SCOPE, SEQUENCE, and COORDINATION

A National Curriculum Project for High School Science Education

This project was funded in part by the National Science Foundation. Opinions expressed are those of the authors and not necessarily those of the Foundation.

SCOPE, SEQUENCE, and COORDINATION

SS&C Research and Development Center
Gerry Wheeler, *Principal Investigator*
Erma M. Anderson, *Project Director*
Nancy Erwin, *Project Editor*
Rick McGolerick, *Project Coordinator*
Arlington, Va., 703.312.9256

Evaluation Center
Frances Lawrenz, *Center Director*
Doug Huffman, *Associate Director*
Wayne Welch, *Consultant*
University of Minnesota, 612.625.2046

Houston SS&C Materials Development and Coordination Center
Linda W. Crow, *Center Director*
Godrej H. Sethna, *School Coordinator*
University of Houston-Downtown, 713.221.8583

Houston School Sites and Lead Teachers
Jefferson Davis H.S., Lois Range
Lee H.S., Thomas Ivy
Jack Yates H.S., Diane Schranck

California Coordination Center
Tom Hinojosa, *Center Coordinator*
Santa Clara, Calif., 408.244.3080

California School Sites and Lead Teachers
Sherman Indian H.S., Mary Yarger
Sacramento H.S., Brian Jacobs

Iowa Coordination Center
Robert Yager, *Center Director*
University of Iowa, 319.335.1189

Iowa School Sites and Lead Teachers
Pleasant Valley H.S., William Roberts
North Scott H.S., Mike Brown

North Carolina Coordination Center
Charles Coble, *Center Co-Director*
Jessie Jones, *School Coordinator*
East Carolina University, 919.328.6172

North Carolina School Sites and Lead Teachers
Tarboro H.S., Ernestine Smith
Northside H.S., Glenda Burrus

Puerto Rico Coordination Center*
Manuel Gomez, *Center Co-Director*
Acenet Bernacet, *Center Co-Director*
University of Puerto Rico, 809.765.5170

Puerto Rico School Site
UPR Lab H.S.

* * * * * * * * * * * *

Pilot Sites
Site Coordinator and Lead Teacher
Fox Lane H.S., New York, Arthur Eisenkraft
Georgetown Day School, Washington, D.C., William George
Flathead H.S., Montana, Gary Freebury
Clinton H.S., New York, John Laffan*

*not part of the NSF-funded SS&C Project.

Student Materials

Learning Sequence Item:

942

Observing Chemical Reactions

March 1996
Adapted by: Michele Mallardi

Contents

Lab Activities

Readings

This micro-unit was adapted by **Michele Mallardi** (Jefferson Davis H.S., Houston, Texas)

Science as Inquiry

The Repeating “Exploding” Flask

What causes the explosion?

Objective:

Why do some chemical reactions produce explosions?

Procedure:

Observe the demonstration using alcohol performed by your teacher.

1. What are some ways of knowing that energy is being produced? Try to name three.
2. Is this reaction producing or using energy? How do you know?
3. What is the average number of explosions in five minutes?
4. The coil is not submerged in the alcohol so how can it be reacting with the alcohol? What did you do to make sure that they would come in contact?

Science as Inquiry

The Heat Is On

How does the temperature vary during a reaction?

Overview:

Measure the water temperature increase during a reaction.

Procedure:

Measure 100 mL of water and pour into the Styrofoam™ cup. Record the temperature of the water. Measure 10–15 g of calcium chloride. Pour the calcium chloride into the water. Record the temperature of the reactants every minute until the temperature levels off. Graph your data.

1. What are three ways you can tell that energy is being released during a chemical reaction?
2. Did the reaction absorb energy or release energy? How do you know?
3. During what five-minute interval of time did you measure the greatest temperature change?
4. Why do you think the greatest temperature change occurred during that time?
5. Plants can absorb energy and release energy. Explain how.

Student Sheet

Science as Inquiry

Disc O' Inferno

How can water cause a reaction?

Objective:

Mix together several chemicals and add a few drops of water.

Procedure:

Place a 5-inch square of aluminum foil on a metal sheet. Do not rest the foil on wood or the desk.

Obtain a soda cap containing ammonium nitrate and place it on the aluminum foil. Using a spatula, sprinkle a pinch of zinc dust on top. Use another spatula to sprinkle a pinch of iodine on top of this. *(Do not touch iodine with your hands.)*

Now add three drops of water. *Warning: the mixture may catch fire—stand back!*

After the reaction ceases, use tongs to dump the soda cap into a beaker of water. Dispose of the contents as directed by your teacher.

1. Was energy absorbed or released during the reaction? How do you know?
2. Is heat a product or a reactant in this chemical reaction?
3. Will the reaction run without iodine? What changes?
4. Predict what would happen if you added twice as much ammonium nitrate.

Science as Inquiry

Caught Red-Handed

How are hand warmers made?

Purpose:

Make your own hand warmer by mixing chemicals.

Procedure:

Place about 25 g of iron powder in a small plastic bag. Add 1 g of sodium chloride. Close the bag and shake to mix the chemicals. Add about 1 tbsp. of vermiculite and mix again. The hand warmer is now ready to be activated. Add 5 mL of water to the bag and seal it with the twist tie or zipper lock. Squeeze and shake the bag to thoroughly mix the contents. Be careful not to break the plastic.

1. Is this an exothermic or endothermic reaction? How do you know?
2. Is heat a reactant or a product in this reaction?
3. Do you think more energy is *(a)* being used to break the chemical bonds or *(b)* being made by bonding together the new compounds (product)? Explain your answer.
4. Does this reaction use or release oxygen? What can you do during this experiment to prove your answer?
5. What is the function of vermiculite in this experiment? Can the reaction work without the vermiculite?
6. How is this reaction similar to rusting?
7. Is there a commercial use for this reaction?

Science as Inquiry

Stuck Between a Beaker and a Hard Place

What happens when reactions absorb energy?

Overview:

In this demonstration, two chemicals will be mixed to produce an endothermic reaction.

Procedure:

Observe the demonstration of the endothermic reaction.

1. What happened to the wood block? How did that happen?
2. Did the reaction absorb energy or release energy? Explain the flow of energy.
3. They say if you put your tongue on a flagpole on a day when the temperature is freezing, your tongue will stick. Based on this lab do you think that is true or false? Be sure to explain why.

Science as Inquiry

Chilling Tales

How are cold packs made?

Overview:

Can you see how the product in this activity is similar to that in a cold pack?

Procedure:

Pour 100 mL of water into a Styrofoam™ cup. Take the temperature of the water. Measure 15 g of ammonium nitrate and pour the solid into the water. Note the temperature of the water at two-minute intervals. Graph your data.

1. Is this reaction exothermic or endothermic? How do you know?
2. Is more energy being used to break the chemical bonds or is more energy being produced by chemical bonding?
3. Why does the temperature drop?
4. In what six-minute interval of time did you notice the greatest temperature change?
5. What is similar in this reaction to the reaction in a cold pack? What is different?

Science in Personal and Social Perspectives

Hot and Cold Packs

It is late in the fourth quarter of the football championship, and the score is tied. Thirty seconds remain. With second down and goal to go, the home team calls its last time out. Robert, the quarterback who led the team all season, has sprained his thumb, and it's too painful to throw the ball. The trainer quickly takes a white plastic bag from his pocket, gives it a sharp punch with his fist, and applies it to Robert's injured hand. Within seconds the bag is ice cold.

The coach briefs the second-string quarterback and sends him into the game. The replacement calls a pass play, but is sacked even before he can raise his arm to throw. Third and goal to go . . . fifteen seconds left. The coach looks at Robert, who nods and drops the plastic bag. His thumb is numb enough for one more play.

Back in the game, Robert takes the snap, fakes right, steps out of the pocket, then runs to his left. Three defenders rush him. He throws the ball, missing the defenders' outstretched hands by inches. The receiver catches it in the end zone. Touchdown! Seconds later the clock runs out . . . it is the winning score.

Cold in a Bag

The plastic bag that the trainer used to cool the quarterback's thumb is an "instant cold pack." It does not need refrigeration and can be stored for months in a first aid kit, yet it produces cold the moment it is needed. How does the instant cold pack work? As shown in Figure 1, the pack has two sealed bags, one inside

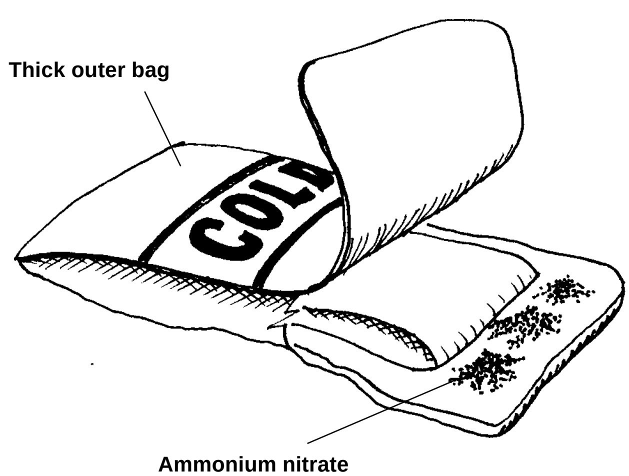


Figure 1. Jack Frost brand instant cold pack. The ammonium nitrate crystals and a plastic bag of water are contained inside a heavier plastic bag. Punching the pack bursts open the inner bag allowing the water and ammonium nitrate to mix.

the other. The outer bag is made of thick plastic and is relatively strong. It contains two things: a white powder, and a second plastic bag. The inner bag is made of weaker plastic and contains water. When the trainer punches the pack the inner bag breaks, and the water mixes with the powder. As the water dissolves the powder—a substance called ammonium nitrate—the solution becomes very cold. The reaction couldn't be simpler: a powder dissolving in water. This particular dissolving reaction absorbs heat, which is a technical way of saying it gets cold. (For why it gets cold, see "The Cold Facts"). Other compounds that get cold on dissolving in water are potassium nitrate, potassium chloride, and, to a very slight extent, plain old table salt (sodium chloride). Some other salts, such as sodium hydroxide, give off heat on dissolving. A reaction that absorbs heat is called an *endothermic* reaction, whereas heat-producing reactions are called exothermic.

Hot Stuff

Most familiar chemical reactions give off heat. Light a match. Hot, isn't it? Where does the

Continued

The Cold Facts

Ammonium nitrate (NH_4NO_3), is classified as a salt. Chemically speaking, there are thousands of salts in addition to sodium chloride, common table salt. Salts contain ions, particles with electrical charges. Because the ions with positive charge are strongly attracted to those with negative charge, they form a solid crystal.

In the first step, the solid crystal separates into ions. Breaking the ionic bonds requires a lot of energy, which means that heat must be absorbed from the surroundings. In the second step the water molecules, which are attracted to the charged ions, attach themselves to the ions. This step releases energy, which means that heat flows to the surroundings. The steps can be written like this:

Step 1:
$NH_4NO_3 + \text{Heat} \rightarrow NH_4^+ + NO_3^-$

Step 2:
$NH_4^+ + NO_3^- + H_20 \rightarrow NH_4^+(H_2O)_X + NO_3^-(H_2O)_X + \text{Heat}$

Several water molecules may bond to each ion, as indicated by $(H_20)_X$. In the first step, heat is absorbed; in the second step, heat is released. Overall, because more heat is involved in the first step than in the second step, heat is absorbed from the surroundings (6 kilocalories per mole of ammonium nitrate). This leaves the surroundings with less thermal energy— colder.

heat energy come from? It wasn't in the match before you struck it, was it?

Yes it was. It was *stored* in the match—in the various chemicals that make up the match. When the match burned, a series of vigorous reactions took place. Combustion occurs in many steps. To break the original bonds, heat must be absorbed. As new bonds are formed, heat is released. In this case, the amount absorbed is less than the amount released. Overall, heat is given off—the surroundings get hot.

Not all exothermic reactions are as vigorous as a burning match. Instant hot packs use slower reactions that take place at lower temperatures. The "Heat Factory" is a brand of hot pack that is sold at many camping stores. It has an outer plastic envelope and an inner paper bag perforated by minute holes (see Figure 2). The paper bag contains a mixture of powdered iron, sodium chloride, activated charcoal, and sawdust, all dampened with water. Remove the envelope from the outer plastic bag and shake it vigorously. It gets hot! What's going on here?

Everyone knows what happens when an iron shovel is left out in the rain for a couple of days—it rusts. The chemical reaction of iron and oxygen (oxidation), produces iron (III) oxide, or rust.

$$4Fe + 3O_2 \rightarrow Fe_2O_3 + Heat$$

In this reaction, ionic bonds are formed between iron and oxygen and heat is released (197 kilocalories per mole of iron (III) oxide). The rusting goes faster if the iron is wet, and faster still if the iron is wet with a salt solution. The shovel left out in the rain rusts too slowly for the heat to be noticeable. In the Heat Factory, though, the ingredients are mixed in precise proportions and ground up finely to make the oxidation go much faster. The Heat Factory is activated by shaking the envelope to get the oxygen in the air circulating through the small holes. The heat is the result of fast rusting.

Another brand of heat pack, called the "Heat Solution," works on a different principle. The Heat Solution looks like a small air mattress and is filled with a liquid the consistency of honey (see Figure 3). To activate it, you squeeze a special compartment in one corner of the pack, which releases a triggering crystal. The liquid then gradually solidifies and gives off heat for several hours.

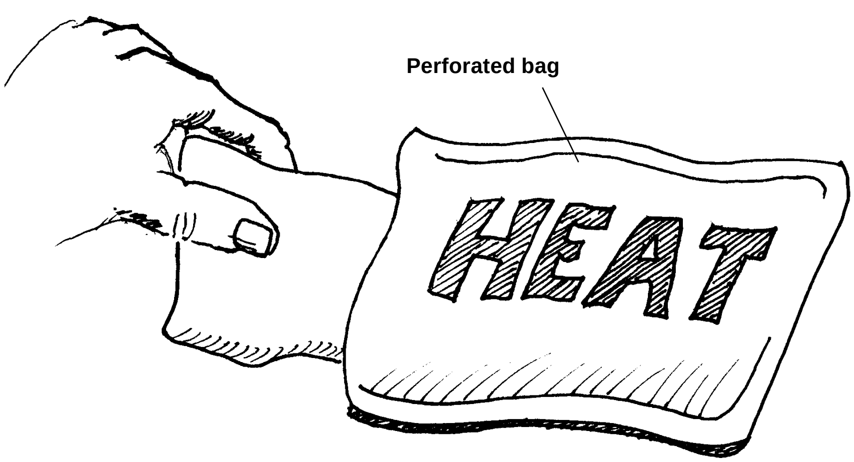
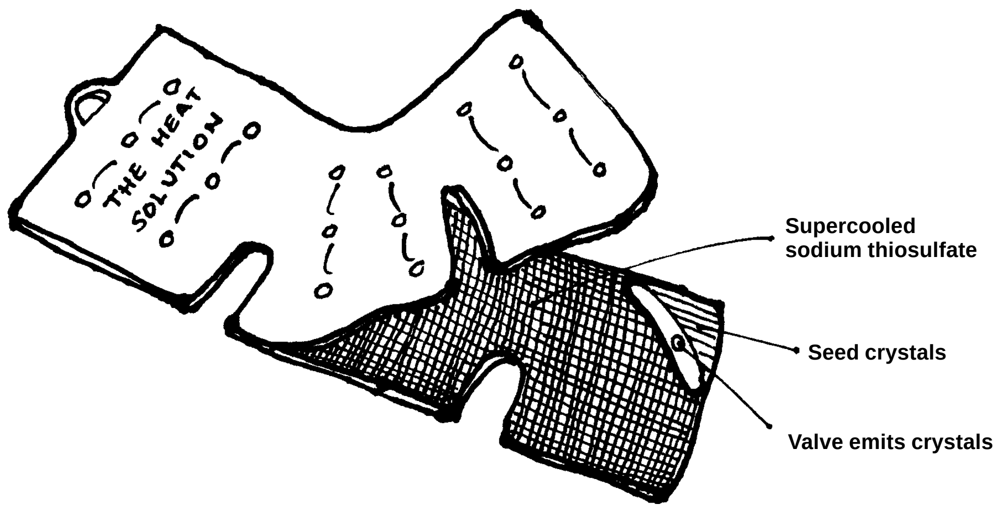


Figure 2. An instant heat pack. Remove the inner bag from the plastic envelope and shake it to start the heat-releasing reaction. Because the reaction needs oxygen from the air, you can stop it by returning it to the airtight envelope, then restart it later.

Figure 3. A constant temperature heat pack. A triangular pocket contains seed crystals that can be released through a valve by squeezing the corner. This triggers the crystallization that warms the pack and maintains a comfortable 48 °C (118 °F).



This heat generator uses a *phase change* instead of a chemical reaction. We are all familiar with the three common phases of matter: solid, liquid, and vapor. The most common example is water, which can exist as ice, liquid water, or steam—same chemical, different phases. Like other substances, water has specific temperatures at which it changes phase—boiling point (100 °C) and freezing point (O °C).

Under certain conditions, a phase can exist outside of the normal temperature limits. Water, for example, can be cooled below O °C. Meteorologists have learned that, high in the atmosphere, the tiny droplets of water in clouds may be as cold as - 30 °C and still be liquid. When a chemical remains liquid at temperatures below its normal freezing point, it is called a *supercooled* liquid.

What does this have to do with hot packs? The liquid in the Heat Solution contains supercooled sodium thiosulfate. To make the supercooled liquid solidify, a *seed crystal* is needed. This is simply a small piece of solid sodium thiosulfate around which more solid can crystallize. When a seed crystal is added, it triggers the change from supercooled liquid to solid. As the sodium thiosulfate solidifies around the seed crystal, the pack heats up. The heat is the result of bonds being formed as the substance crystallizes. The temperature rises to the freezing point of the sodium thiosulfate, a pleasingly warm 48 °C (118 °F). The valuable feature of phase change systems is that they can’t overheat. When a supercooled substance crystallizes, the temperature rises to the freezing point and stays there. It goes no higher or lower until all of the material has solidified. Notice that the inventors picked their chemical carefully. Unlike many compounds that cannot be supercooled at all, sodium thiosulfate supercools easily, and

The Tendency to Mess Up

Portable hot and cold packs depend on reactions that are *spontaneous*. Because the packs must be quick and easy to use, they require reactions that begin as soon as the reactants are placed together and that will continue on their own. Most spontaneous chemical reactions are exothermic—they give off heat. This is because chemical bonds have a tendency to shed their stored energy and release it as heat. The people who designed hot packs found that this natural flow of energy suited their needs perfectly. They selected the appropriate reactants, put them in the same package, but kept them separated. When warmth is needed, the reactants are simply mixed, and heat is produced automatically.

The tendency of stored bond energy to emerge as heat would seem to rule out cold packs. Because endothermic reactions absorb heat, the bonds end up with more stored energy than they started with—which is against the natural flow of things. Yet, this occasionally happens. When ammonium nitrate dissolves in water it gets very cold—spontaneously. Why does this occur? Scientists explain it with a concept called *entropy*.

Entropy is the degree of disorder in a system. Chemical changes tend to go from orderly arrangements of molecules and ions to disorderly arrangements. Nature tends to increase the amount of messiness, or disorder, or entropy.

The natural tendency to increase entropy sometimes opposes the tendency to release heat. When the increase in entropy is great enough, it can drive the heat flow "backward." The drive for high entropy overpowers the drive to release heat. Endothermic reactions happen spontaneously *only* when the reaction permits a large increase in entropy.

In the case of the instant cold pack, the starting materials were highly ordered: The water was pure and sealed in its own container, and the ammonium ions and nitrate ions were arranged in an orderly pattern in solid crystals. The substances were sorted and organized—everything in its place. When the inner plastic bag was broken and the water dissolved the ammonium nitrate, the orderly arrangement of the ions was disrupted. The ions were dispersed randomly throughout the water, and the once-pure water became "contaminated." Disorder reigned. The system went from very ordered to very disordered, and the reaction was partly driven by this increase in entropy.

its freezing point is comfortably warm, but not hot enough to burn the person who has a pulled muscle.

Unlike the other hot or cold packs, the Heat Solution is reusable. Simply heat the pack in boiling water for a while to return the sodium thiosulfate to its supercooled state; let it cool, and it's ready for the next emergency. The pack can be recycled until the supply of seed crystals is used up.

Bond Energy

We have examined three thermal first-aid packs. The instant cold pack uses a dissolving reaction (ammonium nitrate in water) that is endothermic. The Heat Factory uses an exothermic chemical reaction (iron and oxygen rust to iron [III] oxide). The Heat Solution uses an exothermic phase change (crystallization of supercooled sodium thiosulfate). Only in the Heat Factory "fast-rust" system was a new chemical compound formed. Nevertheless, the underlying theory is the same. Chemical processes always involve breaking and making bonds, which cause heat to be absorbed and released. It is the relative balance of heat change that determines whether the overall process feels hot or cold to the touch.

It is a week after the football game. Having won the championship, Robert and some friends are on a weekend camping trip. They had planned a lot of fishing but, at the lake, they mostly talk about the game, eat, tell jokes, and relax. What if it gets cold during the night, and the fire won't start? What if one of them strains a muscle chopping wood? Robert is not worried. In addition to the food, tents, and sleeping bags, he brought some chemical hot and cold packs.

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Science in Personal
and Social Perspectives

Fireside Dreams

You've just finished a day of vigorous skiing and, as the snow falls quietly outside in the darkness, you make a mug of hot chocolate, pick out a good book, and start a fire in the fireplace. You bring in some extra split logs so you won't have to make another trip outside. You relax on the rug and watch the fire for a moment as the wood starts burning.

Actually the wood itself does not burn. When you start a fire, the heat from the match or kindling breaks down the molecular structure of the wood and drives off vapors such as water, carbon monoxide, methane, hydrogen, methanol, and carbon dioxide. All these gases, except for water and CO_2, are combustible.

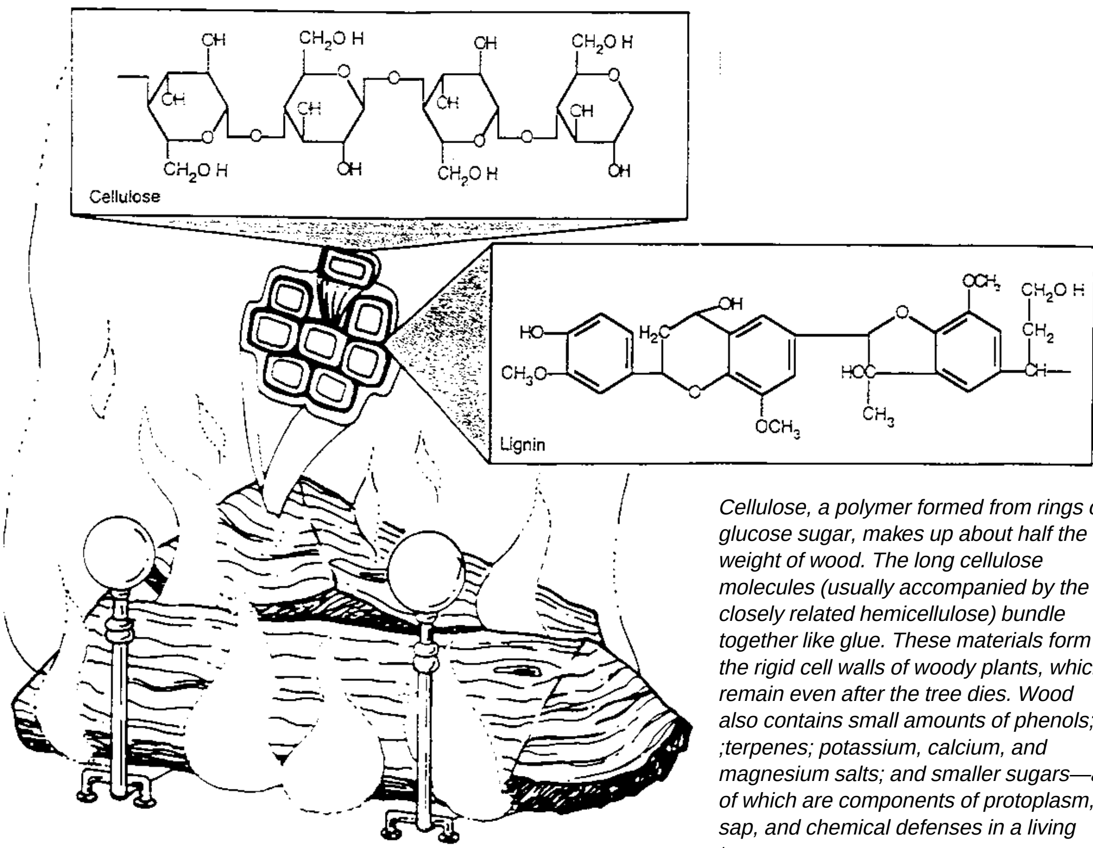
wood + heat → combustible gases + solid char

It is the mixture of gases that burns:

$2H_2 + O_2$ - $2H_20$ + heat

$2CO + O_2$ - $2CO_2$ + heat

$CH_4 + 20_2$ - $CO_2 + 2H_20$ + heat



Cellulose, a polymer formed from rings of glucose sugar, makes up about half the weight of wood. The long cellulose molecules (usually accompanied by the closely related hemicellulose) bundle together like glue. These materials form the rigid cell walls of woody plants, which remain even after the tree dies. Wood also contains small amounts of phenols; ;terpenes; potassium, calcium, and magnesium salts; and smaller sugars—all of which are components of protoplasm, sap, and chemical defenses in a living tree.

By Gail Marsella. Reprinted with permission from *Chem Matters (*December 1988) pp. 13–15. Copyright 1988 American Chemical Society.

The hot gases expand, become buoyant, and rise above the logs, causing a draft that sucks air into the fireplace, feeding the fire with more oxygen. Heat energy must be supplied to the wood to decompose it into combustible gases, but once the gases begin to burn, the wood fire pours out much more energy than it consumes. In these three reactions, the heat energy is all on the product (right) side of the equation.

The heat from the fire increases. The cat purrs contentedly on the rug. The clothes, wet from skiing, dry quickly. Lulled by the muted roar of the draft in the chimney, you start to doze, but it gets uncomfortable near the fire, and you move farther away.

Wood is a favorite fuel, in part, because its combustion products are so stable that a great deal of energy is liberated when it burns. The only reason wood doesn't burst into flame the minute you split the logs is that it needs a little energy input to get started, but it doesn't take much. The energy in a single match can start a ragging fire in dry wood, and that fire will not stop until either the wood or the oxygen runs out.

Wood is composed primarily of polymers-cellulose, hemicellulose, and lignin. When heated, the chemical bonds in these polymers absorb heat energy, become strained, begin to stretch, and finally break. A piece of the polymer flies off. When molecules break like this, they often form radicals, very reactive molecules and molecular fragments that have unpaired electrons. Much of the heat of the fire comes from the reactions in which radicals react with each other to pair up their electrons.

Other reactions occur, too. Let's say the piece of polymer that breaks off is a molecule of methane gas, CH_4. When the molecule is heated, its bonds absorb energy and stretch to the breaking point. An oxygen molecule can react with this "activated" methane to form carbon dioxide and water and release heat energy.

Much of the heat produced in the combustion process goes toward stretching and breaking other bonds and warming the resulting gas molecules to their reaction temperature. These new reactions produce more CO_2, H_2), and heat, making everything hotter still. In short, the fire's own heat keeps it going. It is the excess heat from this process that warms the room or goes up the chimney.

Because there isn't always enough oxygen for complete combustion, some of the molecules may burn only partly. Soot and hot vapors are swept up the chimney where they cool, condense on the chimney walls, and form creosote-a mixture of soot and gummy tars. This mixture collects inside the chimney and must be cleaned out by a chimney sweep.

Chimney Fires

Creosote is a combustible mixture of soot and tarry wood compounds that, if not periodically removed, may catch fire. A chimney fire is especially dangerous because the rising hot gases draw fresh air into the chimney's bottom at a rapid rate, transforming the chimney into a blowtorch. Temperatures may become high enough to damage the masonry of the chimney and even ignite wood that is in contact with the outside of the chimney.

Your eyes are getting tired of reading in the flickering light, so you put down you book and stare at the colors of the flames.

The top part of the flame is bright yellow. The color is caused by small carbon particles that break off of the fireplace grate and are swept upward by the draft. The particles are heated red-hot, then yellow-hot, and begin to react with oxygen to form CO, but the combustion is slow because most of the carbon is inside the solid particle, where it isn't exposed to oxygen. The particles glow yellow-hot, lending their color to the flame as they are carried upward.

If a flame has a lot of oxygen, if doesn't have much yellow color because the carbon particles burn to colorless carbon dioxide before they get a chance to glow. On the other hand, a fire that doesn't have enough oxygen has many carbon particles that don't burn completely, and these are swept as glowing yellow sparks up the chimney, where they cool and form black smoke.

The wood components that do not form gases remain on the fireplace grate as a black solid, called char. Some of the char burns slowly without visible flame-smoldering. Oxygen combines with carbon radicals on the surface of the char, forming carbon monoxide and then carbon dioxide, as well as heat, which causes the surface to glow red.

Near the base of the flame is an area of faint blue. When the gaseous radicals and molecules react to form their products and energy, part of the energy given off may take the form of light instead of heat. The energy produced by the combustion reaction excites electrons to higher energy levels. The energy may be passed from one electron to another, changed into energy of motion, or emitted as a particular wavelength of light. Near the base of the flame the reactions give off mostly blue wavelengths.

The red glow of the embers (char) and the yellow of the carbon particles occur in much the same way. Trillions of electrons from trillions of different hot atoms and molecules emit light of different color—all at the same time. The wavelengths of light that are most intense and most common determine the colors we actually see. The other colors may be there, but they are drowned out by the majority.

The glow of the embers dims. You consider putting another log on the fire, but decide against it. You have burned three logs, and all that remain are a few handfuls of ash, some dark streaks on the bricks, and the faint smell of smoke. The last ember goes out, and the room starts to cool off. You close the chimney flue so the cold outside air can't get in. You're finished with the book, the hot chocolate, and the fire, and it's time to crawl into the sleeping bag.

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