called the ___________________.
28. The _______________ phase of matter has no definite volume.
29. The _______________ is the temperature at which the partial pressure of water vapor in a sample of air equals its saturation vapor pressure.
30. The _______________ is the unique value of pressure and temperature at which the solid, liquid, and gas phases coexist in equilibrium.

Conceptual Questions

Section 3.1

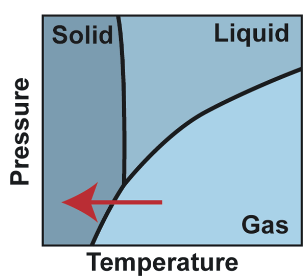
31. Explain why Brownian motion provides evidence that molecules are in constant random motion.
32. Describe the differences between the Fahrenheit and the Celsius temperature scales.
33. What units do we use to measure temperature?
34. How is the kinetic energy of a system related to its temperature?
35. Describe how the random motion of atoms and molecules is related to temperature.

Section 3.2

36. To calculate the heat absorbed by an object, what information is needed besides its specific heat?
37. Objects tend to exchange heat until "thermal equilibrium" is reached. What does this mean? Explain your reasoning.
38. Does the specific heat change when you have a greater mass of a substance, for example, 10.0 g versus 100.0 g? Explain.
39. Copper has a specific heat of 0.34 J/(g·°C). What does this tell you about the structure of copper? (Explain what the unit of specific heat means.)
40. Adding heat to a substance can cause its temperature to increase. Explain why.
41. The specific heat of water is much higher than that of other substances. What does this mean?
42. You have 20.0 g samples of two metals, one with a high specific heat and one with a low specific heat. Upon heating them with the same amount of heat, which one will have the greatest change in temperature? Explain.
43. Suppose you have a flat, hot bar of iron that weighs 25.0 g at 90°C, and you place a 10.0 g bar of copper at 25°C on top of it. What will happen at the molecular level?
44. Tropical islands often have what is referred to as temperate climates, which means the daily temperature is generally the same. It does not get really hot or really cold. Since islands are surrounded by water, explain why the temperature remains relatively constant.

Section 3.3

45. Make a sketch of the heating curve for water. Label your axes appropriately, and label the solid, liquid, and gas phases.
46. Why are there flat places on your sketch for the heating curve of water?
47. Explain why the heat of fusion is always less than the heat of vaporization for any substance.
48. Explain why your skin feels cool when you sweat.
49. Some substances are liquids at room temperature. What does this tell you about the substances' melting points?
50. Two identical glasses full of water are left in two different rooms. One room is maintained at a temperature of 10°C and the other at 50°C. What would you find when you come back a week later?
51. What phase transition does the red arrow represent in the phase equilibrium diagram shown here?



52. Describe the process of evaporation using simple words.
53. Water is the substance with the highest values of heat of fusion (335 J/g) and heat of vaporization (2,256 J/g). Describe the possible role that these large values have for life on Earth.
54. Suppose the air is at a relative humidity of 100%, and then the temperature goes down.
    a. Does the partial pressure of water vapor change? Why or why not?
    b. If the partial pressure changes, what happens to the lost water vapor?
    c. What is the relative humidity at the new temperature?

Quantitative Problems

Section 3.1

55. The normal temperature for a human body is about 98.6°F. Calculate this temperature in Celsius.
56. Ethylene glycol is a liquid organic compound, known as antifreeze. It freezes at –12°C. Calculate the freezing temperature on the Fahrenheit scale.
57. Paper catches on fire at about 451°F. What is this temperature on the Celsius scale?
58. Liquid nitrogen is used as a commercial refrigerant to flash freeze foods. It boils at –196°C. What is this temperature in Fahrenheit?
59. Convert –20°C to Kelvin.
60. Convert 273 K to degrees Fahrenheit.

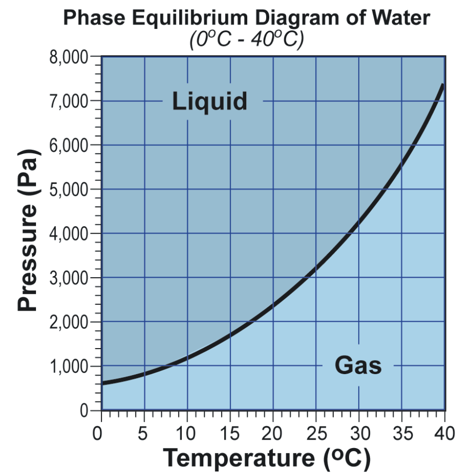
Section 3.2

61. How many joules (J) are needed to increase the temperature of 15.0 g of lead from 20°C to 40°C? [Use C_p of Pb = 0.128 J/(g·°C).]
62. An unknown metal substance absorbs 387 J of energy. If that substance weighs 50.0 g and changes temperature from 30°C to 50°C, what is the metal? (Refer to your table of specific heats.)
63. A 5.0 g piece of gold (Au) changes temperature from 22 °C to 80 °C. How much heat does the gold absorb? [Use C_p of Au = 0.129 J/(g·°C).]
64. A 6.22 kg piece of copper metal is heated from 21.5 °C to 310 °C. Calculate the heat absorbed in kilojoules by the metal. [Use C_p of Cu = 0.34 J(g·°C).]
65. If 856 J of heat was absorbed by a block of iron that weighs 55.0 g, calculate the change in temperature of the block of iron. [Use C_p of Fe = 0.44 J/(g·°C).]
66. An 80.0 g mass of a metal was heated to 100°C and then plunged into 100 g of water at 24°C. The temperature of the resulting mixture was 30°C. [Use 4.18 J/(g·°C) for the specific heat of water]
    a. How many joules of heat did the water absorb?
    b. How many joules of heat did the metal lose?
    c. What is the specific heat of the metal?

67. One beaker contains 145 g of water at 22°C and a second beaker contains 76.5 g of water at 96°C. What is the final temperature of the water after the beakers are mixed?
68. A 20.5 g piece of zinc was heated to 99.5°C in boiling water. It was then added to a beaker containing 50.0 g of water at 23 °C. When the water and the metal come to thermal equilibrium, the temperature is 25.8°C. What is the specific heat of zinc?

Section 3.3

69. What quantity of heat is released (evolved) when 1.5 L of water at 0°C solidifies to ice at 0°C?
70. The heat energy required to melt 1.0 g of ice at 0°C is 335 J. If one ice cube has a mass of 58.0 g, and you have 12 ice cubes, what quantity of energy is required to melt all of the ice cubes to form liquid water at 0°C?
71. What quantity of heat is required to vaporize 110 g of benzene (C_6H_6) at its boiling point? Benzene's boiling point is 80.1°C and its heat of vaporization is 30.8 kJ/mole.
72. A teacher says the following in a lesson. "The specific heat of water is the highest of all common substances. With a value of 4.18 J/(g°C), water is very important in regulating the temperature of the earth. It takes a lot of energy to change the temperature of water by one degree Celsius. This is true for increasing or for decreasing the temperature of water. The large amount of water that covers the surface of the Earth, provides a very strong buffer both for rasing and for lowering the temperature of the atmosphere."
Choose the most appropriate notes you might take:
    a. Water has high specific heat.
    b. Water maintain the temperature of the Earth
    c. H_2O has high specific heat.
    Takes a lot of energy to change the temp. of H_2O.
    Water helps to maintain constant Earth temp.
73. What quantity of heat energy, in joules, is required to raise the temperature of 1.50 kg of ethanol from 22.0°C to its boiling point, and then to change the liquid to vapor at that temperature? The boiling point of ethanol is 78.3°C. [The specific heat capacity of liquid ethanol is 2.44 J/(g·°C) and the heat of vaporization for ethanol is 855 J/g.]]



Use the phase equilibrium diagram above to solve problems 74–78.

74. What is the partial pressure of water vapor in saturated air at 15°C?
75. The water vapor in a sample of air has a partial pressure of 4,400 Pa. At what temperature is this air saturated (what is its dew point)?
76. Some warm air has a temperature of 30°C and a water vapor partial pressure of 2,000 Pa. This air is cooled to 10°C as night falls. Will some liquid condense or not? Explain how you get your answer.
77. Desert air at 35°C has a partial pressure of water vapor of 1,200 Pa. What is the relative humidity?
78. Which has more water vapor?
    a. saturated air (100% relative humidity) at 15°C
    b. air at 28°C with a relative humidity of 50%.
79. A student is planning an investigation to determine if the boiling point of salt solution increases with concentration. To do the investigation and understand the results, which of the following is the most important reason to select a computer?
    a. to email the conclusions to colleagues or the instructor
    b. to collect and graph the data with a spreadsheet program
    c. to record a video of the experimental procedure
    d. to write the lab report using a word processing program